RS•C

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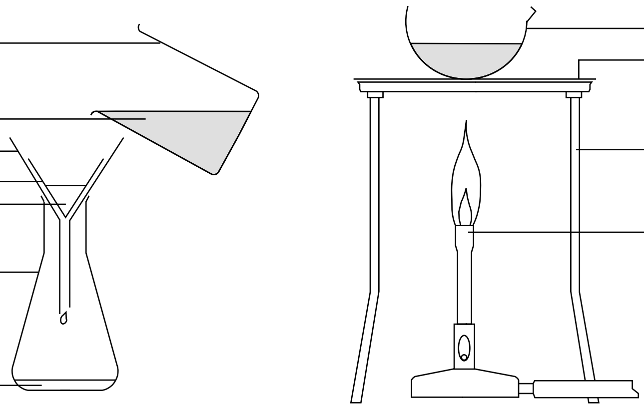
Separating a sand and salt mixture

Introduction

In this experiment simple processes are used to separate salt from a sand and salt mixture.



What to do

1. Mix about 5 g of the mixture with 50 cm^3 of water in a 250 cm^3 beaker. Stir gently.
2. Filter the mixture into a conical flask and pour the filtrate into an evaporating basin.
3. Heat the salt solution gently until it starts to 'spit'. **Care:** do not get too close.
4. Turn off the Bunsen burner and let the damp salt dry.

Safety

Wear eye protection.

Questions

1. Why is the salt, sand and water mixture stirred in step 1?
2. What happens when this mixture is filtered in the step 2?
3. Why is the salt heated in step 3?

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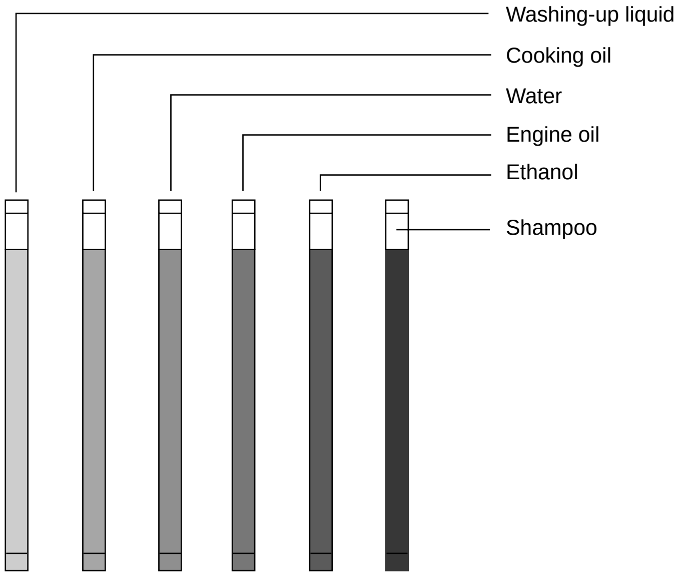
Viscosity

Introduction

The viscosity of a liquid is another term for the thickness of a liquid. Thick treacle-like liquids are viscous, runny liquids like water are less viscous. Gases exhibit viscosity in the same way. In this experiment, the viscosity of various liquids are compared.



What to record

Complete a table like this:

Liquid	**Time taken /s**
Washing up liquid	
Water	

What to do

1. Take one of the tubes provided.
2. Ensure the bubble is at the top and the tube is held vertical.
3. Quickly invert the tube and measure the time it takes for the bubble to reach the top.
4. Repeat this measurement for all the samples.

Questions

1. Which liquid is the most viscous?
2. Which liquid is the least viscous?
3. Design a different experiment for comparing the viscosity of liquids.

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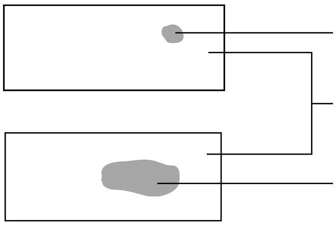
Rate of evaporation

Introduction

Evaporation is the conversion of a liquid into vapour, without necessarily reaching the boiling point. In this experiment the rate of evaporation is measured and compared under various different conditions.



What to record

Complete the following table.

Condition	Evaporation time (s)
Unspread, cool, air movement	
Unspread, cool, no air movement	
Spread out, cool, no air movement	
Spread out, warm, no air movement	
Unspread, warm, air movement	
Spread out, cool, air movement	
Spread out, warm, air movement	
Unspread, warm, no air movement	

What to do

1. Consider the following conditions for the evaporation of a drop of propanone on a microscope slide.

Condition	How achieved
Warm	Warm slide in hands and hold on a flat palm. Alternatively, place the slide in warm water then dry the slide.
Cool	Room temperature.
Spread out drop	Spread the drop of propanone on the slide with a matchstick.
Unspread	Drop left as one drop on the slide.
Cool air flow	Fan with book.
Warm air flow	Blow across drop.

2. Place a microscope slide in one of the conditions listed.
3. Add the single drop of propanone.
4. Measure the time for the drop to evaporate.
5. Repeat the experiment using different conditions.

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Safety

Ensure there are no sources of ignition nearby. Wear eye protection.

Questions

1. Name three factors that affect the rate of evaporation.
2. Why does evaporation produce a cooling effect?

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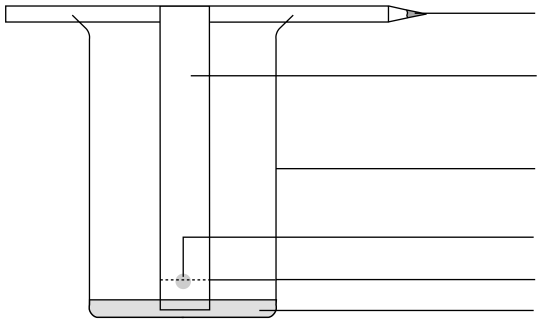
Chromatography of leaves

Introduction

Most leaves are green due to chlorophyll. This substance is important in photosynthesis (the process by which plants make their food). In this experiment, the different pigments present in a leaf are separated using paper chromatography.



What to record

The chromatogram produced in this experiment can be dried and kept.

What to do

1. Finely cut up some leaves and fill a mortar to about 2 cm depth.
2. Add a pinch of sand and six drops of propanone from the teat pipette.
3. Grind the mixture for at least three minutes.
4. On a strip of chromatography paper, draw a pencil line 3 cm from the bottom.
5. Use a fine glass tube to put liquid from the leaf extract onto the centre of the line. Keep the spot as small as possible.
6. Allow the spot to dry, then add another spot on top. Add five more drops of solution, letting each one dry before putting on the next. The idea is to build up a very concentrated small spot on the paper.
7. Put a small amount of propanone in a beaker and hang the paper so it dips in the propanone. Ensure the propanone level is below the spot.
8. Leave until the propanone has soaked near to the top.
9. Mark how high the propanone gets on the paper with a pencil and let the chromatogram dry.

Safety

Propanone is highly flammable. Wear eye protection.

Questions

1. How many substances are on the chromatogram?
2. What colours are they?
3. Which colour moved furthest?

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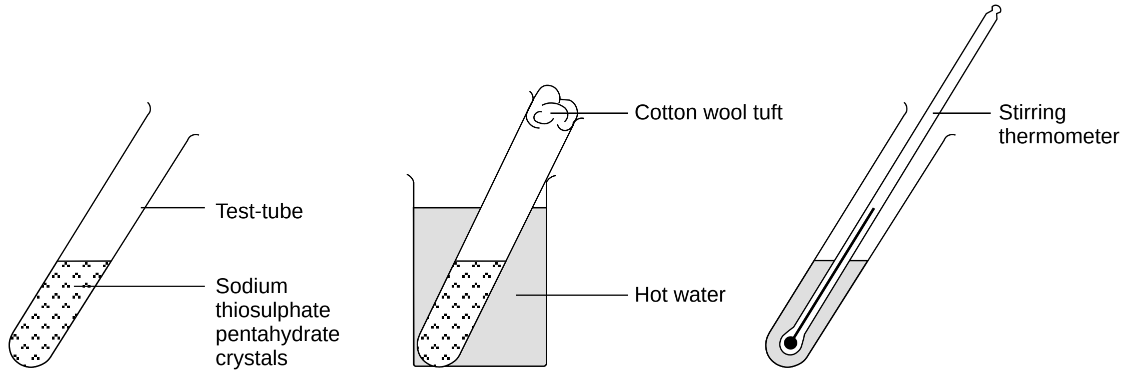
Energetics of freezing

Introduction

When a substance changes state, energy can be produced or absorbed. This experiment illustrates the energy change when a liquid freezes to form a solid.



What to record

Record the temperature of the liquid, record the temperature as the liquid solidifies (this is the melting point of sodium thiosulfate pentahydrate).

What to do

1. Half fill a test-tube with crystals of sodium thiosulfate pentahydrate.
2. Warm the test-tube gently in a beaker of hot water to melt the crystals.
3. Put a tuft of cotton wool in the top of the test-tube to exclude dust.
4. Stand the test-tube in an empty beaker and leave in a still place to cool.
5. Remove the cotton wool, put a thermometer in the melt, and record the temperature.
6. Stir with the thermometer and observe the temperature change at regular intervals as it solidifies.

Safety

Wear eye protection.

Questions

1. When a liquid turns into a solid is the process exothermic or endothermic?
2. When all the liquid has turned into solid the temperature begins to drop. Why is this?

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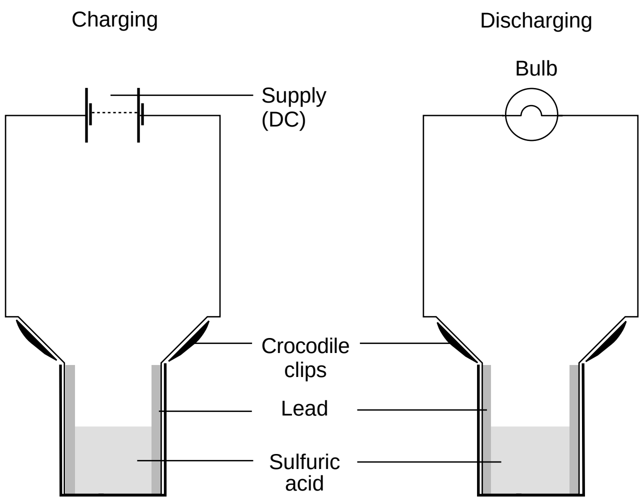
Accumulator

Introduction

Some types of cell are rechargeable. These cells store electricity. The most common rechargeable cell is the lead-acid type, which is the basis of car batteries. This experiment illustrates the charging and discharging of a lead-acid cell.



What to record

Complete the table:

Charging time /s	Time bulb is lit /s
180	
210	
240	
270	
300	

What to do

1. Connect the apparatus as shown.
2. Charge the cell at 4.5 V for three minutes.
3. Connect the cell for discharge.
4. Time how long the cell keeps the bulb lit.
5. Recharge the cell for a longer time and see how long the bulb stays lit.
6. Wash hands after handling lead.

Safety

Wear eye protection. Care with sulfuric acid.

Questions

1. Draw a line graph of your results. Charging time along the horizontal (*x*) axis and time lit along the vertical (*y*) axis.

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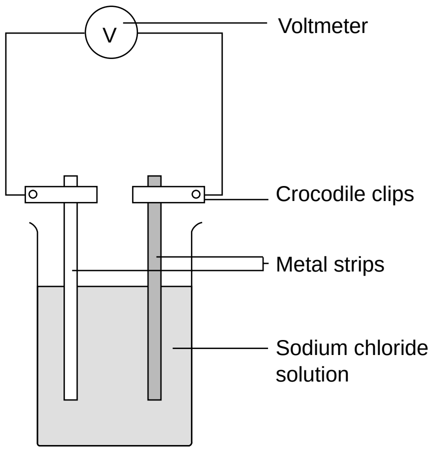
Electricity from chemicals

Introduction

Reactive metals form ions more readily than less reactive metals. This experiment illustrates the tendency of various metals to form ions. Two different metals and an electrolyte form a cell. The more reactive metal becomes the negative pole from which electrons flow.



What to record

Complete the table.

What to do

1. Set up the apparatus as shown.
2. Record the voltage.
3. Try all the combinations of metals.
4. Wash hands after handling lead.

Safety

Wear eye protection.

Metals used	**Which metal forms the positive terminal (+ve)**	**Which metal forms the negative terminal (-ve)**	**Voltage (V)**
Zinc and copper			
Copper and lead			
Lead and iron			
Zinc and lead			
Iron and magnesium			
Zinc and iron			
Zinc and magnesium			
Lead and magnesium			
Copper and magnesium			
Copper and iron			

Questions

1. Place zinc, magnesium, copper, lead, and iron in order of reactivity.

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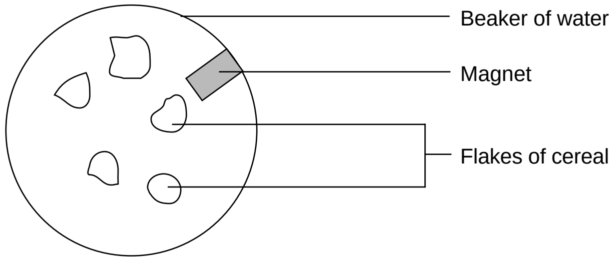
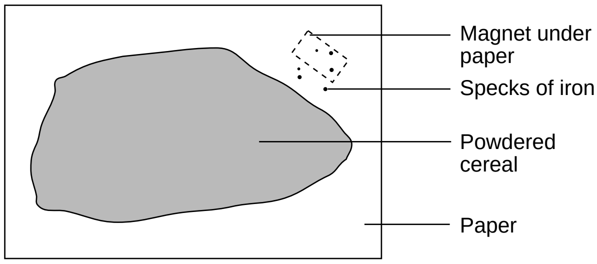
Iron in breakfast cereal

Introduction

Many breakfast cereals are fortified with iron. This iron is metallic and is added to the cereal as tiny particles of food grade iron before packaging. This experiment involves extracting the iron.





What to do

1. Float four to six pieces of cereal on the surface of a beaker of water.
2. Hold a magnet close to the cereal and see if this can cause a piece to move.
3. Put some cereal into a mortar and use a pestle to produce a very fine powder.
4. Spread the powder on a piece of paper.
5. Put a magnet under the paper and move the paper over the magnet.
6. Observe closely in the region of the magnet as the cereal moves over it.

Questions

1. Are all metals attracted to a magnet?
2. What are the symptoms of iron deficiency in the diet?

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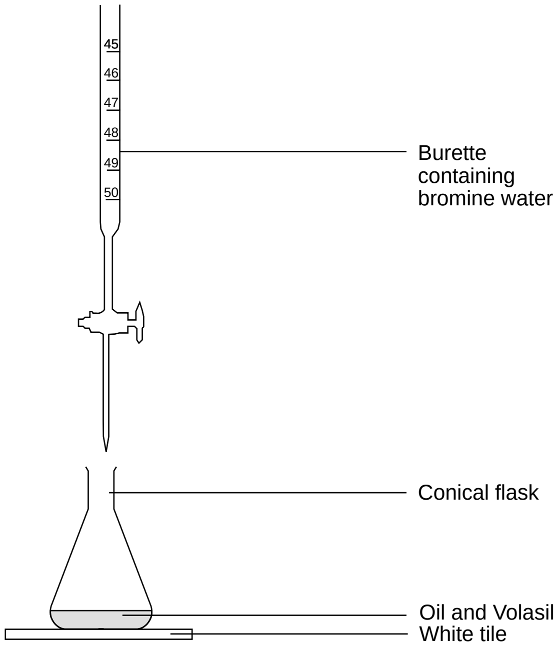
Unsaturation in fats and oils

Introduction

Advertisements often refer to unsaturated fats and oils. This experiment gives a comparison of unsaturation in various oils.



What to record

Volume of bromine water required for each oil.

What to do

1. Using a teat pipette, add five drops of olive oil to 5 cm^3 of Volasil in a conical flask.
2. Use a burette filled with a dilute solution of bromine water (0.02 mol dm^{-3}) **(Harmful and irritant**). Read the burette.
3. Run the bromine water slowly into the oil solution. Shake vigorously after each addition. The yellow colour of bromine disappears as bromine reacts with the oil. Continue adding bromine water to produce a permanent yellow colour.
4. Read the burette. Subtract to find the volume of bromine water needed in the titration.
5. Repeat the experiment with: five drops of cooking oil (vegetable) and five drops of cooking oil (animal).

Safety

Wear eye protection.

Questions

1. Which sample is the most saturated and which is the most unsaturated?
2. This comparison is only approximate. How could the method be improved?
3. What does unsaturated mean?

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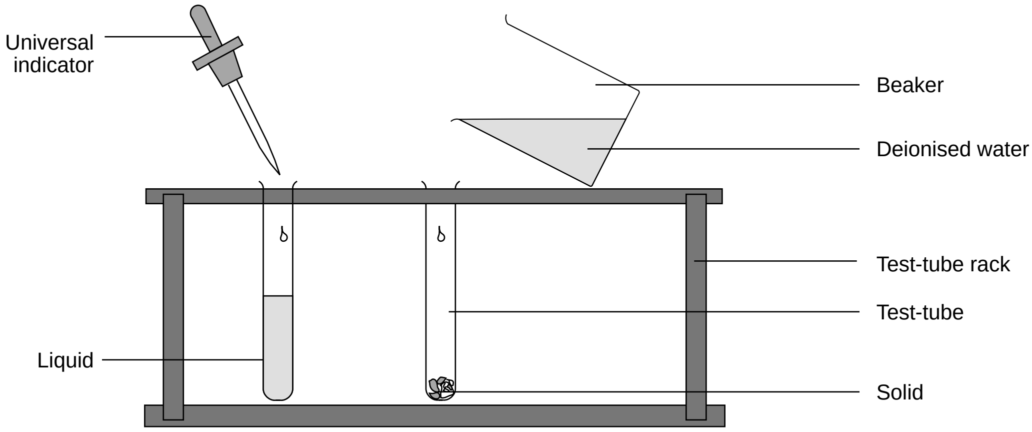
The pH scale

Introduction

The pH of a substance can be found by dissolving a small amount of the substance in deionised water and adding a few drops of Universal Indicator solution. The colour produced is compared with a pH chart.



What to record

Prepare a table for your results

Solution	Colour with Universal Indicator	pH

What to do

1. Place one spatula measure of solid, or pour a few drops of liquid into a test-tube.
2. Half-fill the test-tube with deionised water from a small beaker, and shake to dissolve the solid or mix the liquid.
3. Add a few drops of Universal Indicator to the test-tube. Make a note of the colour in the table. Compare it against the pH colour chart and record the pH of the nearest colour in the table.

Safety

Wear eye protection.

Questions

1. List the substances that were acidic, substances that were alkaline and substances that were neutral.
2. Why might a scientist prefer to use Universal Indicator rather than a different indicator like litmus?
3. What would happen if equal amounts of vinegar and limewater were mixed?

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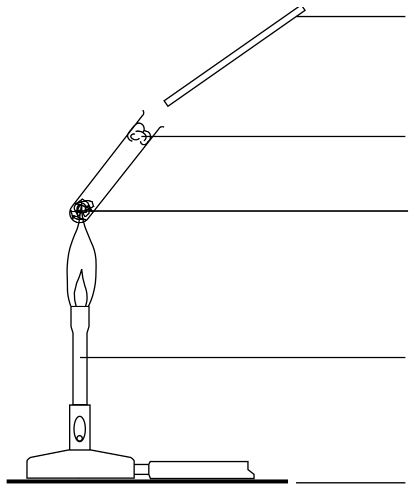
The preparation and properties of oxygen

Introduction

Potassium manganate(VII) produces oxygen when heated. In this experiment oxygen is produced and identified with a glowing splint.



What to record

What was done and what was observed.

What to do

1. Place two spatula measures of potassium manganate(VII) in a test-tube.
2. Place a small piece of ceramic wool near the top of the test-tube. This stops fine dust escaping.
3. Gently heat the test-tube containing the potassium manganate(VII).
4. Light a splint and extinguish it, to make a 'glowing splint'.
5. Place the glowing splint just above the top of the test-tube. Keep heating the test-tube. The splint should relight.
6. Scrape out the ceramic wool. Let the test-tube cool to room temperature and then wash it out.
7. Notice the colours produced when the test tube is washed out.

Safety

Wear eye protection.

Potassium manganate(VII) is harmful if swallowed. It assists fire.

Questions

1. What is the chemical formula for potassium manganate(VII)?

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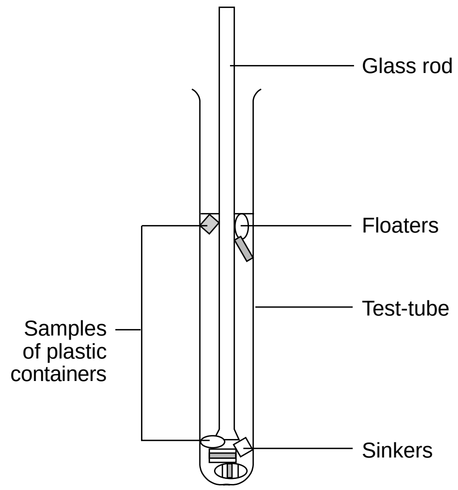
Identifying polymers

Introduction

In this experiment solutions with known densities are used to identify the polymers used in everyday materials.



What to record

Sample	Colour (or shape)	Solutions 1	2	3	4	5	6

What to do

1. Fill six test-tubes with solutions 1 to 6 and label each tube.
2. Place a sample of each type of polymer into solution 1.
3. Use a glass rod to stir the contents of the tube. Observe whether the waste plastics float or sink.
4. For samples that sink, write the letter S in column 1 of the results table.
5. Wash the glass rod and dry it on a tissue or paper towel.
6. Repeat the test for solutions 2 to 6. Use a new sample each time.

Safety

Wear eye protection.

Solutions 1,2 and 3 are highly flammable and toxic. Solutions 5 and 6 are irritants.

Questions

1. Why were the solutions stirred once the plastics were added?
2. Use the following table to identify the plastics. Fill in the table.

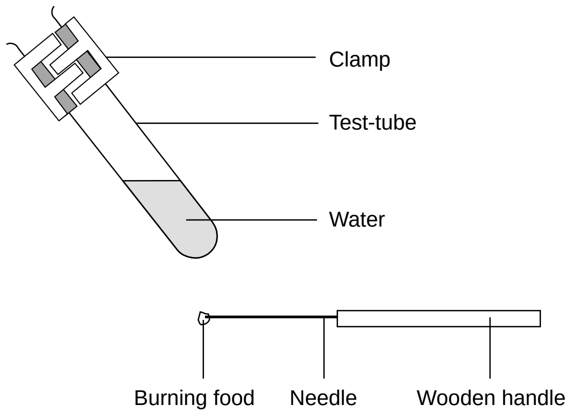
Polymer name	Acronym	Colour	Fingerprint					
			1	2	3	4	5	6
Polyethylene terephthalate	PET	Any	S	S	S	S	S	S
Polyvinyl chloride	PVC	Any	S	S	S	S	S	–
Polystyrene	PS	Any	S	S	S	S	–	–
High density polyethylene	HDPE	Any	S	S	S	-	–	–
Low density polyethylene	LDPE	Any	S	S	–	-	–	–
Polypropylene	PP	Any	S	–	–	–	–	–
Expanded polystyrene	EPS	Any	–	–	–	–	–	–

Polymer name	Acronym	Sample
Polyethylene terephthalate	PET	
Polyvinyl chloride	PVC	
Polystyrene	PS	
High density polyethylene	HDPE	
Low density polyethylene	LDPE	
Polypropylene	PP	
Expanded polystyrene	EPS	

Energy values of food

Introduction

In this experiment various foods are tested to find how much energy they contain.



What to record

Measurement	**Food**		
Mass/g			
Temperature of water before heating/°C			
Temperature of water after heating/°C			
Change in temperature/°C			
Heat absorbed by water/J (Temperature change x 4.2)			
Heat absorbed by water per gram of food/J			

What to do

1. Put 10 cm^3 of water in a test-tube. Clamp the test-tube in the retort stand at an angle as shown in the diagram.
2. Weigh a small piece of food and record the mass in your table.
3. Take the temperature of the water in the test-tube and record it in the table.
4. Fix the food on the end of the mounted needle. If the food is likely to melt when heated put it on a teaspoon instead of on the needle.
5. Light the food using a bunsen burner. As soon as the food is alight, hold it about 1 cm below the test-tube. If the flame goes out, quickly relight it.
6. When the food stops burning, stir the water in the test-tube with the thermometer and note the temperature. Record it in your table.
7. Empty the test-tube and refill it with another 10 cm^3 of water. Repeat the experiment using a different food each time.

Safety

Wear eye protection.

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Questions

1. Suggest reasons why this experiment may not be a fair test?
2. Burning gives out heat. What is the name given to this sort of reaction?
3. The label on a packet of cheese says 100 g provides 1638 kJ. Calculate how many joules this is per gram of cheese and compare it to the cheese in your experiment. (1 kJ = 1000 J)

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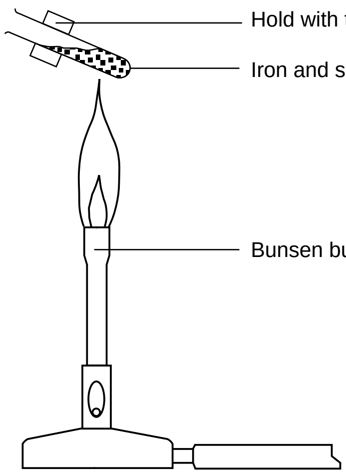
A compound from two elements

Introduction

A mixture of iron and sulfur can easily be separated. This is because there are no chemical bonds between the sulfur and the iron. The iron is magnetic and is therefore easily removed from the sulfur. In this experiment, a mixture of iron and sulfur are heated to make a new compound.



What to do

1. Examine the plastic bag of sulfur, the bag of iron and the bag containing a mixture of the two.
2. Run a magnet over each of the bags.
3. Set up the apparatus as shown in the diagram.
4. Light a Bunsen burner and half open the air-hole to give a medium flame.
5. Heat the very end of the tube strongly. When the mixture starts to glow, move the Bunsen burner to one side.
6. Watch the mixture in the tube. (If the glow just goes out, heat the tube again.)
7. Let the tube cool down completely.
8. The substance from the tube is a new compound called iron sulfide.
9. Test the iron sulfide with a magnet. Does the magnet pick it up?

Safety

Wear eye protection. Do not get too close to the fumes.

Questions

1. Write a word equation for this reaction.
2. What has happened to the iron and the sulfur in forming iron sulfide?
3. What is the chemical formula for iron sulfide?

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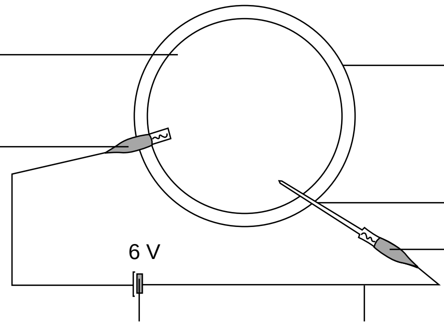
Chemistry and electricity

Introduction

In this experiment, electricity and some indicators are used to make coloured writing.



What to do

1. Dissolve a spatula measure of sodium chloride in 2 cm^3 of water. Add three drops of methyl orange indicator.
2. Lay a filter paper inside a plastic petri dish. Drop the solution onto the paper using a dropping pipette, until the paper holds no more solution.
3. Attach the positive end of a 6 V battery to a lead ending in a crocodile clip. Use the crocodile clip to grip one end of the paper.
4. Attach the negative end of the battery to a carbon electrode.
5. Write lightly on the wet paper, using the carbon electrode. What colour is the writing?
6. Repeat the experiment using Universal Indicator. Describe what happens.

Safety

Wear eye protection.

Questions

1. What would happen if the lead were attached to the positive electrode using Universal Indicator? Try this if there is time.
2. Explain what reactions have occurred to produce the colours.

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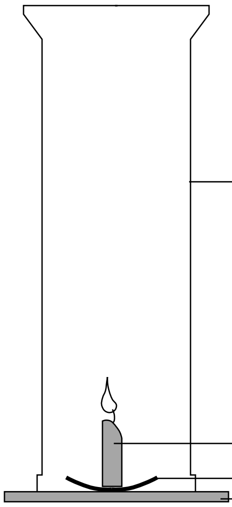
Combustion

Introduction

Hydrocarbons produce carbon dioxide and water when they burn. In this experiment the products of combustion are captured and tested.



What to record

What was done and what was observed.

What to do

1. Set up the apparatus as shown in the diagram. The gas jar should be placed over the lit candle on a heatproof mat.
2. When the candle goes out, put a lid on the gas jar.
3. Test to see if the candle made water by adding a piece of blue cobalt chloride paper, test the sides of the jar. If it turns pink, water is present.
4. Now test to see if carbon dioxide was produced. Pour a little limewater into the gas jar. Swill it around a little. If carbon dioxide is present, the limewater turns cloudy.

Safety

Wear eye protection.

Questions

1. What is the gas that reacts with the hydrocarbon when it burns?
2. What gases does the candle produce when it burns?
3. Name another fuel that produces the same gases when it burns.

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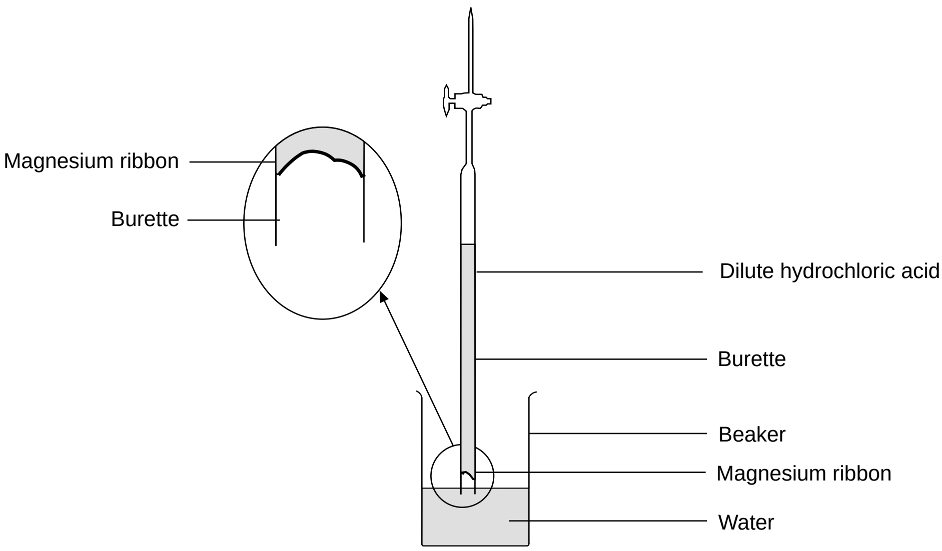
The determination of relative atomic mass

Introduction

One mole of any gas occupies the same volume when measured under the same conditions of temperature and pressure. In this experiment, the number of moles of hydrogen produced from a known mass of magnesium is measured. The relative atomic mass of magnesium can therefore be calculated.



What to record

The mass of magnesium used and the volume of hydrogen produced.

What to do

1. Clean a piece of magnesium ribbon (about 3.5 cm long) and weigh it accurately (This should weigh between 0.02 g and 0.04 g; if not adjust the amount used.)
2. Measure 25 cm^3 of dilute hydrochloric acid into the burette. Carefully add 25 cm^3 of water on top of this.
3. Push the magnesium in the end of the burette so it stays in position with its own tension.
4. Add 50 cm^3 of water to a 250 cm^3 beaker.
5. Quickly invert the burette into the water, (if this is done quickly and carefully very little will be lost. It is important that the liquid level in the burette starts on the graduated scale. If it is not on the scale; momentarily open the tap, this allows the level to drop).
6. Clamp the burette vertically.
7. Take a burette reading (**NB:** it is upside down!)
8. Observe how the magnesium reacts as the acid diffuses downwards, wait until all the magnesium has reacted.
9. Note the new volume on the burette (**NB:** it is upside down).
10. Record your results.

Safety

Wear eye protection.

Questions

The equation for the reaction is

$Mg + 2HCl \rightarrow MgCl_2 + H_2$

1. Copy out and fill in the gaps:

____ g magnesium was produced from _____cm^3 of hydrogen.

____ g magnesium was produced from 1 cm^3 of hydrogen

____ g magnesium was produced from 24000 cm^3 of hydrogen.

____ g magnesium would be produced from 1 mole of hydrogen.

This is the mass of 1 mole of magnesium. Numerically, this number is the relative atomic mass of magnesium.

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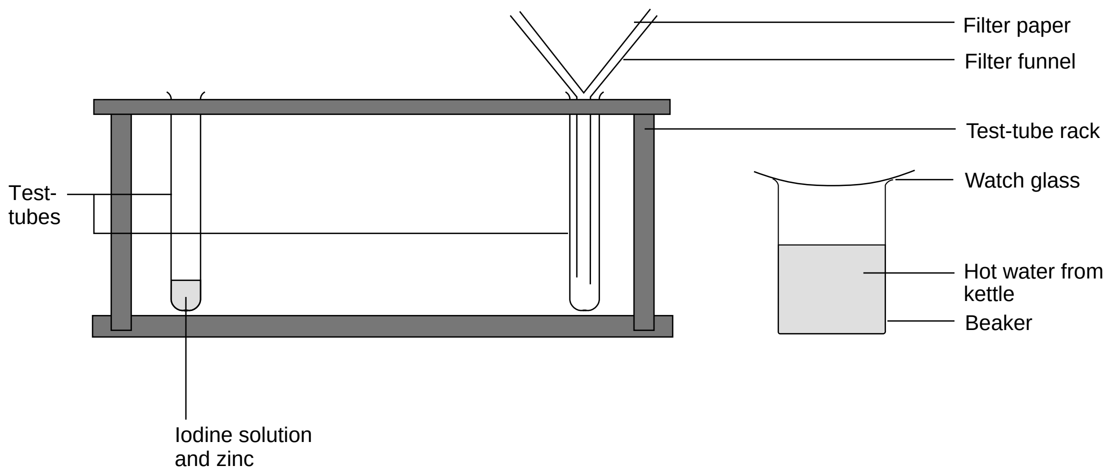
The reaction of a Group 7 element (iodine with zinc)

Introduction

This experiment involves producing a salt by reacting a Group 7 element (iodine) with zinc. This is an example of salt preparation by direct synthesis.



What to do

1. Use the measuring cylinder to add 4 cm^3 of alcohol to a test-tube.
2. Dissolve a few crystals of iodine in the alcohol.
3. Note the temperature.
4. Add some powdered zinc (about the volume of a grain of rice) to the iodine solution.
5. Stir with the thermometer and note the highest temperature.
6. Record any changes which suggest a chemical reaction has taken place.
7. Filter the liquid into another test-tube (the excess zinc is left behind).
8. Collect a beaker of hot water, and place a watch glass on the top
9. Pour the liquid filtrate onto the watch glass and allow the alcohol to evaporate. Observe the product.
10. The product is classified as a SALT.

Safety

Wear eye protection. Do not touch the iodine.

Questions

1. What is the collective name for the Group 7 elements?
2. What does this name mean?
3. Complete the word equation:
   Zinc + iodine →

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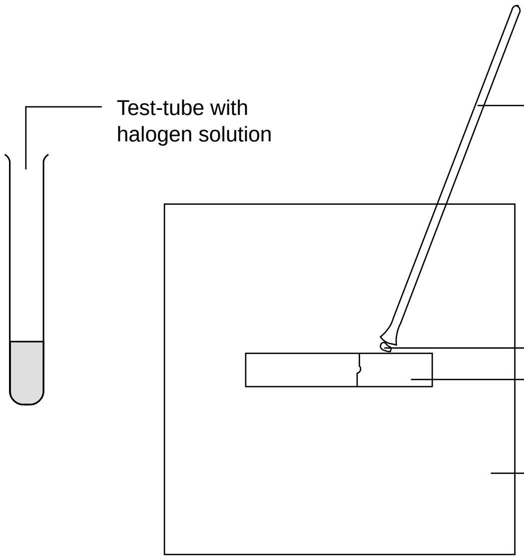
Reactions of halogens

Introduction

The Group 7 elements are called the halogens. This experiment involves some reactions of the halogens.



What to record

Complete the table.

	Effect on indicator paper	**Reaction with potassium chloride solution**	**Reaction with potassium bromide solution**	**Reaction with potassium iodide solution**
Chlorine water				
Bromine water				
Iodine water				

What to do

1. Put a piece of Universal Indicator paper onto a white tile.
2. Use a glass stirring rod to transfer a few drops of the first solution onto the indicator paper.
3. Repeat this with a fresh piece of paper and the different solutions.
4. In a test-tube, add some chlorine solution to a solution of potassium bromide.
5. Add some chlorine solution to a solution of potassium iodide.
6. Now try mixing solutions of bromine and potassium iodide. If there is time, mix the other combinations of solutions to complete the table.

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Safety

Wear eye protection. Do not breathe chlorine gas.

Questions

1. Which halogen solution is the strongest bleaching agent?
2. Which halogen is the most reactive?
3. Write a word and symbol equation for the reaction of chlorine with potassium bromide.

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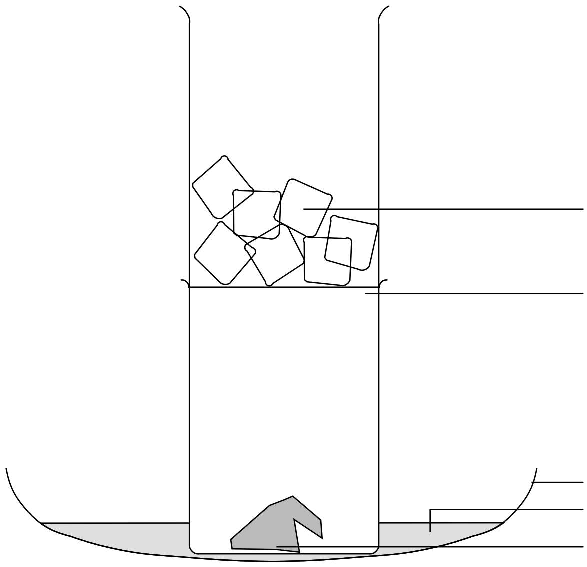
The sublimation of air freshener

Introduction

Sublimation is an interesting physical change. When a substance sublimes, it changes directly from a solid to a gas without passing through the liquid state. Dry ice sublimes, as do iodine and mothballs. This experiment involves the study of another common substance that sublimes – air freshener.



What to do

1. Place a few lumps of air freshener in the bottom of the 100 cm^3 beaker.
2. Put the 100 cm^3 beaker carefully on top of the other 100 cm^3 beaker.
3. Fill the top beaker three quarters full with ice. Ensure no ice enters the beaker below.
4. Fill the shallow dish or pan about one-third full of hot water (at a higher temperature than 45 °C).
5. Place the sublimation apparatus in the shallow dish in a fume cupboard.
6. Observe what happens to the solid. Be patient, it may take a while.

Safety

Wear eye protection. Use a fume cupboard.

Questions

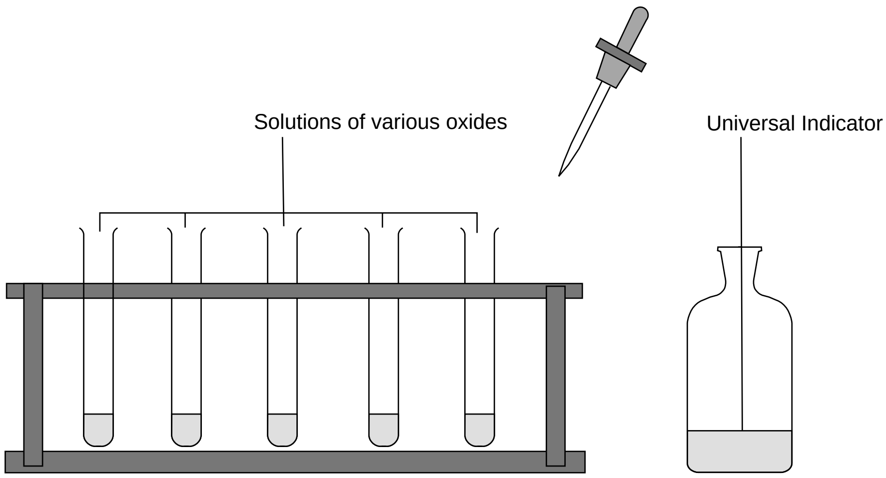
1. What might be the significance of 45 °C? Try lower and higher temperatures if there is time.
2. Define ‘sublimation’.
3. Use the particle theory of matter to explain what is happening and include a particle diagram.

RS•C

Testing the pH of oxides

Introduction

In this experiment the pH of various oxides is tested.



What to record

Name of oxide	Colour of Universal Indicator	pH value	Acid, alkali or neutral
Nitrogen oxide			
Sodium oxide			
Potassium oxide			
Phosphorus(V) oxide			
Calcium oxide			

What to do

1. Using separate test-tubes, collect a sample (about 2 cm^3) of each oxide in water.
2. Add three drops of Universal Indicator solution to each sample.
3. Record the results in a table showing the oxide, the colour of the Universal Indicator, the pH and whether the oxide is acidic, alkaline or neutral in water.

Safety

Wear eye protection

Questions

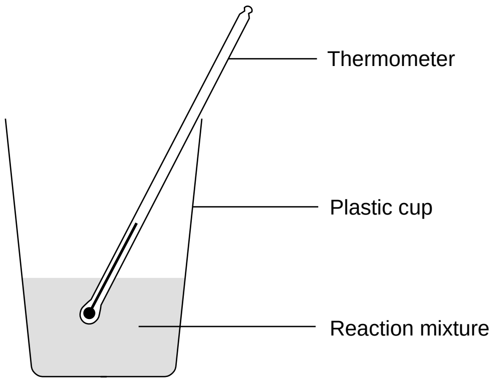
1. Which compounds in the table are metal oxides?
2. Comment on your results for the metal oxides?
3. Some metal oxides do not react with water. Predict the pH of these compounds.

RS•C

Exothermic or endothermic?

Introduction

Some reactions give out heat and others take in heat. In exothermic reactions the temperature goes up, in endothermic reactions the temperature goes down. In this experiment, various reactions are examined. Temperatures are measured to decide whether a particular reaction is exothermic or endothermic.



What to record

Complete the table

Reaction	Temperature before mixing/°C	Temperature. after mixing/°C	Exothermic or endothermic
Sodium hydroxide solution + dilute hydrochloric acid			
Sodium hydrogen carbonate solution + citric acid			
copper(II) sulfate solution + magnesium powder			
Dilute sulfuric acid + magnesium ribbon			

What to do

1. Use the apparatus as shown.
2. Put 10 cm^3 of sodium hydroxide solution in the beaker, record the temperature then add 10 cm^3 of dilute hydrochloric acid, stirring with the thermometer. Record the maximum or minimum temperature.
3. Repeat the procedure for the following reactions: (*a*) sodium hydrogen carbonate solution and citric acid; (*b)* copper(II) sulfate solution and magnesium powder; and (*c*) dilute sulfuric acid and magnesium ribbon.

Safety

Wear eye protection. Some of the solutions are irritant.

RS•C

Questions

1. The first reaction is between an acid and an alkali, what do we call this type of reaction?
2. Which gas is given off when sodium hydrogen carbonate reacts with citric acid?
3. Which type of reaction takes place between copper(II) sulfate and magnesium?
4. Which reactions are exothermic and which are endothermic?
5. Describe in terms of bond breaking and bond making, why some reactions are exothermic and some are endothermic.

RS•C

Water expands when it freezes

Introduction

Water expands when it freezes. Most liquids contract when they freeze so this property of water is unusual. This property is clearly shown in this experiment. This process is used to explain how ice can break rocks apart.

What to record

What happens.

What to do

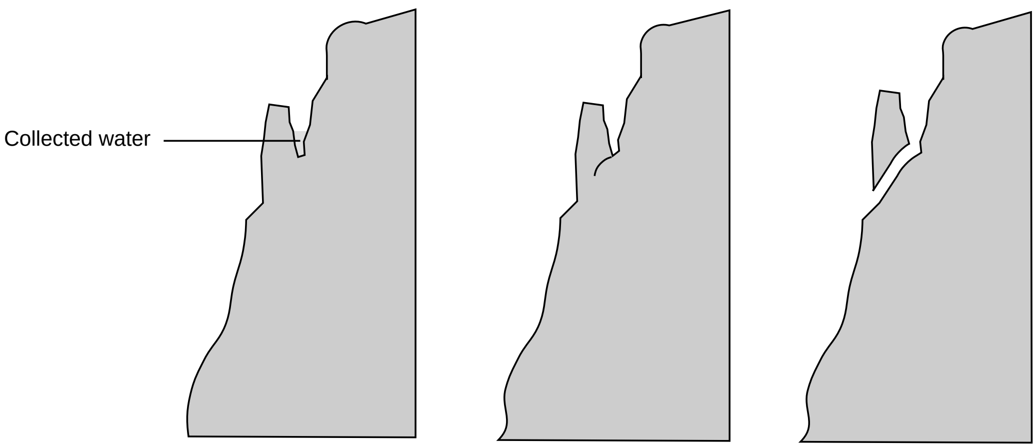
1. Fill a screw top bottle right up to the top with water.
2. Screw on the top tightly.
3. Tie the sealed bottle in a clear plastic bag.
4. Leave overnight in a freezer.

Safety

Care when removing frozen bottle from freezer.

Questions

1. Use what happened to explain why water pipes sometimes burst in winter.
2. What happens when your milk freezes on the doorstep in winter?
3. Use this knowledge to add captions to the diagram that explain how rocks crack.



RS•C

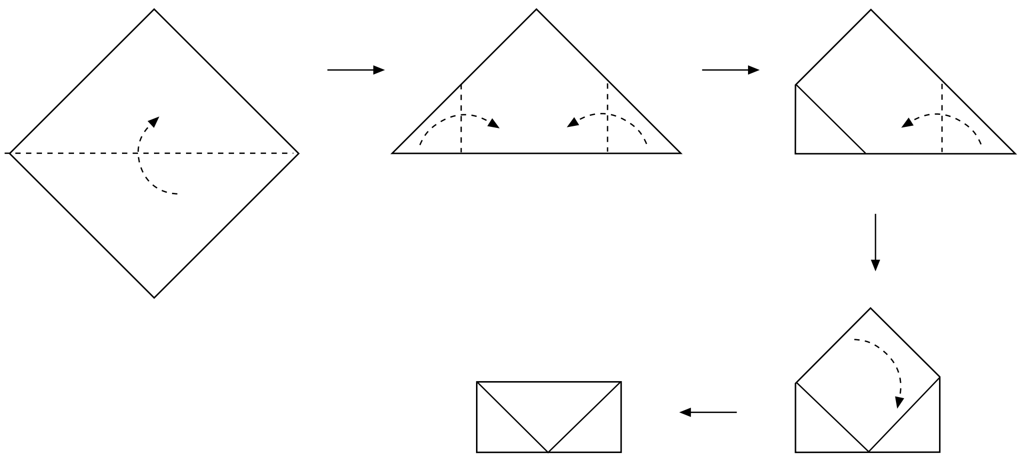
The chemical properties of the transition metals – the copper envelope

Introduction

Transition metals are situated between Groups 2 and 3 of the Periodic Table. They have important uses. One well-known transition metal is copper. Transition metals have similar reactions and properties.

Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd
La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg

This is the sequence of folds to produce the copper envelope.



What to do

1. Fold a square of copper foil into an envelope as shown in the diagram. Care – sharp corners.
2. Hold the envelope in tongs.
3. Heat strongly in a Bunsen burner flame for 5 min.
4. Place the copper envelope on the heatproof mat to cool. Care – hot.
5. When the envelope is cool enough to hold; open it up and compare the inside with the outside.

Safety

Wear eye protection.

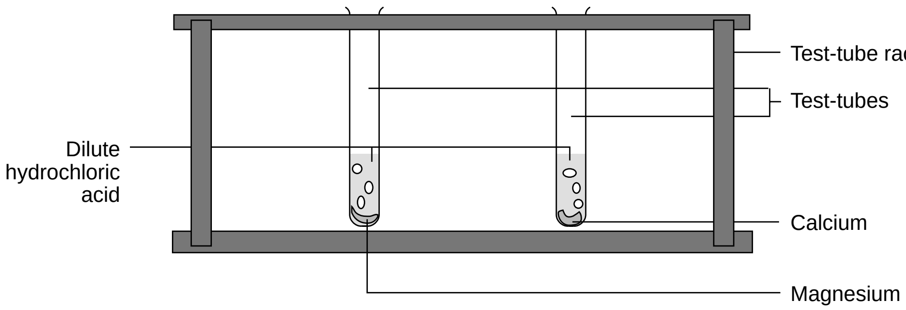
Questions

1. Does the copper burst into flames like magnesium?
2. What does the copper look like after it has cooled?
3. Which gas in the air has copper reacted with?
4. What is the black coating on the surface called?
5. What is the appearance of the inside of the envelope and why is this?

The reactivity of Group 2 metals

Introduction

Metals in Group 2 of the Periodic Table are less reactive than those in Group 1. This experiment indicates the relative reactivity of elements within the group.



What to do

1. Fill two test-tubes a quarter full with dilute hydrochloric acid.
2. Into one test-tube drop a small piece of magnesium.
3. Into the other, drop a small piece of calcium.
4. Compare the reactivity of the two metals.
5. Drop another bit of magnesium into the first test-tube and put your thumb over the end.
6. When the pressure can be felt, take your thumb off and test the gas with a lighted splint.
7. Record what happens.

Safety

Wear eye protection.

Questions

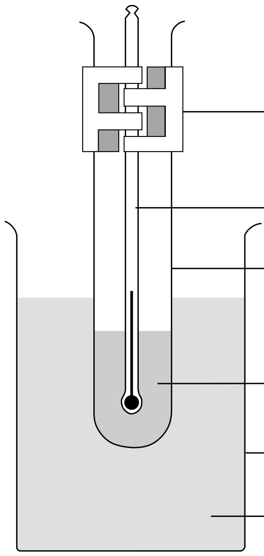
1. Which is the more reactive, magnesium or calcium?
2. Write word equations for these reactions.
3. Write formula equations for these reactions.

RS•C

Melting and freezing

Introduction

In this experiment, a solid turns into a liquid and then the liquid turns into a solid. The energy changes are examined.



What to record

Complete the table.

Time/min	Temperature/°C

What to do

1. Fill a 250 cm^3 beaker with about 150 cm^3 tap water.
2. Heat it on a tripod and gauze until the water just starts to boil.
3. Set up the apparatus as shown in the diagram and start the timer.
4. Try and maintain the temperature of the water. It should be just boiling but not boiling vigorously.
5. Record the temperature every minute as the stearic acid heats up, until it reaches about 70 °C. Show in your table the temperature where the solid starts to melt.
6. Use the clamp stand to lift the tube from the hot water. Record the temperature every minute as the stearic acid cools down until it reaches about 50 °C. Note the temperature in your table when the first signs of solid formation are observed.
7. Plot a line graph of your results. Put time along the bottom and temperature up the side. Label your graph to show where stearic acid is a solid, a liquid or present in both states.

Safety

Wear eye protection.

Questions

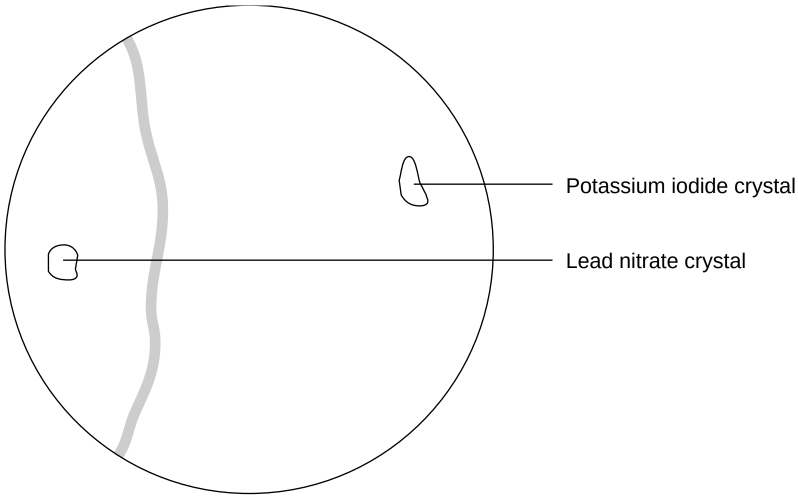
1. What is the melting point of stearic acid?
2. What is the freezing point of stearic acid?
3. Why are there flat sections on your graph? Explain this in terms of the forces between particles.

RS•C

Diffusion in liquids

Introduction

Diffusion occurs in liquids but more slowly than in gases. The particles are not as free to move about. This experiment illustrates diffusion in a liquid.



What to do

1. Fill a petri dish with deionised water.
2. Use forceps to drop crystals of lead nitrate and potassium iodide on opposite sides of the dish.
3. The solid crystals form solutions that react. Observe what happens.
4. Watch diffusion occurring, as a yellow solid slowly forms between the two crystals.

Safety

Wear eye protection.

Questions

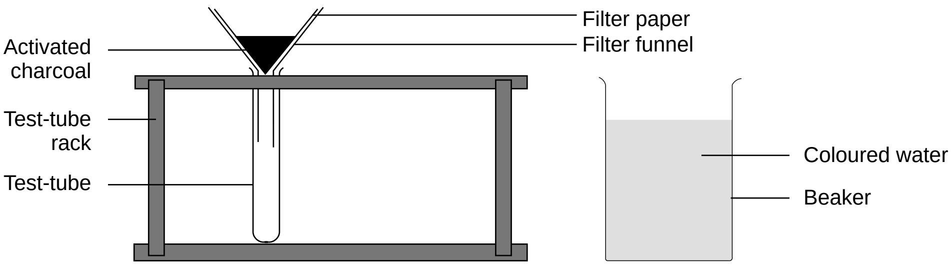
1. Write a word equation for the reaction.
2. Write a formula equation for the reaction.
3. How can this diffusion process be explained?
4. With crystals of lead nitrate and potassium iodide in water, the yellow solid forms closer to the lead nitrate crystal. Which reactant diffuses faster, the lead ion or the iodide ion? Why?

RS•C

Chemical filtration

Introduction

In this experiment, carbon that has undergone special treatment to make it into decolourising carbon is shown to remove colour and odour from various solutions. This form of carbon is sometimes called activated charcoal. This method is used to remove objectionable taste and odours from drinking water.



What to do

1. Fold a piece of filter paper, place it in a funnel, and put the stem of the funnel into a test-tube in a test-tube rack.
2. Add about five spatulas of decolourising carbon to the funnel
3. Add one drop of ink or food colour to 100 cm^3 of water in a beaker.
4. Carefully pour some of the coloured water onto the charcoal in the filter paper.
5. Prepare another filter paper with the same amount of carbon. This time filter a solution made by adding two or three crystals of potassium manganate(VII) to 100 cm^3 of water.
6. Repeat the activity, this time filter sauerkraut juice, dill pickle juice or vinegar.

Safety

Wear eye protection.

Questions

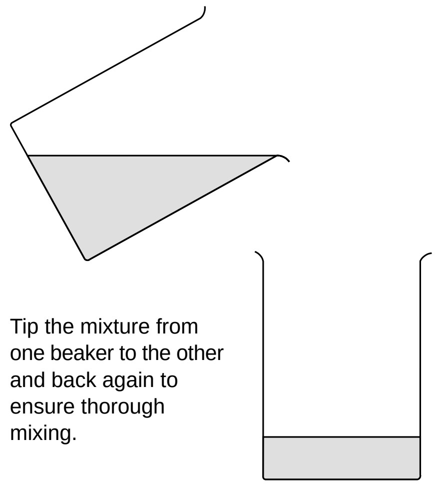
1. Describe the material before and after filtration in each of the three activities.
2. How does carbon remove colour and odour?
3. How could this process be used to provide pure water for drinking?

RS•C

Rate of reaction – the effect of concentration and temperature

Introduction

In this experiment, two colourless solutions are mixed to make a solution which becomes dark blue. Changing the concentration or temperature of the solutions changes the time required for the blue colour to develop.



What to record

The conditions and the times for reactions to occur.

What to do

1. Place 50 cm^3 of solution A in a 250 cm^3 beaker.
2. Place the same volume of solution B in a second beaker.
3. Mix the two solutions by pouring from one beaker into the other several times.
4. Note the time required for a reaction to occur (formation of blue colour).
5. Repeat, but use solution A that has been diluted to one half the concentration. Note the time for the reaction to occur.
6. Repeat using solution A warmed to 35 °C. Note the time for a reaction to occur.

Safety

Wear eye protection

Questions

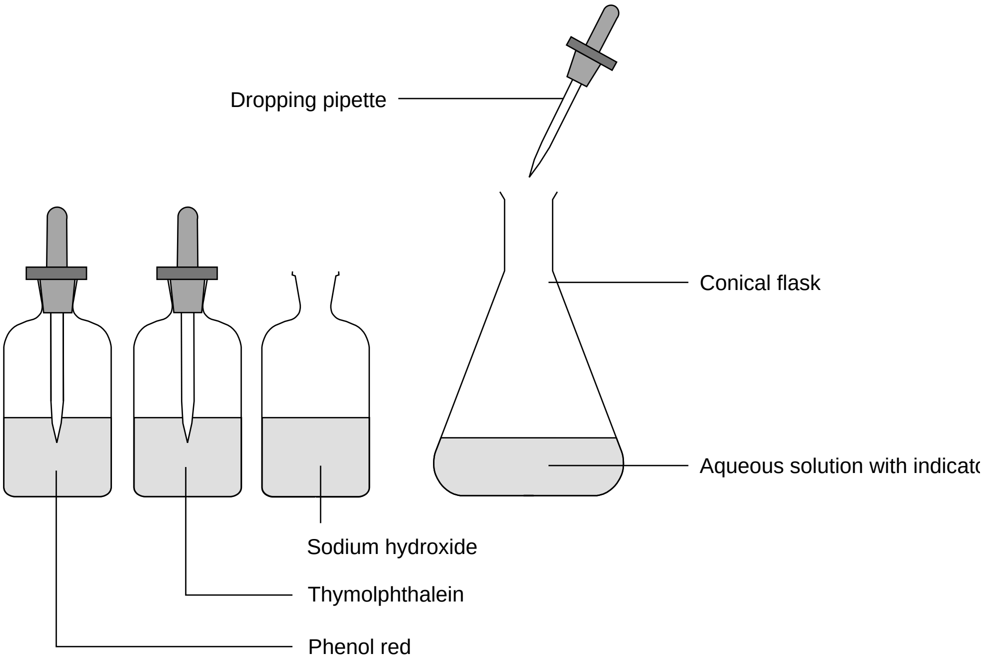
1. Why does increasing the concentration usually result in an increased rate of reaction?
2. Why does increasing the temperature usually result in an increased rate of reaction?
3. How could this experiment be set up so it took exactly 10 min to turn blue?

RS•C

Reaction between carbon dioxide and water

Introduction

When carbon dioxide reacts with water a weak acid is formed. Carbon dioxide is present in exhaled breath. Observing a colour change using an acid-base indicator shows the reaction between carbon dioxide and water.



What to do

Activity 1

1. Place about 125 cm^3 of ethanol (**Highly flammable**) in a 250 cm^3 conical flask.
2. Add five or six drops of thymolphthalein indicator to the alcohol.
3. Add just enough dilute sodium hydroxide (**Irritant**) (dropwise) to produce a blue colour.
4. Talk or blow gently into the flask – *ie* add the carbon dioxide.
5. Continue adding the carbon dioxide until a colour change is observed.

Activity 2

1. Place about 125 cm^3 of water in a 250 cm^3 conical flask.
2. Add one or two drops of phenol red to the water.
3. Add two drops of sodium hydroxide solution (**Irritant**) to produce a red solution.
4. Talk or blow gently into the flask – *ie* add carbon dioxide.
5. Continue adding the carbon dioxide until a colour change is observed.

Safety

Wear eye protection.

Questions

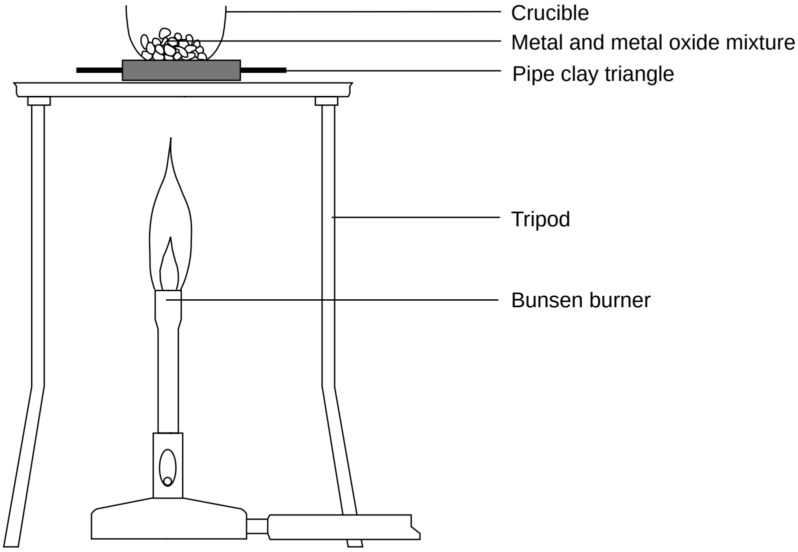
1. Why does the colour change not occur instantly?
2. Why are a few drops of sodium hydroxide solution (NaOH) added before the experiment?

RS•C

Competition for oxygen

Introduction

This experiment involves the reaction of a metal with the oxide of another metal. When reactions like these occur, the two metals compete for the oxygen. The more reactive metal finishes up with the oxygen (as a metal oxide). If the more reactive metal starts as the oxide then no reaction takes place.



What to record

Decide whether a reaction takes place in each case.

What to do

1. Set up the apparatus as shown in the diagram.
2. Place one spatula measure of one of the reaction mixtures into the crucible.
3. Heat the mixture gently at first and then more strongly. Watch carefully to see what happens but do not lean over the crucible.
4. Allow the mixture to cool. Look for evidence that a reaction has taken place.
5. Use your observations to decide which of the two metals has ‘won’ the competition for oxygen – which is more reactive?
6. Choose another mixture and repeat the experiment.

Safety

Wear eye protection. Do not lean over the crucible.

RS•C

Questions

1. Complete the following table.

Reaction mixture	Does this mixture react?
Magnesium oxide and iron	
Lead oxide and iron	
Lead oxide and zinc	
Iron oxide and zinc	

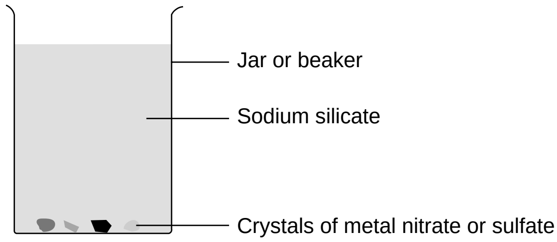
2. Write word equations for any reactions that occur.

RS•C

Making a crystal garden

Introduction

The formation of molten silicates in the Earth's mantle involves the formation of silicon dioxide and its subsequent reaction at high temperatures with metal oxides. In this experiment coloured silicates are formed in the laboratory.



What to record

What is observed.

What to do

1. Pour sodium silicate solution (**Irritant**) into a glass jar to a depth of 3 cm. Add hot water to this solution, stirring well during the addition. The final depth of liquid required is about 12 cm. Stirring should continue until no separate silicate layer is visible.
2. Allow the solution to stand until the liquid is quite still.
3. Use forceps to drop a few crystals into the liquid, try and choose different colours. Ensure that the crystals do not fall close to each other.
4. Cover the jar and leave overnight.

Safety

Wear eye protection.

Questions

1. Why are the silicate crystals different colours?
2. Why could silicates not be formed in the laboratory in exactly the same way as they are formed in the Earth's mantle?

RS•C

Extracting metal with charcoal

Introduction

This extraction experiment consists of two competition reactions. A metal oxide is reacted with charcoal. If the charcoal (carbon) is more reactive it will remove the oxygen from the metal oxide and leave a trace of metal in the reaction vessel. Start with an oxide of lead, then observe what happens to an oxide of copper.

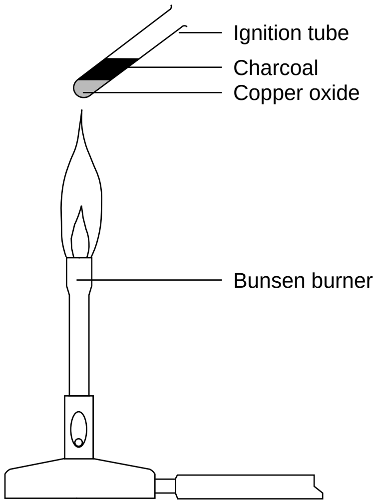
What to do

Lead oxide

1. Add one spatula of lead oxide (**Toxic**) to the ignition tube.
2. Add one spatula of charcoal powder.
3. Mix this mixture using an unfolded paper clip.
4. Strongly heat this mixture for five minutes in the Bunsen burner flame.
5. Allow to cool.
6. Tip the mixture onto a heatproof mat.

Copper oxide

1. Add one spatula of copper oxide (**Harmful**) to the ignition tube.
2. Carefully add one spatula of charcoal powder on top without any mixing.
3. Strongly heat these two layers for five minutes in the Bunsen burner flame.
4. Allow to cool and then look closely at the ignition tube where the powders meet.



Safety

Wear eye protection.

Questions

1. Why should the mixture be cool before it is tipped out of the ignition tube at the end?
2. What happens to the carbon if it takes oxygen from the metal oxide?
3. Write the equation for the reaction of lead oxide.
4. Which element is oxidised and which is reduced in this reaction?

RS•C

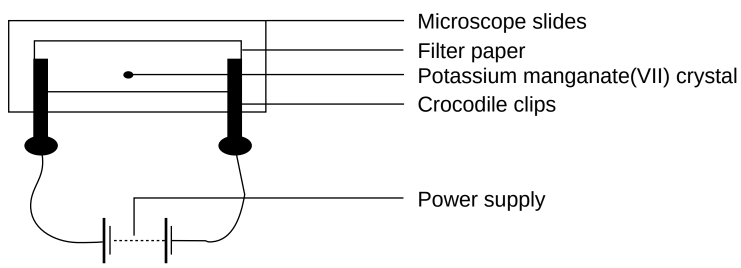
The migration of ions

Introduction

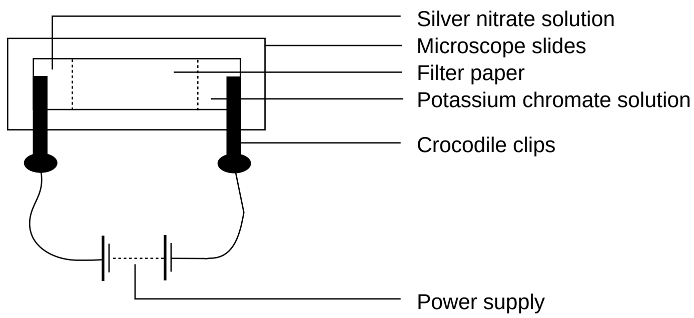
In an electrolysis experiment, the ions migrate towards electrodes of opposite charge. In the first part of this experiment the migration of manganate ions is observed.

In the second part of this experiment, silver and chromate ions meet as they migrate towards opposite electrodes. Silver chromate, an insoluble red compound is formed.

Part 1



Part 2



What to record

Draw a diagram to show the filter paper at the end of both experiments. Mark the ends of the paper with a + or – to show which terminal of the power supply each end is connected to.

What to do

Part 1

1. Cut a piece of filter paper slightly smaller than a microscope slide. Draw a faint pencil line across the middle.
2. Moisten the filter paper with tap water. Fasten the paper to the slide with crocodile clips.
3. Use forceps to put a small crystal of potassium manganate(VII) in the centre of the paper.
4. Connect the clips to a power supply set at not more than 20 V DC. Switch on and wait about ten minutes.

Part 2

1. Cut a piece of filter paper slightly smaller than a microscope slide. Draw a faint pencil line across the middle.
2. Moisten the paper with tap water. Fasten the paper to the slide with crocodile clips.

3. At the end of the paper where the positive electrode will be placed, moisten the paper with silver nitrate solution. (**Will stain fingers**)
4. At the end of the paper where the negative electrode will be placed, moisten the paper with potassium chromate solution. (**Toxic**).
5. Connect the clips to a power supply set at no more than 20 V DC. Switch on and wait about 10 min, or until you see a change.

Safety

Wear eye protection.

Questions

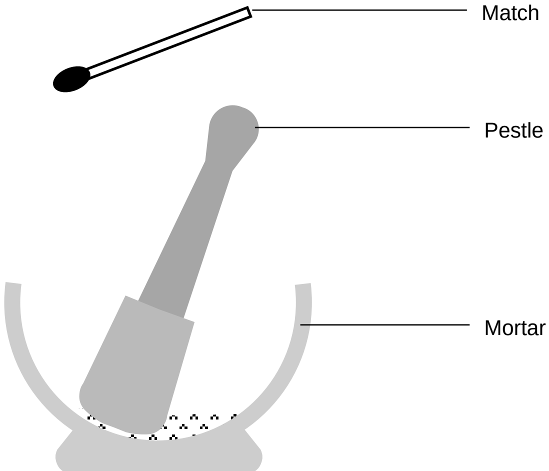
1. Potassium manganate (VII) consists of two ions - potassium ions and manganate (VII) ions. One of these ions is coloured. Which is it likely to be?
2. From the direction of movement, what does this indicate about the charge on the manganate ion?
3. What is happening in the second experiment?
4. The formula of silver nitrate is $AgNO_3$ and potassium chromate is K_2CrO_4. Write the ionic formula equation for the reaction.

RS•C

The reduction of iron oxide by carbon

Introduction

Metals high in the reactivity series will reduce the oxides of those lower in the series. The oxides of metals between zinc and copper in the reactivity series can be reduced by carbon. In this experiment, sodium carbonate is used to fuse the reactants in intimate contact.



What to do

1. Char the point of a used match, moisten it with a drop of water and rub on some sodium carbonate crystals.
2. Rub the point in some powdered iron(III) oxide (Fe_2O_3) and heat in a blue Bunsen burner flame until the point glows strongly.
3. Allow to cool.
4. Crush the charred head in a mortar and pestle then run a magnet through the pieces.

Safety

Wear eye protection.

Questions

1. What does 'reduction' mean?
2. Carbon does not reduce aluminium oxide. Where would carbon be placed in this reactivity series?
   Potassium
   Sodium
   Calcium
   Magnesium
   Aluminium
   Zinc
   Iron
   Lead
   Copper

What other information would you need to determine carbon's exact place?

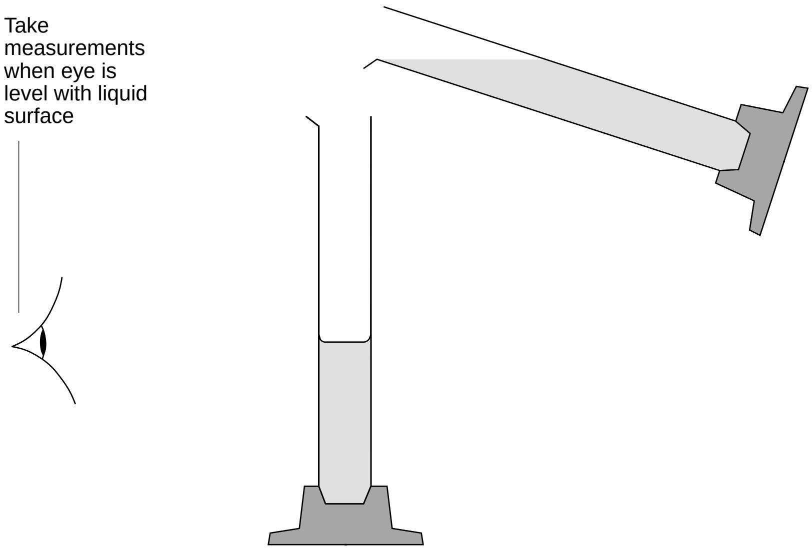
3. Explain why calcium oxide cannot be reduced using carbon.

RS•C

Experiments with particles

Introduction

When materials are added together, they may acquire new properties. When a solid and a liquid are mixed, the solid may or may not dissolve. When two liquids are mixed they may become one liquid or stay separate. These experiments provide an opportunity to predict and then observe what happens.



What to record

Activity 1

Volume of peas/cm^3	Volume of sand/cm^3	Combined volume/cm^3

Activity 2

Volume of alcohol/cm^3	Volume of water/cm^3	Combined volume /cm^3

Activity 3

Initial volume of water/cm^3	Final volume of salt solution/cm^3

RS•C

What to do

Activity 1

1. Add approximately 25 cm^3 of dried peas and 25 cm^3 of sand to separate measuring cylinders. Accurately measure and record the volumes.
2. The contents of one cylinder is added to the other and shaken until the two substances are mixed together.
3. Place the measuring cylinder on the bench and gently shake from side to side to allow the mixture to settle.
4. Read the combined volume.

Activity 2

1. Add approximately 25 cm^3 of ethanol (**Highly flammable**) and 25 cm^3 of water to separate measuring cylinders. Accurately measure and record the volumes.
2. The contents of one cylinder is added to the other and shaken from side to side for 15–30 seconds until the two substances are mixed together and then left to stand for one minute.
3. Read the combined volume.

Activity 3

1. To the measuring cylinder add approximately 75 cm^3 of water. Accurately measure and record the volume.
2. Spatulas of salt should then be added one at a time until the salt begins to be left at the bottom of the cylinder, despite continued stirring.
3. The volume reading on the side of the cylinder should again be recorded.

Safety

Wear eye protection.

Questions

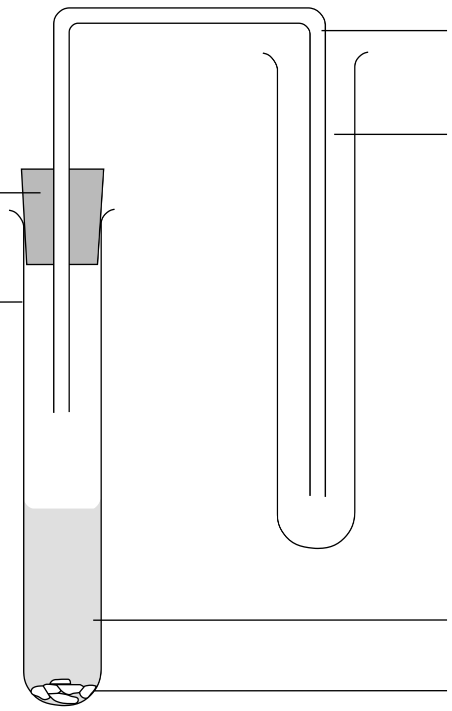
1. What is the similarity between the first two activities?
2. What is an explanation for the result in the last activity?

RS•C

Particles in motion?

Introduction

These two activities suggest that particles in a gas are in motion.



What to do

1. Set up the apparatus as shown in the diagram.
2. Put a spatula measure of calcium carbonate into the first test-tube.
3. Add 10 cm^3 of hydrochloric acid and quickly replace the bung and delivery tube. Ensure the delivery tube reaches almost to the bottom of the second test-tube.
4. Allow the gas to pass into the second test-tube for about one minute, then remove the delivery tube and cork the test-tube.
5. Hold the test-tube upside down over a similar test-tube containing air.
6. Remove the cork and place the tubes mouth to mouth.
7. After 5 min, cork both tubes and test the contents for carbon dioxide (swirl a little limewater round in the test-tube). Write down what happens in both tubes.
8. Repeat this experiment but this time at step 5 hold the test-tube of air upside down over a test-tube of carbon dioxide.

Safety

Wear eye protection.

Questions

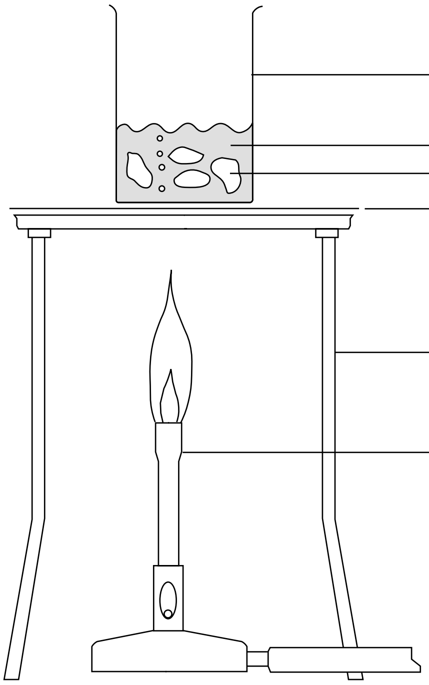
1. Which of the four test-tubes contained carbon dioxide at the end of the experiment?
2. Is air or carbon dioxide more dense?
3. Does this experiment support the idea that the particles of a gas are in motion? Give your reasons.

RS•C

Making a pH indicator

Introduction

A pH indicator is a substance that has a different colour when added to acid or alkali. In this experiment a pH indicator is made from red cabbage.



What to record

Record the colour of the cabbage indicator in the three solutions.

What to do

1. Boil about 100 cm^3 of tap water in a beaker.
2. Add three to four pieces of red cabbage to the boiling water.
3. Boil for about 5 min. The water should have turned blue or green.
4. Turn off the Bunsen burner and allow the beaker to cool for a few minutes.
5. Place three test-tubes in a rack. Half fill one with alkali, one with acid and one with deionised water.
6. Decant approximately 2–3 cm height of cabbage solution into each test-tube.

Safety

Wear eye protection.

Questions

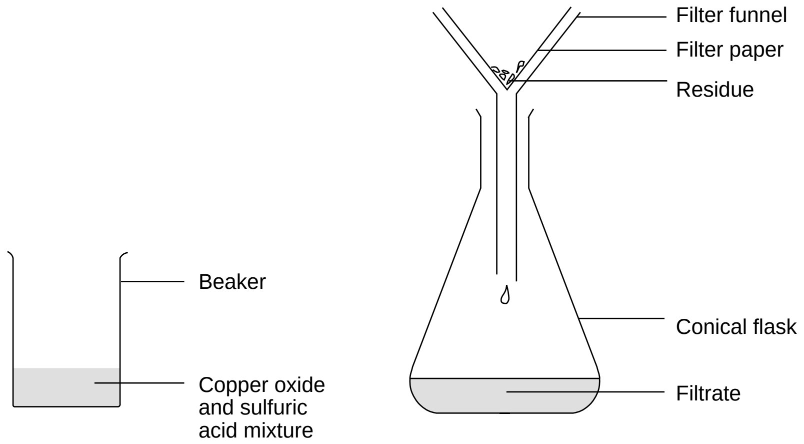
1. What colour is the cabbage indicator when neutral?
2. What colour is the cabbage indicator when alkali?
3. What colour is the cabbage indicator when acid?

RS•C

The reaction between a metal oxide and a dilute acid

Introduction

Many metal oxides react with dilute acid. In this experiment copper(II) oxide is reacted with dilute sulfuric acid.



What to do

1. Pour about 20 cm^3 of dilute sulfuric acid into a beaker.
2. Heat on a tripod and gauze using a Bunsen burner until the acid just boils.
3. Add copper(II) oxide to the hot acid, a spatula measure at a time and stir after each addition, continue until no more dissolves.
4. Filter the mixture, while still hot, into a conical flask.

Safety

Wear eye protection.

Questions

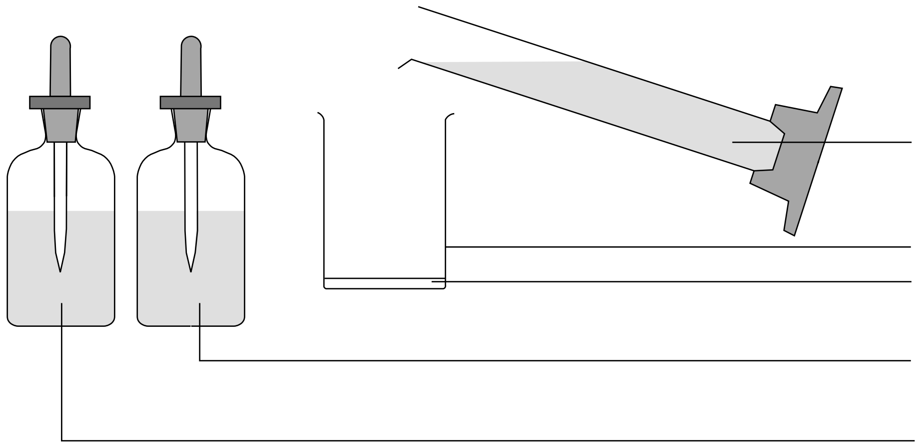
1. What is the residue on the filter paper?
2. What does the filtrate contain?
3. Write an equation for the reaction between solid copper(II) oxide and dilute sulfuric acid.
4. How could the solid product be isolated from the filtrate?

RS•C

Disappearing ink

Introduction

A blue liquid is made. This liquid is tested on a white page, it leaves a blue ink spot. In a few seconds, the blue spot disappears.



What to do

1. Place 10 cm^3 of ethanol (**Highly flammable**) in a small beaker.
2. Add a few drops of thymolphthalein (**Highly flammable**) indicator solution.
3. Add just enough NaOH solution (**Irritant**), dropwise, to produce a deep blue colour in the solution.
4. Using a small paint brush test the 'disappearing ink' on a white page.

Safety

Wear eye protection. Care with teat pipettes that contain sodium hydroxide.

Questions

The colour change occurs because sodium hydroxide reacts with a gas in the air.

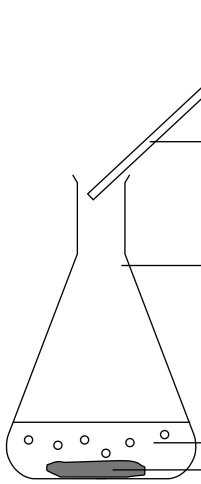
1. Which gas in the air causes this colour change?
2. Write a word equation for the reaction.
3. Write a formula equation for the reaction.

RS•C

Testing for enzymes

Introduction

Enzymes are biological catalysts, they increase the speed of a chemical reaction. They are large protein molecules and these enzymes are very specific to certain reactions. Hydrogen peroxide decomposes slowly in light to produce oxygen and water. There is an enzyme called catalase that can speed up (catalyse) this reaction.



What to record

What do you see? What gas is produced, and which enzyme source makes the most effective catalyst?

What to do

1. Using a measuring cylinder, put 25 cm^3 of hydrogen peroxide solution into a conical flask.
2. Add a small piece of liver.
3. Test the gas given off with a glowing splint.
4. Dispose of this mixture, including the liver, into a bucket, and put another 25 cm^3 of hydrogen peroxide solution in the flask.
5. Add a small piece of potato.
6. Test the gas given off with a glowing splint.
7. Repeat this experiment with a piece of celery instead of potato.

Safety

Wear eye protection.

Questions

1. Which gas is produced in this reaction?
2. What is the test for this gas?
3. Which enzyme source produces the fastest reaction (liver, potato or celery)?
4. Write a word equation for this reaction.
5. How could the rate of gas production be measured?

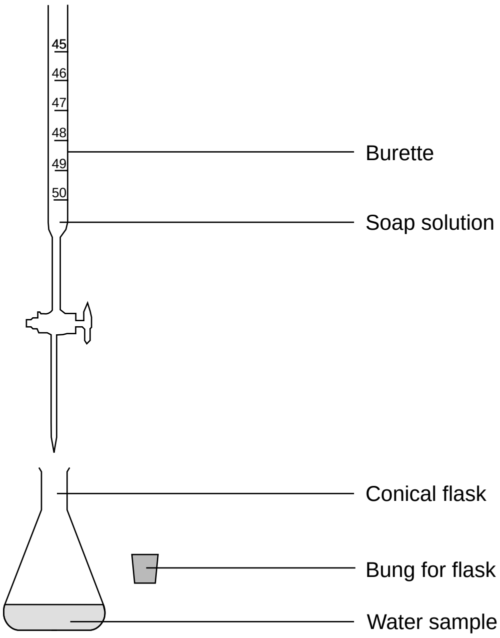
RS•C

Testing water hardness

Introduction

Tap water in some parts of the country is very pure and is said to be 'soft'. It easily makes a lather with soap. Water from other parts may contain various dissolved impurities and is described as 'hard' water. Temporary hardness may be removed by boiling, but permanent hardness survives the boiling process.

In this practical activity, water hardness can be measured by finding out the volume of soap solution required to form a permanent lather with a known volume of water.



What to record

Record the volume of soap needed to produce a permanent lather with each type of water. Note any difference between the appearance of the samples after the addition of soap solution.

Water type	**Volume of soap required a to produce permanent lather /cm^3**
Rainwater	
Seawater	
Temporary hard water	
Seawater, boiled then cooled (permanent hard water)	
Temporary hard water, boiled then cooled	

RS•C

What to do

1. Collect a conical flask and bung. Check the bung is a good fit.
2. Measure 10 cm^3 of water sample into a conical flask using a measuring cylinder.
3. Using the burette add 1 cm^3 of soap solution to the water. Stopper the flask and shake vigorously. If no lather is produced, add another 1 cm^3 of soap solution. Continue in this way until a permanent lather (one that lasts for 30 seconds) is obtained. Record the volume of soap solution needed to produce a permanent lather.
4. Repeat this procedure for the other water samples.

Safety

Wear eye protection.

Questions

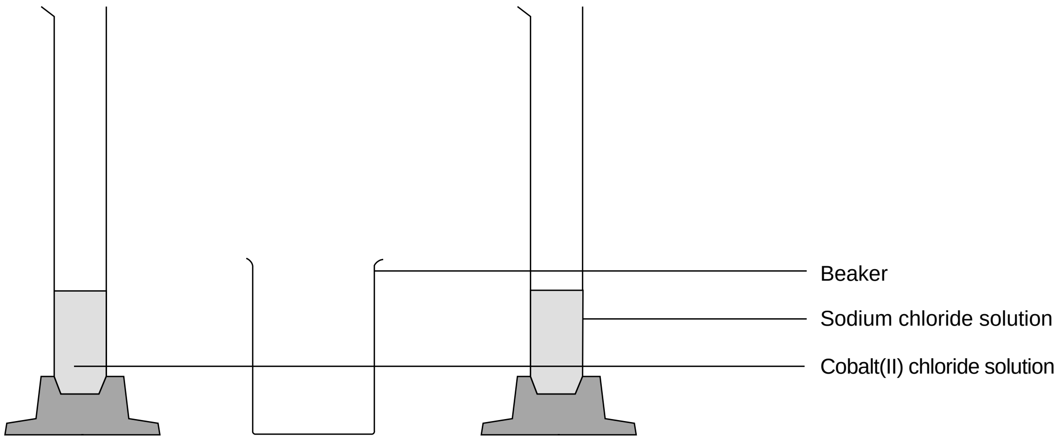
1. Is the rainwater hard or soft?
2. Is seawater hard or soft?
3. Does seawater contain temporary hardness, permanent hardness or both?

RS•C

A chemical test for water

Introduction

Some chemicals change colour when water is added to them. Some coloured chemicals owe their colour to the water molecules that are associated with them. Cobalt(II) chloride is one colour when dry and another colour when damp. In this experiment these colours are identified.



What to do

1. Add 4 cm^3 of cobalt(II) chloride solution to a small beaker.
2. Add 4 cm^3 of salt solution.
3. Dip half of the filter paper into the solution, using tongs.
4. Boil a beaker of water and carefully lift the beaker off the tripod onto the bench.
5. Allow the paper to dry. Wrap it around the beaker of hot water to speed up drying.
6. Observe differences in colour between the wet and the dry cobalt(II) chloride paper.
7. Place the dry cobalt(II) chloride paper near an open window on a humid day and observe what happens.

Safety

Wear eye protection.

Avoid contact with the cobalt(II) chloride paper and the cobalt(II) chloride solution. Cobalt(II) chloride is toxic.

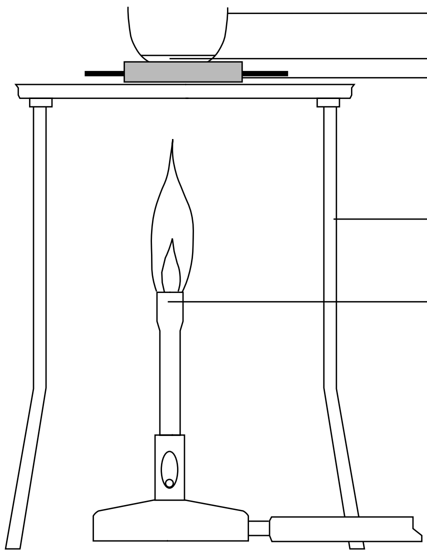
Questions

1. For what other purposes might the sodium chloride be on the paper, in addition to supplying more chloride ions?
2. How many times can the cobalt(II) chloride paper cycle between colours?
3. Suggest a practical application for the cobalt(II) chloride paper.

Forming glass

Introduction

The aim of this experiment is first to make some glass and secondly to make some coloured glass by adding other compounds to the molten glass mixture.



What to record

Record the colour of the glass produced when a speck of a particular oxide is added.

What to do

1. Weigh 6.5 g of lead(II) oxide, 3.5 g of boric acid and 0.5 g of zinc oxide. Mix thoroughly in the boiling tube.
2. Pour the mixture into a crucible and place it on a pipe clay triangle.
3. Heat strongly until it becomes molten and runny. Using tongs pour one or two drops of the molten glass onto your heatproof mat.
4. Allow the beads to cool for 5 min and then examine them.
5. Using an unfolded paper clip pick up a tiny amount of transition metal oxide and drop it into the remaining molten mixture. Stir in the powder using the paper clip.

Do not add too much powder or you will produce a very dark piece of glass.

Safety

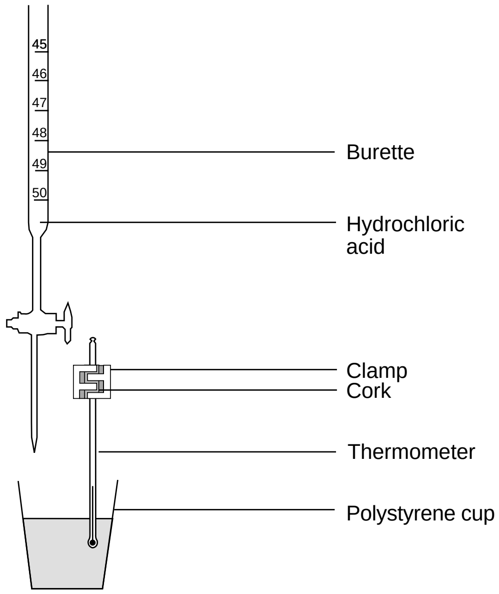
Wear eye protection. Care must be taken with the lead(II) oxide, as it is toxic. Avoid raising the dust. Some of the other chemicals are harmful.

RS•C

Thermometric titration

Introduction

The aim of this experiment is to measure the maximum temperature reached during the reaction between hydrochloric acid and sodium hydroxide solution. The solutions of acid and alkali do not have the same concentration. The volumes that have reacted at the highest temperature reached, represent the ‘end point’ of the titration.



What to record

Record your results in a table.

Volume of acid added/cm^3	Temperature/ °C

What to do

1. Using a measuring cylinder, place 15 cm^3 of sodium hydroxide (**Corrosive**) into the polystyrene cup and measure the temperature.
2. Using the burette add a small portion (3–5 cm^3) of dilute hydrochloric acid to the solution in the polystyrene cup. Swirl the solution and measure the highest temperature reached.
3. Immediately add a second small portion of the dilute hydrochloric acid, swirl and again measure the highest temperature.
4. Continue in this way until there are enough readings to decide the highest temperature for the experiment.

Safety

Wear eye protection.

Questions

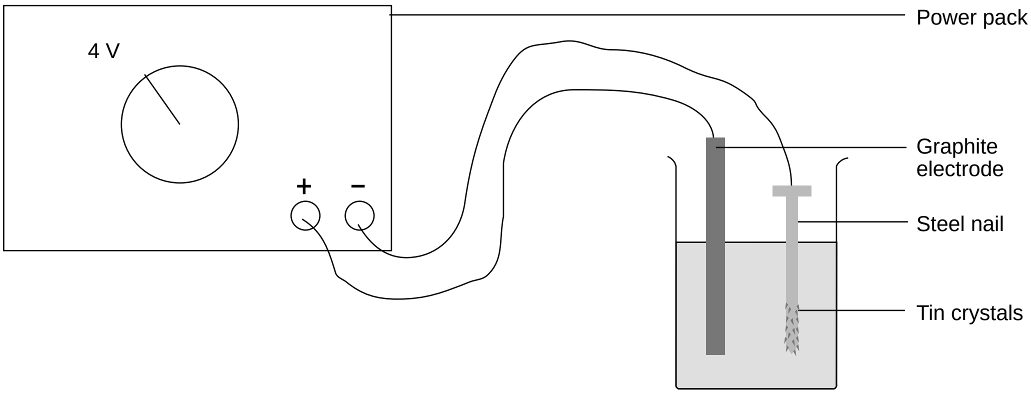
1. What is the highest temperature reached in this reaction?
2. Draw a graph of your results.

RS•C

The formation of metal crystals

Introduction

Metal crystals can be grown using several methods – *eg* displacing one metal by another from a salt solution, by cooling a liquefied metal and by electrolysing a salt solution. This experiment illustrates the electrolysis of tin(II) chloride solution.



What to record

What was done and what was observed.

Note your observations at each electrode.

What to do

1. Set up an electrolysis experiment as shown in the diagram.
2. Ensure the nail is connected as the cathode (-ve terminal) and the carbon rod as the anode (+ve terminal).
3. Set the voltage to 4 V
4. Observe the formation of tin crystals on the nail cathode.

Safety

Wear eye protection. Do not inhale the fumes produced. Do not electrolyse for too long. The tin(II) chloride is dissolved in strong acid (**Irritant solution**).

Questions

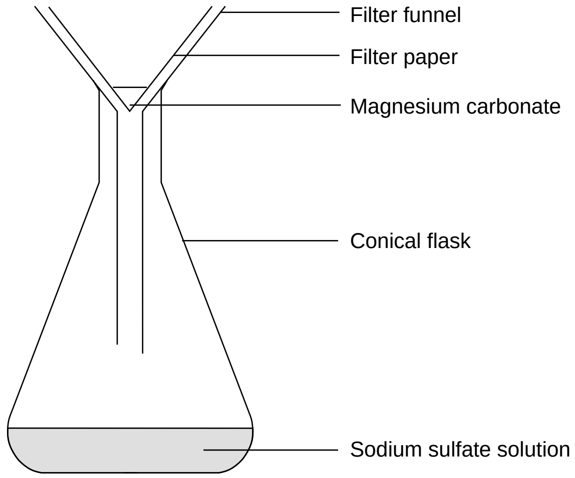
1. What is the product formed at the carbon anode?
2. Give one common household use of steel electroplated with tin.
3. Write an ionic equation for the reaction at the cathode.

RS•C

Formation of a salt which is insoluble in water

Introduction

When solutions of two soluble salts are mixed, a solid may form. The solid is called a precipitate, and the reaction is called a precipitation reaction. Precipitation reactions are used to make insoluble salts.



What to record

What was done and what was observed.

What to do

1. Mix 25 cm^3 of magnesium sulfate solution and 25 cm^3 of sodium carbonate solution in a conical flask.
2. Filter the mixture to remove the spectator ions.
3. Remove the filter paper with the magnesium carbonate and leave to dry.

Safety

Wear eye protection.

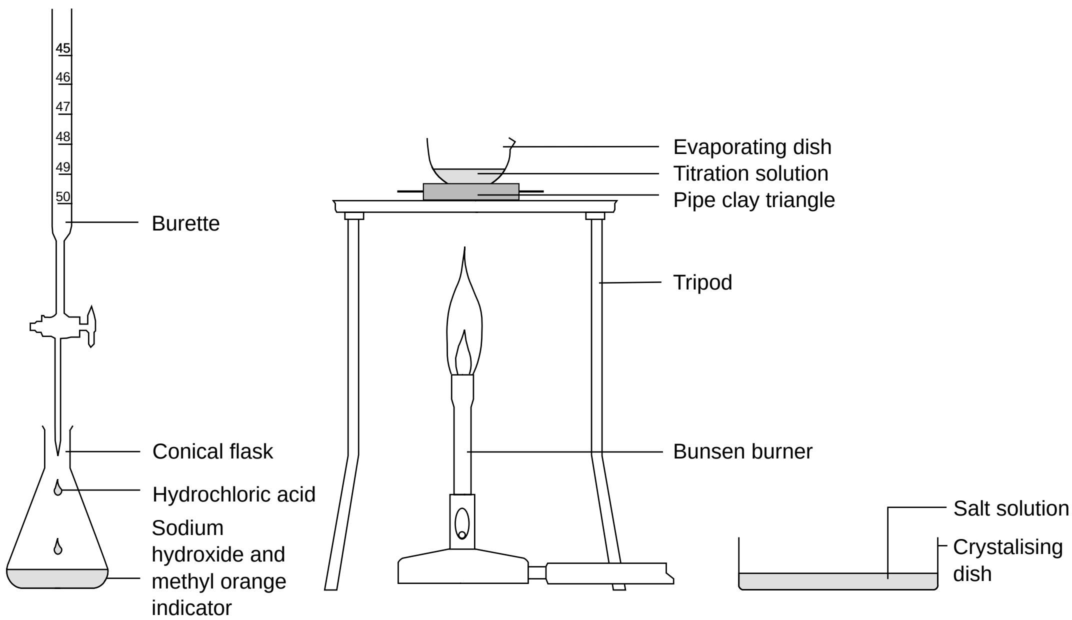
Questions

1. Write a word equation for this reaction.
2. Write a formula equation for this reaction.
3. Write an ionic equation for this reaction.
4. What is meant by spectator ions?

Titration of sodium hydroxide with hydrochloric acid

Introduction

In this experiment sodium hydroxide is neutralised with hydrochloric acid to produce the soluble salt sodium chloride. This is then concentrated and crystallised in a crystallising dish.



What to record

What was done and what was observed.

What to do

1. Add 25 cm^3 of sodium hydroxide solution (**Corrosive**) to a conical flask using a measuring cylinder and add a couple of drops of methyl orange indicator.
2. Fill the burette with hydrochloric acid and run through to the zero mark (use a funnel to fill the burette and a beaker to collect the excess acid).
3. Add the hydrochloric acid to the sodium hydroxide solution in small volumes swirling after each addition. Continue until the solution turns red and record this reading on the burette.
4. Carefully add this volume of fresh hydrochloric acid to another 25 cm^3 of sodium hydroxide solution to produce a neutral solution.
5. Reduce to about half the volume using an evaporating dish on a gauze over a Bunsen burner flame.
6. Leave to evaporate in a crystallising dish to produce a white crystalline solid.

RS•C

Safety

Wear eye protection.

Questions

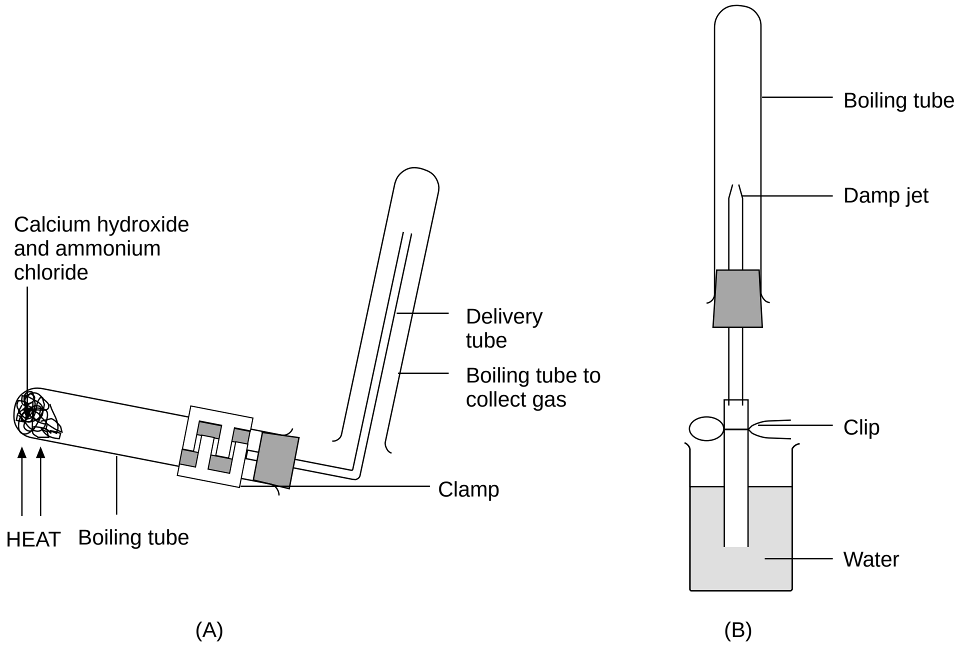
1. What is the everyday name for sodium chloride?
2. This reaction is a specific example of the general reaction:
   Acid + alkali → salt + water.
   Write a word equation for this specific reaction.
3. Write a formula equation for this reaction.
4. Why must you use another 25 cm^3 of sodium hydroxide solution to make *pure* sodium chloride?

RS•C

The properties of ammonia

Introduction

In this experiment, ammonia is produced and collected. The gas is tested and the solubility in water is illustrated by the use of a fountain experiment.



What to record

Observe what happens at each stage. Record the results in the table.

Test	Observations
Heating the mixture	
Lighted splint test	
Damp red litmus paper	
Damp Universal Indicator paper	
Hydrochloric acid bottle stopper	
Opening the clip	

What to do

1. Add two spatulas of calcium hydroxide and two spatulas of ammonium chloride to a boiling tube and mix them.
2. Set up the apparatus as shown in the diagram (A). Warm gently.

3. Test the gas produced with a lighted splint
4. Test the gas with damp red litmus paper
5. Test the gas with damp Universal Indicator paper
6. Test the gas with a stopper from a hydrochloric acid bottle.
7. Fill a dry boiling tube with the gas by heating for several minutes. (The tube *must* be dry.)
8. Fit the tube quickly with a bung carrying a damp glass jet.
9. Set up the apparatus as shown in diagram (B).
10. Open the clip.

Safety

Wear eye protection. Take care not to inhale the ammonia produced. Work in a well ventillated area.

Questions

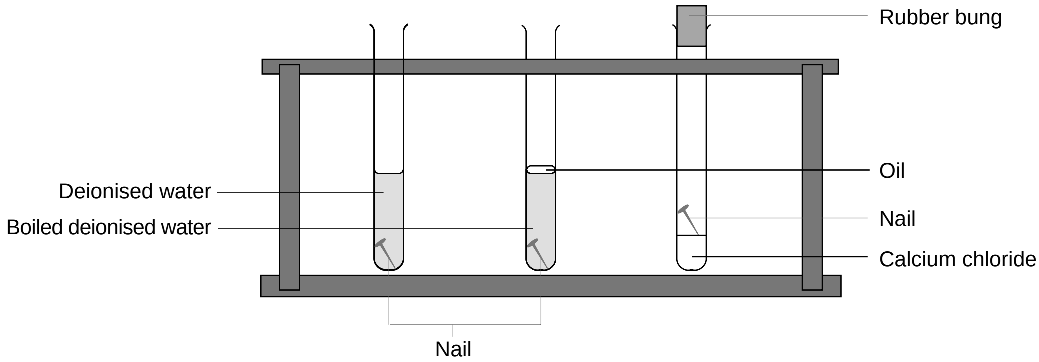
1. Fill in the missing gaps: Ammonia is a __________ gas with a strong, choking _________. It does not ___________ and puts out a lighted _____________. It is ____ _____ than air. It is _____________ to litmus and Universal Indicator. It is very __________ in water, as shown by the ____________ experiment. It reacts with hydrochloric acid to form ______________ chloride.
2. Complete the following word equations:
   Ammonium chloride + calcium hydroxide → ______ + water + _______

RS•C

Causes of rusting

Introduction

Rusting of iron and steel is a commonly occurring process with which we are all familiar. This experiment investigates the conditions needed for rusting to occur.



What to record

Observe what happens at each stage. Complete the table.

Tube number	Conditions	Observations when re-examined
1		
2		
3		

What to do

1. Place a clean nail into a test-tube that contains a little deionised water.
2. Place a clean nail into a test-tube that contains a little boiled deionised water. Pour about 1 cm depth of oil onto the surface.
3. Place about 2 cm depth of anhydrous calcium chloride granules into a test-tube, add a nail on top and place a bung on the tube.
4. Leave the tubes for more than three days and then re-examine the nails.

Safety

Wear eye protection.

Questions

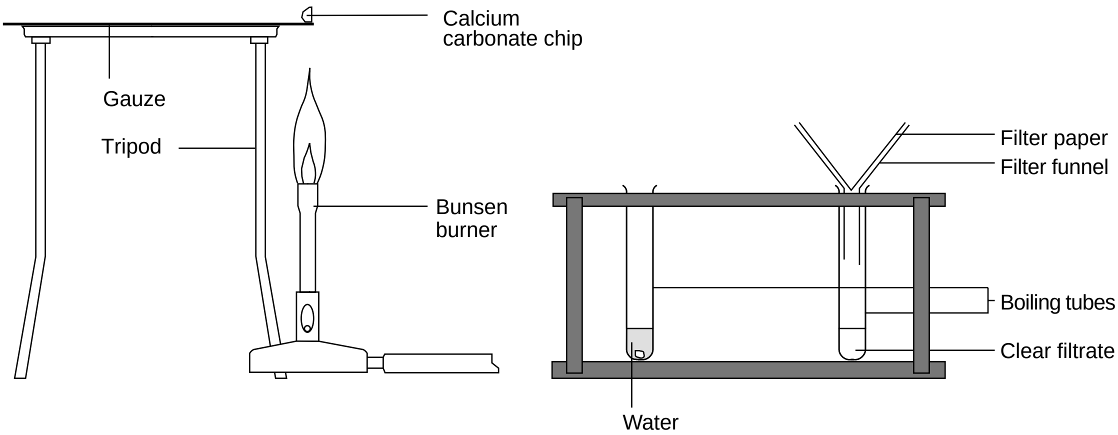
1. Explain why the water is boiled, and oil added in tube 2.
2. What conditions are required for rust to form?
3. Suggest another method to prevent rust formation.

RS•C

Reactions with calcium carbonate

Introduction

Limestone and chalk are mainly calcium carbonate. In this experiment, calcium carbonate is heated to form calcium oxide. This is reacted with a few drops of water, and the resulting calcium hydroxide is dissolved in water. Carbon dioxide is bubbled through the water and the milky suspension of calcium carbonate characteristic of limewater and carbon dioxide is observed.



What to record

Observe what happens at each stage. Complete the table.

Method	Observation
Heat for 10 mins	
Add 2–3 drops of water	
Blow bubbles through solution	
Add Universal Indicator	

What to do

1. Set the chip of calcium carbonate, $CaCO_3$, on a gauze. If your gauze has a coated circle use the edge where there is no coating. Heat strongly for 10 minutes.
2. Let the chip cool and use tongs to move to a boiling tube. Add 2–3 drops of water with a dropper.
3. Add about 10 cm^3 more water to the solid. Then filter half the mixture into the other boiling tube.
4. Gently blow a few bubbles through the filtrate.
5. Test the remaining half with Universal Indicator solution.

Safety

Wear eye protection. Take care not to suck the limewater into your mouth.

RS•C

Questions

Write word equations for the reactions that occur at the following stages.

1. Calcium carbonate is heated.
2. Water is added to the product.
3. Carbon dioxide is bubbled through limewater.

Write formula equations for the reactions that occur at these stages.

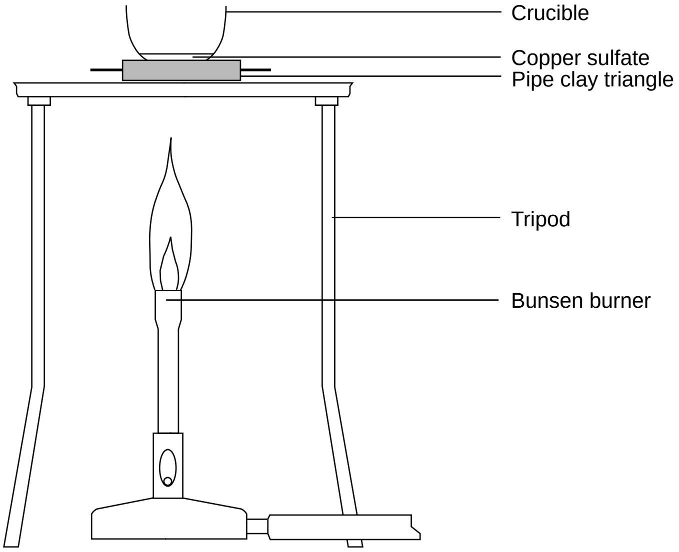
4. Calcium carbonate is heated.
5. Water is added to the product.
6. Carbon dioxide is bubbled through limewater.

RS•C

To find the formula of hydrated copper(II) sulfate

Introduction

In this experiment, the water of crystallisation is removed from hydrated copper(II) sulfate. The mass of water is found by weighing before and after heating. This information is used to find *x* in the formula: $CuSO_4.xH_2O$.



What to record

Complete the table.

Relative atomic mass H=1, O=16, S=32, Cu=64

It is necessary to calculate the relative molecular mass of H_2O and $CuSO_4$.

What to do

1. Find the mass of your crucible.
2. Place 2–3 spatulas of blue copper(II) sulfate in the crucible and weigh.
3. Heat until the powder has gone completely white, but do not heat so strongly that it starts to blacken.
4. Allow to cool then reweigh.

RS•C

(*a*) Mass of crucible	–	g
(*b*) Mass of crucible + blue copper(II) sulfate	–	g
(*c*) Mass of crucible + white copper(II) sulfate	–	g
(*d*) Mass of blue copper(II) sulfate	(*b*-*a*)	g
(*e*) Mass of white copper(II) sulfate	(*c*-*a*)	g
(*f*) Mass of water	(*d*-*e*)	
(*g*) Moles of white copper(II) sulfate	*e*/RMM ($CuSO_4$)	
(*h*) Moles of water	*f*/RMM(H_2O)	
(*I*) Moles of water/Moles copper(II) sulfate	*h*/*g*	
(*j*) Formula of hydrated copper(II) sulfate.	–	

Safety

Wear eye protection.

Question

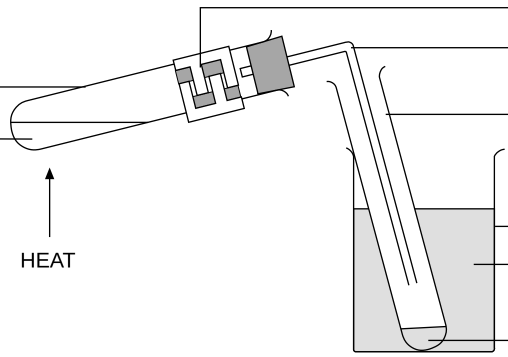
1. What is the formula of hydrated copper(II) sulfate?

RS•C

Heating copper(II) sulfate

Introduction

In this experiment the water of crystallisation is removed from hydrated blue copper(II) sulfate. After cooling the anhydrous copper(II) sulfate formed is then rehydrated with the same water.



What to record

Record any observations when the water was poured back onto the white copper(II) sulfate.

What to do

1. Set up the apparatus as shown.
2. Heat the blue copper(II) sulfate until it has turned white.
3. Act quickly to prevent suck back. Lift the clamp stand so that the delivery tube does not reach into the water in the test-tube.
4. Allow the anhydrous copper(II) sulfate to cool.
5. Hold the tube containing anhydrous copper(II) sulfate in one hand and pour the condensed water onto the powder.

Safety

Wear eye protection.

Questions

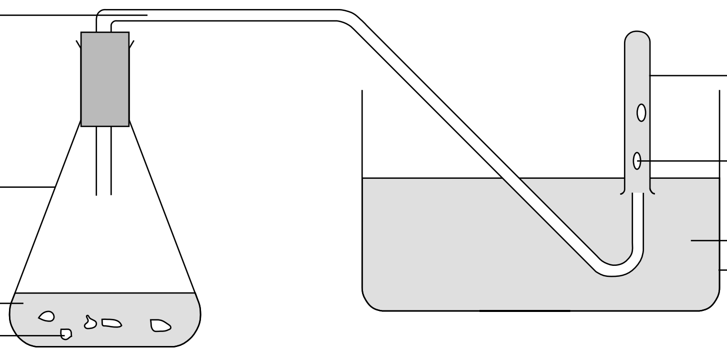
1. Why is one test-tube placed in a beaker of cold water?
2. What do the following words mean (*a*) hydrated, (*b*) anhydrous, (*c*) product, (*d*) condensed and (*e*) reaction?
3. The reaction
   Hydrated copper(II) sulfate + heat $\rightleftharpoons$ anhydrous copper(II) sulfate + water
   is called a reversible reaction. Why?
4. Anhydrous copper(II) sulfate could be used as a fuel for heating ('just add water to get the heat'). Explain why it would not be a very economical fuel.

RS•C

The oxidation of hydrogen

Introduction

In this experiment hydrogen is burnt with some air. The aim of the experiment is to find out how much air is needed to burn hydrogen most efficiently.



What to record

How loud/shrill is the pop for each mixture.

What to do

1. Mark one test-tube where it is a quarter full. Mark one at half full, another at three quarters full.
2. Set up the equipment as shown – a few drops of copper(II) sulfate speeds up the hydrogen production.
3. Using the mark as a guide, fill one test-tube a quarter, one a half, one three quarters and another completely full of water.
4. Invert the quarter filled tube in the trough to collect the hydrogen by displacing the water.
5. Put a bung in the test-tube.
6. Repeat this with the tubes half full, three quarters full and completely full of water.
7. Keep the 4 test tubes in the test-tube rack and clear away the hydrogen generator.
8. Using a lighted splint, carry out the ‘pop’ test on each of the tubes.

Safety

Wear eye protection.

Questions

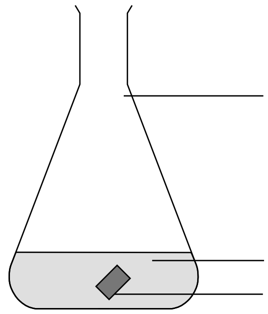
1. Which mixture of hydrogen and air gives the most powerful explosion?
2. The formula of the product is H_2O, what is the ideal mixture of hydrogen and oxygen?
3. If the air is one-fifth oxygen calculate the ideal mixture of hydrogen to air?

RS•C

Investigating the reactivity of aluminium

Introduction

This experiment illustrates the displacement of copper from copper(II) sulfate solution using aluminium foil.



What to record

Write yes or no.

Observations	Before the sodium chloride is added	After the sodium chloride is added
Bubbles observed		
Colour changes		
Temperature change		
Copper observed		

What to do

1. Measure approximately 20 cm^3 of copper(II) sulfate solution into the conical flask.
2. Add a square of aluminium foil.
3. Look for signs of a reaction.
4. Add a spatula of sodium chloride and stir to dissolve.
5. Observe any changes. If nothing happens, add more sodium chloride. Has displacement of copper from copper(II) sulfate occurred?

Safety

Wear eye protection

Questions

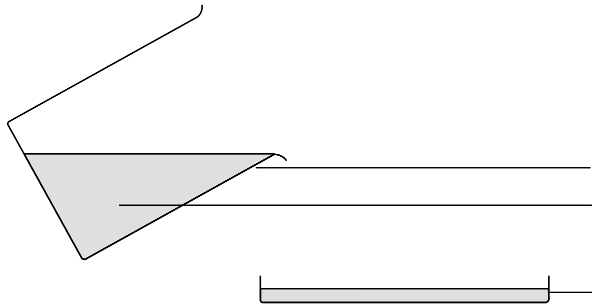
1. Before the sodium chloride is added, does any reaction occur?
2. After adding sodium chloride, does the aluminium appear more or less reactive?
3. How does the salt affect this change?

RS•C

An oscillating reaction

Introduction

Several solutions are mixed in a petri dish. After about 5 min, the colour of the solution oscillates between red and blue.



What to do

1. Prepare the oscillating solution as follows: Place 6 cm^3 of solution A in a small beaker using a measuring cylinder.
2. Add 0.5 cm^3 of solution B using a syringe.
3. Add 1.0 cm^3 of solution C using a syringe. A brown colour appears. When it disappears, add 1.0 cm^3 of ferroin using a syringe.
4. Add 1 drop of Photoflo using the dropper pipette.
5. Add enough of this solution in the beaker to a petri dish to half-fill it.
6. Wait for oscillations to begin.

Safety

Wear eye protection. Solutions may be irritant.

Questions

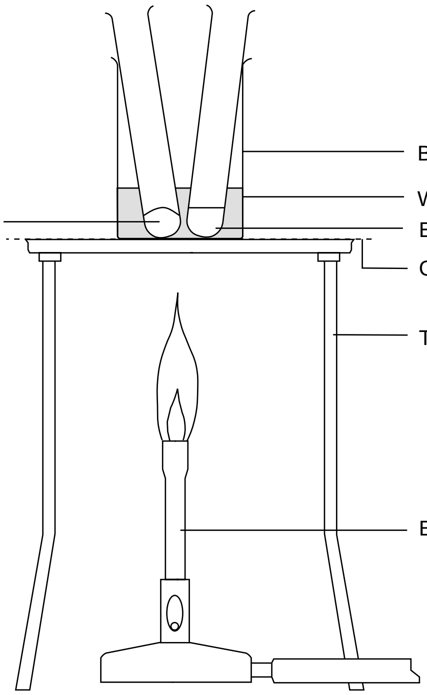
1. What is the role of ferroin in this reaction?
2. What is a surfactant or surface-active agent?

RS•C

Chocolate and egg

Introduction

This experiment shows some changes that happen when different substances are heated.



What to record

Make a table of your observations.

Substance	What happens on heating	What happens on cooling

What to do

1. Set up the equipment as shown in the diagram.
2. Heat the boiling tubes in the beaker of water. The water in the beaker should boil.
3. Watch what happens to the substances in the tubes as they are heated.
4. Turn off the Bunsen burner and use the test-tube holder to transfer the tubes to the rack to cool.
5. Watch what happens to the substances in the tubes as they cool.

Safety

Wear eye protection. Do not taste foods in a laboratory. The food or the apparatus may be contaminated.

Questions

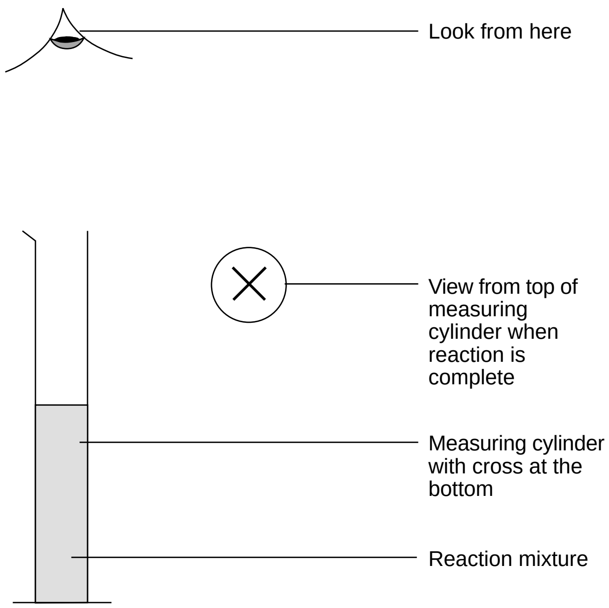
1. Which of these changes is reversible? This is called a physical change.
2. Which of these changes produces a new substance? This is called a chemical change.
3. Give one more example of a physical and a chemical change.

RS•C

Catalysis

Introduction

In this experiment the speed of a reaction is measured. Various metals in solution are tested as possible catalysts.



What to record

Complete the following table.

	No Catalyst	Nickel(II) Sulfate	Copper(II) sulfate	Iron(II) sulfate	Cobalt(II) chloride
Time (s)					

What to do

1. Draw a cross on a piece of scrap paper and put it underneath the measuring cylinder so it can be seen when looking down the cylinder from the top.
2. Using a 50 cm^3 measuring cylinder measure 50 cm^3 of iron(III) nitrate solution.
3. Using a 100 cm^3 measuring cylinder measure 50 cm^3 of sodium thiosulfate solution.
4. Pour the iron(III) nitrate solution into the sodium thiosulfate solution and start the timer.
5. Look through the reaction mixture from above until the cross can first be seen. Stop the timer and record the time.
6. Repeat this experiment but add one drop of catalyst to the iron(III) nitrate solution before mixing. Test the various catalysts and fill in the table.

Safety

Wear eye protection.

Questions

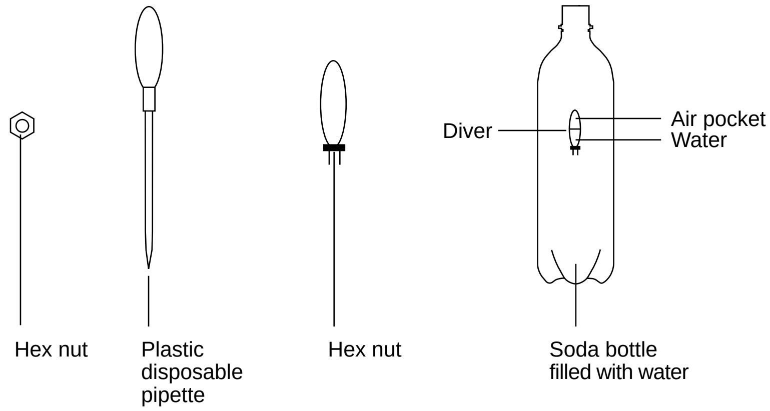
1. Which metal compound is the best catalyst?
2. Why were only very dilute solutions of metal compounds used?
3. A catalyst speeds up a reaction. Suggest one way of slowing down this reaction.

RS•C

A Cartesian diver

Introduction

This is an experiment named after René Descartes (1596–1650). Descartes was a French scientist and philosopher. The Cartesian diver can be used to illustrate the behaviour of gases and liquids when compressed. In this experiment a Cartesian diver is constructed and some of the properties observed.



What to do

1. Screw the hex nut onto the base of the pipette until it is held tightly in place.
2. Cut off all but 1 cm of the pipette stem. (This is the diver.)
3. Place the diver in a beaker of water. Squeeze the bulb of the pipette to force air out and release to allow water up into the diver. Repeat this until the diver is about half full of water.
4. Does the diver still float? If adjusted properly the diver should barely float in the water. If it sinks squeeze a little water out.
5. Carefully transfer the diver to the soda bottle that is full to the brim with water. Take care not to lose water from the diver. Place the cap on the bottle.
6. Use both hands and squeeze the bottle. Watch the diver sink when the bottle is squeezed, or float when pressure is released.

Safety

Wipe up any water spillage.

Questions

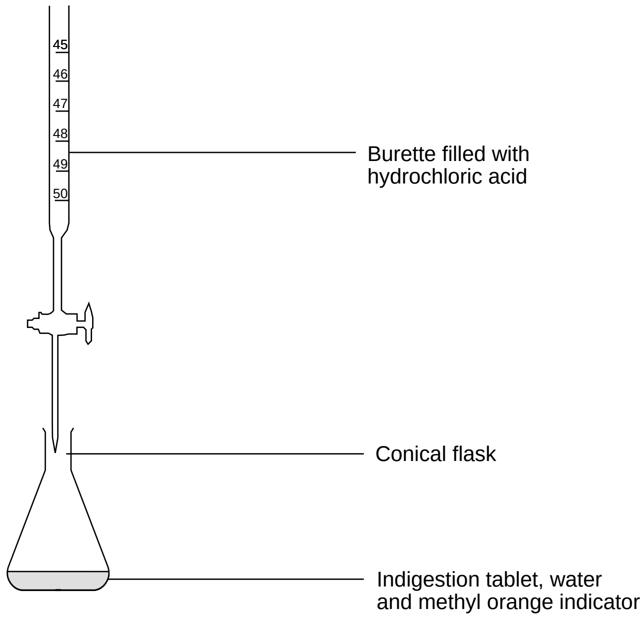
1. What happens to the air in the diver when the bottle is squeezed?
2. Why does the air behave in this way?
3. Write a sentence that explains how the Cartesian diver works.

RS•C

Neutralisation of indigestion tablets

Introduction

Indigestion is caused by excess acid in the stomach. The tablets neutralise some of this acid. In this experiment the amount of acid neutralised by one tablet is found. This may be considered a direct measurement of the effectiveness of the tablet.



What to record

Record the volume and concentration of the acid added.

What to do

1. Crush a tablet using a mortar and pestle and carefully transfer it to a conical flask.
2. Add about 25 cm^3 of water and three drops of methyl orange indicator.
3. While slowly swirling the flask, add acid from the burette, 0.5 cm^3 at a time.
4. Continue adding acid in 0.5 cm^3 portions until the liquid goes red and stays red for one minute.
5. Record the volume of acid used.

Safety

Wear eye protection.

Question

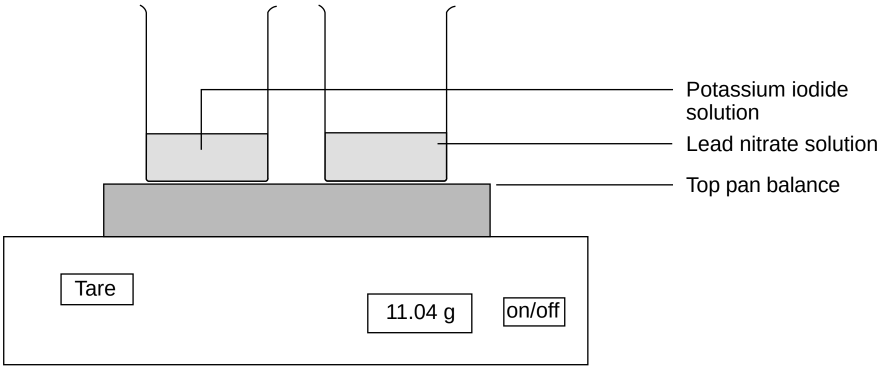
1. Samples of various brands of indigestion tablets and the cost of each packet are provided. Describe how this experiment could be used to determine which brand represents the best value.

RS•C

Mass conservation

Introduction

The aim of this experiment is to show that mass is not gained or lost in a chemical reaction.



What to record

Record the total mass of the reactants and the products.

What to do

1. Measure approximately 5 cm^3 of potassium iodide in the measuring cylinder and pour into one beaker.
2. Rinse the measuring cylinder.
3. Measure approximately 5 cm^3 of lead nitrate in the measuring cylinder and pour into another beaker.
4. Zero the balance then place both beakers on the balance at the same time. Accurately measure their total mass.
5. Take them off the balance.
6. Carefully pour the contents of one beaker into the other making sure there is none spilt.
7. Zero the balance again, place both beakers back on the pan and measure the mass.

Safety

Wear eye protection. Solutions of lead compounds may be toxic – wash your hands thoroughly at the end of the lesson.

Questions

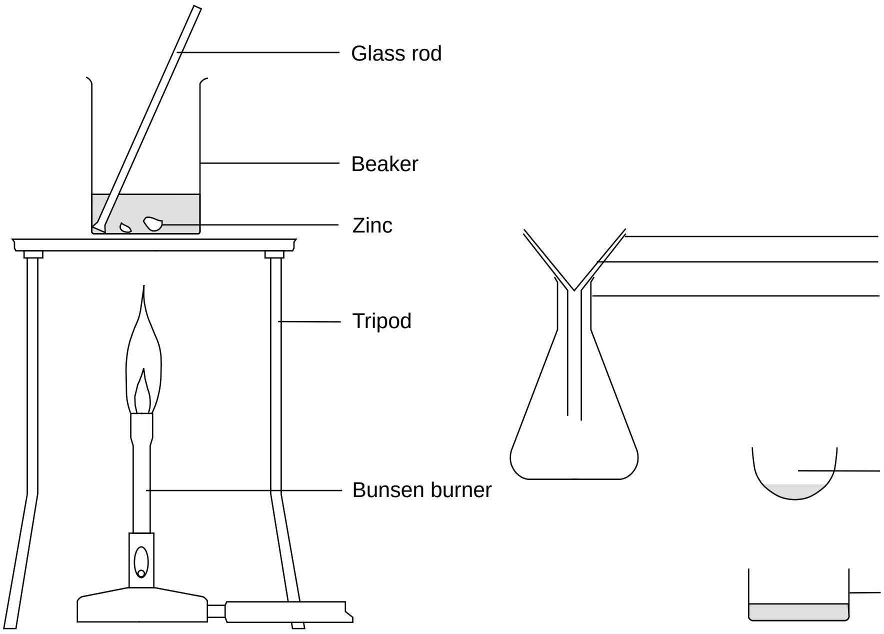
1. Has a chemical reaction occurred?
2. Complete the word equation:
   potassium iodide + lead nitrate → __________ + ___________
3. Comment on your result.

RS•C

Metals and acids

Introduction

Many, but not all, metals react with acids. Hydrogen gas is formed and the metal reacts with the acid to form a salt.



What to do

Lesson 1

1. Measure 50 cm^3 of dilute sulfuric acid and pour it into the beaker. Warm this acid but turn off the Bunsen burner before it reaches the boiling point.
2. Carefully remove the beaker of acid from the tripod and stand it on the bench.
3. To this acid, add two lumps of zinc.
4. If all the zinc reacts, add two more lumps. Add more zinc until no more bubbles form. The acid is now used up.
5. Filter into the conical flask to remove the excess zinc and transfer the filtrate into an evaporating basin.
6. Gently heat the filtrate. Dip in a glass rod and hold it up to cool. When small crystals form on the glass rod stop heating.
7. Pour the solution into a crystallising dish. Label the dish and leave it to crystallise for next lesson.

Lesson 2

1. Examine the crystals with a hand lens.

Safety

Wear eye protection. Care with hot acid.

Questions

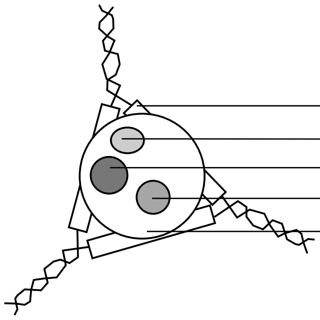
1. Write a word equation for the reaction between zinc and sulfuric acid.
2. Write word equations for the reactions of:
   (a) zinc and hydrochloric acid.
   (b) magnesium and sulfuric acid.
3. Write equations for these three reactions using chemical formulas.

RS•C

Solid mixtures – a lead and tin solder

Introduction

Electrical solder is an alloy of tin and lead. In this experiment a simple method is used to compare the melting points of lead, tin and solder.



What to record

- What was done.
- The order in which they melt.
- The order in which they solidify.

What to do

1. Place a small piece of lead (Pb), tin (Sn), and solder (a Pb/Sn mixture) on an inverted crucible lid over a Bunsen flame. Remember which is which.
2. Observe the three metals, to see which one melts first.
3. When all three are molten, turn off the Bunsen burner and allow the metals to cool.
4. Observe the order in which they solidify.

Safety

Wear eye protection.

Questions

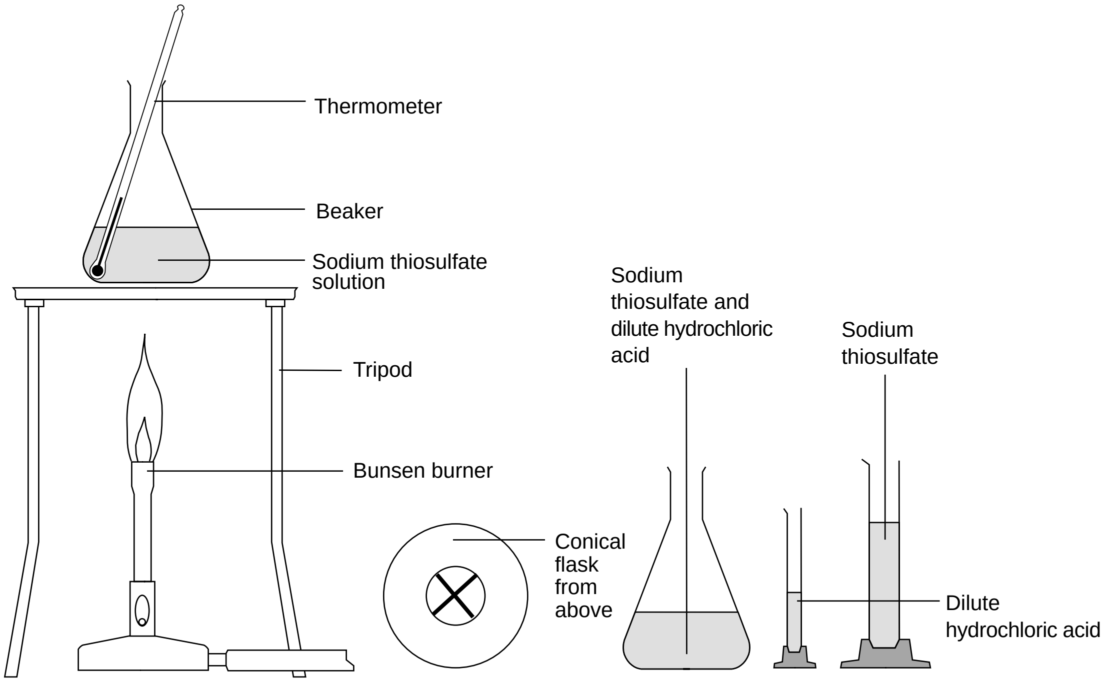
1. Which of the three metals melts first?
2. Which of the three metals melts last?
3. Which of the three metals solidifies last?
4. Which of the three metals solidifies first?
5. Write down the metals in order of their melting points, lowest melting point first.

RS•C

The effect of temperature on reaction rate

Introduction

In this experiment the effect of temperature on the rate of reaction between sodium thiosulfate and hydrochloric acid is investigated.



What to record

Record your results in the table.

Initial temperature of the mixture in the flask/°C	Final temperature of the mixture in the flask/°C	Average temperature of the mixture in the flask/°C	Time taken for the cross to disappear/s	1/time taken /s^{-1}

What to do

1. Put 10 cm^3 of sodium thiosulfate solution and 40 cm^3 of water into a conical flask. Measure 5 cm^3 of dilute hydrochloric acid in a small measuring cylinder.
2. Warm the thiosulfate solution in the flask if necessary to bring it to the required temperature. The object is to repeat the experiment five times with temperatures in the range 15–65 °C.
3. Put the conical flask over a piece of paper with a cross drawn on it.

4. Add the acid and start the clock. Swirl the flask to mix the solutions and place it on a piece of white paper marked with a cross. Take the initial temperature of the mixture.
5. Look down at the cross from above. When the cross disappears, stop the clock and note the time taken. Record the final temperature of the mixture in the flask.
6. As soon as possible, pour the solution down the sink (in the fume cupboard if possible) and wash away.

Safety

Wear eye protection. Take care not to inhale fumes.

Questions

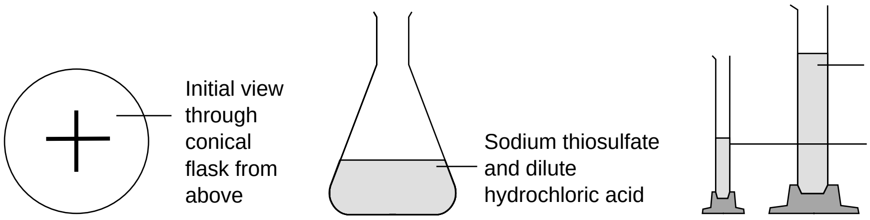
1. For each set of results, calculate the value of 1/time. (This value can be taken as a measure of the rate of reaction for this experiment).
2. Plot a graph of 1/time on the vertical (y) axis and average temperature on the horizontal (x) axis.

RS•C

The effect of concentration on a reaction rate

Introduction

In this experiment, the effect of the concentration of sodium thiosulfate on the rate of reaction is investigated.



What to record

1. Complete the table:

Volume of sodium thiosulfate solution/cm^3	Volume of water/cm^3	Time taken for cross to disappear /s	Original concentration of sodium thiosulfate solution/g dm^{-3}	1/time taken /s^{-1}
50	0		50	
40	10			
30	20			
20	30			
10	40			

What to do

1. Put 50 cm^3 of sodium thiosulfate solution in a flask.
2. Measure 5 cm^3 of dilute hydrochloric acid in a small measuring cylinder.
3. Add the acid to the flask and immediately start the clock. Swirl the flask to mix the solutions and place it on a piece of paper marked with a cross.
4. Look down at the cross from above. When the cross disappears stop the clock and note the time. Record this in the table.
5. Repeat this using different concentrations of sodium thiosulfate solution. Make up 50 cm^3 of each solution. Mix different volumes of the sodium thiosulfate solution with water as shown in the table.
6. As soon as possible, pour the solution down the sink (in the fume cupboard if possible) and wash away.

RS•C

Safety

Wear eye protection. Take care not to inhale fumes.

Questions

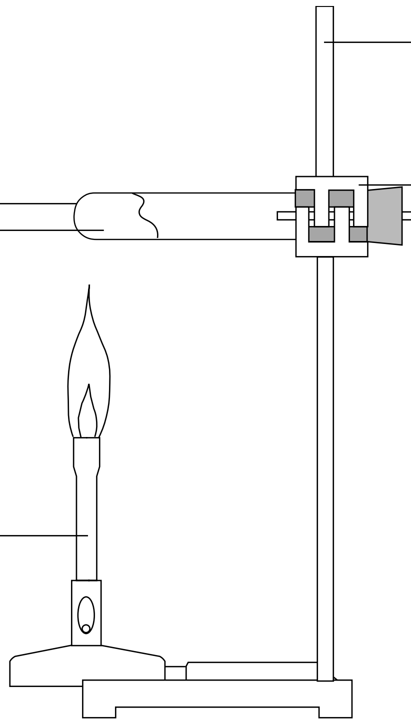
1. Calculate the concentration of sodium thiosulfate in the flask at the start of each experiment. Record the results in the table.
2. For each set of results, calculate the value of 1/time. (This value can be taken as a measure of the rate of reaction).
3. Plot a graph of 1/time taken on the vertical (y) axis and concentration on the horizontal (x) axis.

RS•C

The effect of heat on metal carbonates

Introduction

Metal carbonates decompose when heated. Some carbonates are more reactive than others. The aim of this experiment is to compare the reactivity of some different metal carbonates.



What to record

Complete the table

Carbonate tested	Colour of metal carbonate before heating	Gas evolved, if any	Decomposition easy or difficult	Colour of solid after heating
Sodium carbonate				
Lead carbonate				
Potassium carbonate				
Copper carbonate				

RS•C

What to do

1. Put a large spatula of the carbonate to be tested in a test-tube.
2. Fit a delivery tube and then clamp the tube so that the delivery tube dips into a second test-tube containing limewater.
3. Heat the solid gently at first then more strongly.

Safety

Wear eye protection.

Remove the delivery tube by lifting the clamp stand as soon as heating is stopped. Some metal compounds are toxic. Avoid raising dust. Wash hands thoroughly at the end of the experiment.

Questions

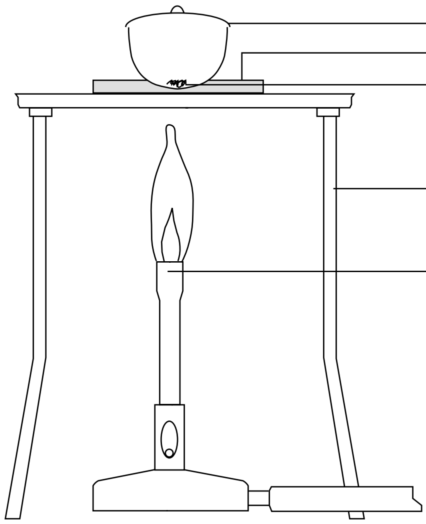
1. Why do some gas bubbles pass through limewater when heating is first started?
2. Why must the delivery tube be removed as soon as heating is stopped?

RS•C

The change in mass when magnesium burns

Introduction

Many areas of chemistry involve careful measurement. One example is measuring the change in mass before and after a chemical reaction. This experiment shows how the mass of magnesium changes when it combines with oxygen.



What to record

What was done.

The mass of the magnesium and the magnesium oxide product.

What to do

1. Clean a 10–20 cm length of magnesium ribbon with emery cloth to remove the oxide layer. Loosely coil it.
2. Weigh a clean crucible and lid. Place the magnesium inside and reweigh.
3. Heat the crucible for 5–10 minutes, lifting the lid a little from time to time with tongs. Ensure that as little product as possible escapes.
4. Continue heating until glowing ceases.
5. Cool the crucible and reweigh.

Safety

Wear eye protection.

Questions

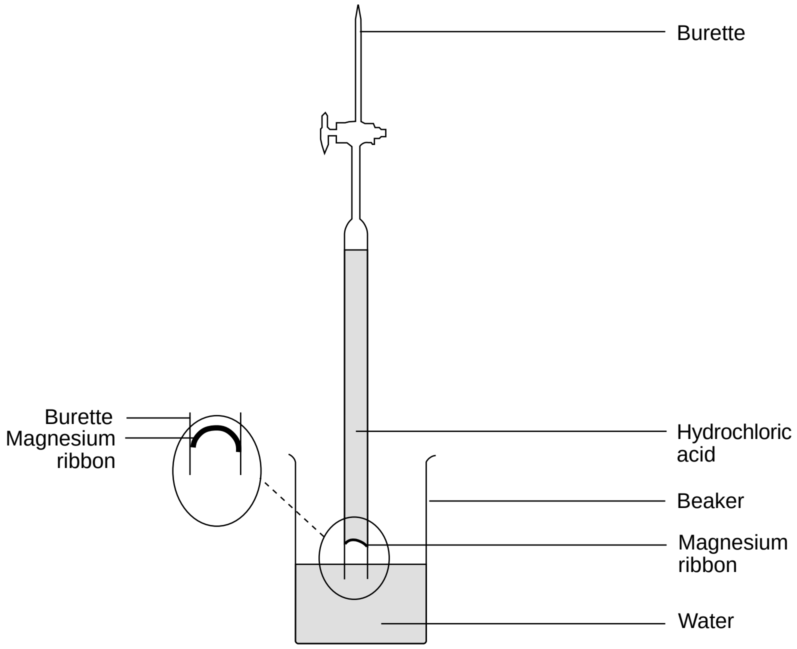
1. Why is the magnesium ribbon cleaned before the experiment?
2. Why is the lid needed?
3. Why is the lid lifted from time to time?
4. How could weighing be used to show when the reaction is complete?

RS•C

The volume of 1 mole of hydrogen gas

Introduction

One mole of any gas occupies the same volume when measured under the same conditions of temperature and pressure. In this experiment, the volume of one mole of hydrogen is calculated at room temperature and pressure.



What to record

What was done.

The mass of magnesium used and the volume of hydrogen produced.

What to do

1. Clean a piece of magnesium ribbon about 3.5 cm long and weigh accurately. (This should weigh between 0.02 and 0.04 g; if not adjust the amount used.)
2. Measure 25 cm^3 of dilute hydrochloric acid (**Irritant**) into the burette. Carefully add 25 cm^3 of water on top of this.
3. Push the magnesium into the end of the burette so it will stay in position with its own tension.
4. Add 50 cm^3 of water to a 250 cm^3 beaker.
5. Quickly invert the burette into the water. If this is done quickly and carefully very little is lost. It is important that the liquid level in the burette starts on the graduated scale. If it is not on the scale; momentarily open the tap, this allows the level to drop). Clamp the burette vertically.
6. Take the burette reading (care: it is upside down!)
7. Observe the magnesium react as the acid diffuses downwards, wait until all the magnesium has reacted.

8. Note the new volume on the burette (care: it is upside down).
9. Record your results.

Safety

Wear eye protection.

Questions

The equation for the reaction is

$Mg + 2HCl \rightarrow MgCl_2 + H_2$

The relative atomic mass of magnesium is 24.

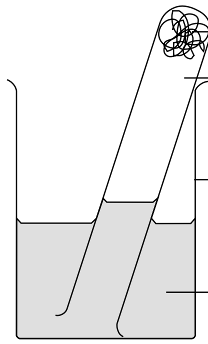
1. Copy out and fill in the gaps:
   ____ g Magnesium has produced ____cm^3 hydrogen
   ____ /24 moles magnesium produces _____ cm^3 hydrogen.
   1 mole magnesium produces ________cm^3 hydrogen which is the volume of one mole of hydrogen gas.

RS•C

How much air is used during rusting

Introduction

This experiment illustrates how much of the air is used in the rusting process. It is the oxygen component of air which reacts in the rusting process. This experiment allows calculation of the percentage of oxygen in the air.



What to record

Initial length of column of air	
Final length of column of air in the tube	

What to do

1. Place approximately 3 cm depth of iron wool in the bottom of a test-tube. Wet the iron wool with water.
2. Invert the test-tube in a beaker of water (approximately 20 cm^3).
3. Measure the length of the column of air.
4. Leave the test-tube for at least one week, and then measure the new length of the column of air. Take care not to lift the test-tube out of the water.

Questions

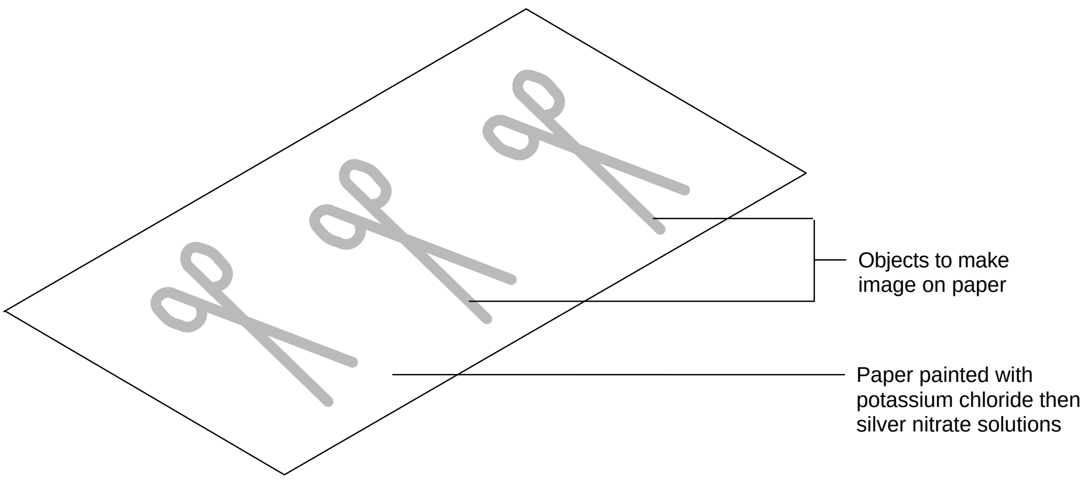
1. Write a word equation for this reaction.
2. Calculate the percentage of oxygen in air.
3. How could it be shown that the reaction is complete?

RS•C

Making a photographic print

Introduction

Only a very small amount of energy is needed to break down silver halide compounds (silver chloride, silver bromide or silver iodide) to the silver metal. This small amount of energy is available from many sources including light, X-rays and radiation from a radioactive substance. The above three silver halides can be used to make photographic film and photographic paper. In this experiment, a photographic print is produced.



What to do

1. In a darkened room, take a piece of paper and paint one side of it with the potassium chloride solution. Dry the paper with a hairdryer.
2. Paint the dried paper, on the same side, with the silver nitrate solution. Dry the paper with a hairdryer.
3. Place your chosen objects(s) on top of the paper and place under an ultraviolet light for 30 min or in sunlight for 2 h.
4. Switch off the ultraviolet light (if used) and remove the objects from the top of the paper. Observe what has happened.

Safety

Wear eye protection and protective gloves. Care with ultraviolet light; do not look directly at the light, it can damage your eyes.

Questions

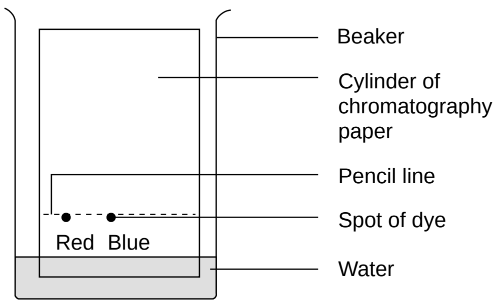
1. What happens to the paper when the silver nitrate is painted onto it?
2. Write a word equation for the above reaction.
3. Explain what happens when the paper is exposed to light.

RS•C

'Smarties' chromatography

Introduction

In this experiment dye is removed from the surface of various Smarties. A spot of each colour is put on a piece of chromatography paper and water is allowed to soak up the paper. The results show which mixtures are used to produce particular colours for the Smarties.



What to record

Record the dyes used to make each colour.

What to do

1. Draw a pencil line 1 cm from the bottom of the chromatography paper.
2. Use a clean paintbrush and clean water to remove the colour from a Smartie. Paint the colour in a small spot on to the line on the chromatography paper.
3. Clean the brush and paint the colour of another Smartie on a small spot about 2 cm from the previous spot. Repeat this until all the colours are on the paper.
4. Using the pencil write the name of each colour by the corresponding spot.
5. Roll the paper into a cylinder, hold in place with paper clips. Put the cylinder in a beaker containing 1 cm of water. Allow the water to rise up the paper.
6. When it reaches the top take the cylinder out of the water, carefully unroll it and examine it.

Safety

Do not eat in the laboratory.

Questions

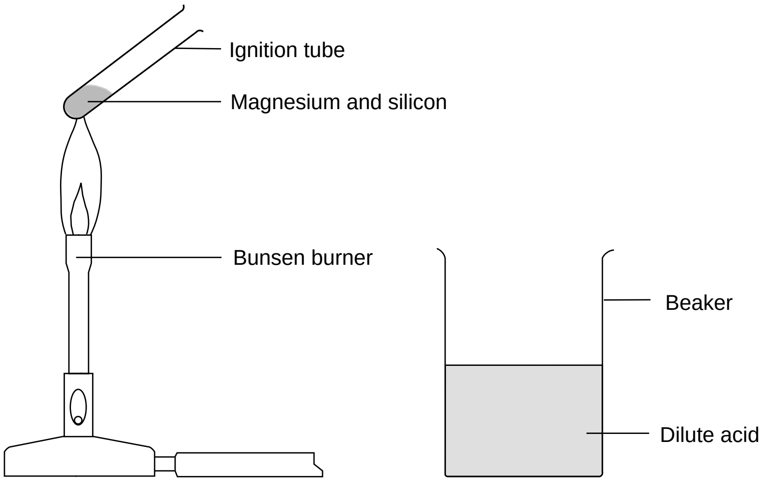
1. Why do some dyes separate into different colours yet others do not?
2. Why do some dyes move further up the chromatography paper than others?
3. Look on the side of the Smarties packet for the list of coloured dyes used. Try and identify which dyes correspond to the spots on the chromatogram.

RS•C

The decomposition of magnesium silicide

Introduction

This experiment illustrates a reaction with low activation energy. Magnesium reacts with silicon to produce magnesium silicide. This then decomposes in dilute acid to produce silane, which spontaneously combusts on contact with air.



What to do

1. Carefully heat equal amounts of silicon and magnesium powder (1 spatula of each) in an ignition tube (behind a safety screen).
2. Use the clamp stand base to crush the cooled tube between two sheets of paper.
3. Put 100 cm^3 of dilute hydrochloric acid into a 1 dm^3 beaker and add 800 cm^3 of tap water.
4. Put the crushed ignition tube with its contents into the dilute acid, stand back and observe. (Do not inhale the fumes, use a fume cupboard if possible, or ensure the room is well ventilated.)

Safety

Wear eye protection.

Questions

The equations for this reaction are:

$2Mg + Si \rightarrow Mg_2Si$

$Mg_2Si + 2H_2O \rightarrow SiH_4(g) + 2MgO$

$SiH_4 + 2O_2 \rightarrow SiO_2 + 2H_2O$

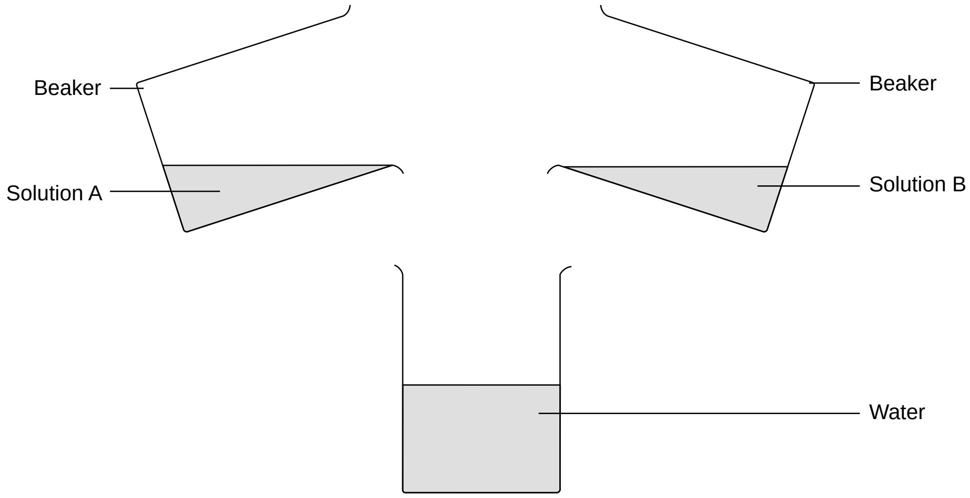
1. What happens when silane meets oxygen?
2. Why does methane (CH_4) not behave in the same way?

RS•C

An example of chemiluminescence

Introduction

Chemiluminescence is the emission of light during a chemical reaction. In this experiment two solutions are mixed to produce chemiluminescence.



What to do

1. Collect 30 cm^3 of solution (A).
2. Collect 30 cm^3 of solution (B).
3. In a dark room: add 10 cm^3 of each of the two solutions A and B simultaneously to 50 cm^3 of water in a 100 cm^3 beaker.
4. Repeat this but add dyes such as Fluorescein or Methylene blue to the water before mixing A and B.

Safety

Wear eye protection.

Questions

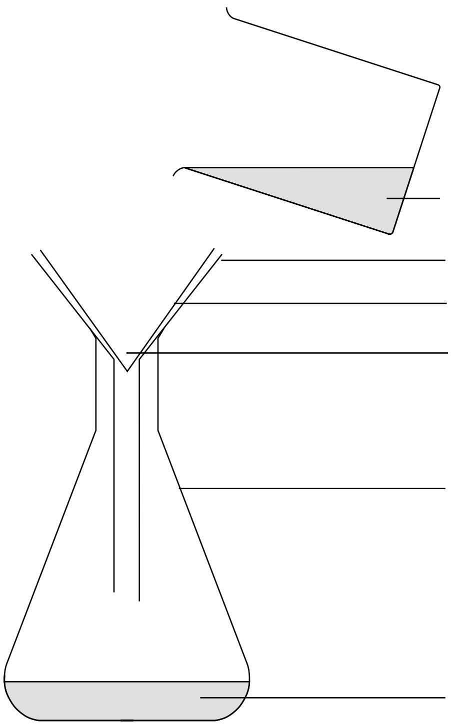
1. Describe one use for chemiluminescence

RS•C

Colorimetric determination of a copper ore

Introduction

An ore is any rock from which a metal may be extracted. Ores contain a mineral of the metal together with waste material. To decide whether an ore is worth mining it is necessary to find out how much of the useful mineral it contains, and how much is waste. This experiment illustrates an example of how this might be done.



What to record

How much copper the ore is estimated to contain.

What to do

1. Weigh 10 g of the ground ore into a beaker.
2. Add 40 cm^3 of 2 mol dm^{-3} sulfuric acid in small amounts. Do not let the mixture go over the top.
3. When the reaction finishes filter the mixture into a conical flask.
4. Add deionised water until the total volume of liquid in the flask is 100 cm^3.
5. Using the laboratory copper(II) sulfate solution, prepare six tubes of diluted copper(II) sulfate, according to the following table. Ensure the solutions are well mixed.

RS•C

Tube number	1	2	3	4	5
Volume of copper(II) sulfate solution/cm^3	8	6	4	2	0
Volume of deionised water/cm^3	2	4	6	8	10

6. Pour a sample of the solution from your conical flask into another test-tube.
7. Compare the colour of your tube from part 6 with those from part 5. Which one matches the colour best?
8. Estimate the mass of copper mineral in 10 g of the ore using the following table:

Tube of best match	1	2	3	4	5
Mass of compound in 10 g of ore/g	10	7.5	5	2.5	0

Safety

Wear eye protection. Dilute sulfuric acid is corrosive. When gases are made in a reaction, a mist of fine acid spray is often produced which is dangerous to your eyes and causes irritation if inhaled.

Questions

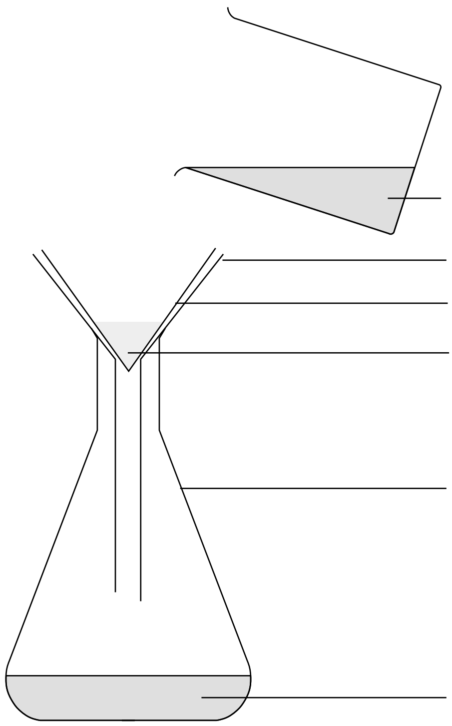
1. Which part of the ore (copper mineral or waste) causes the blue colour of the solutions?
2. Which part of the ore (copper mineral or waste) was removed by filtration in part 3 of the experiment, and why was this done?
3. How could this experiment be adapted to check the result?

RS•C

Glue from milk

Introduction

Glue can be made from the protein in milk called casein. In this experiment, polymer glue is prepared from milk. The casein is separated from milk by processes called coagulation and precipitation.



What to do

1. Place 125 cm^3 of skimmed milk into a 250 cm^3 beaker. Add approximately 25 cm^3 of ethanoic acid (or vinegar).
2. Heat gently with constant stirring until small lumps begin to form.
3. Remove from the heat and continue to stir until no more lumps form.
4. Allow the curds to settle, decant some of the liquid (whey) and filter off the remainder using the filter funnel resting on the 250 cm^3 conical flask.
5. Gently remove excess liquid from the curds using the paper towel.
6. Return the solid to the empty beaker. Add 15 cm^3 of water to the solid and stir.
7. Add about half a teaspoon of sodium hydrogen carbonate to neutralise any remaining acid. (Watch for bubbles of gas to appear then add a little more sodium hydrogen carbonate until no more bubbles appear).
8. The substance in the beaker is glue.
9. Find a way to test your glue.

Safety

Wear eye protection.

Questions

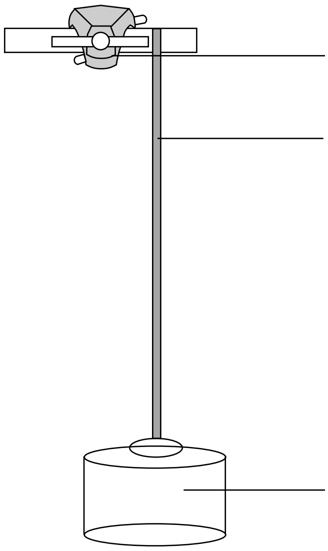
1. What is the purpose of the ethanoic acid (vinegar) in this experiment?
2. Why is sodium hydrogencarbonate added?
3. Write an equation for this reaction between ethanoic acid and sodium hydrogencarbonate.

RS•C

Rubber band

Introduction

This experiment involves an investigation into the effect of heat on a stretched rubber band.



What to record

What was done and what was observed.

What to do

1. Take the rubber band. Quickly stretch it and press it against your lips. Note any temperature change compared with the unstretched band.
2. Now carry out the reverse process. First stretch the rubber band and hold it in this position for a few seconds. Then quickly release the tension and press the rubber band against your lips. Compare this temperature change with the first situation.
3. Set up the apparatus as shown in the diagram. Make sure that if the rubber band breaks the weight cannot drop on toes!
4. Predict what happens if this rubber band is heated with a hair dryer. Write down your prediction. Measure the length of the stretched rubber band.
5. Now heat the rubber band using the hair dryer and observe the result. Does this observation match your prediction? Measure the new length.

Questions

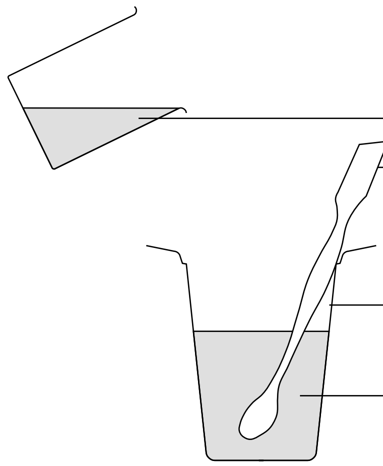
1. Based on your initial testing (by placing the rubber band against your lips) decide which process is exothermic (heat given out): stretching or contracting of the rubber band?
2. The chemist Le Chatelier made the statement '… an increase in temperature tends to favour the endothermic process'. Explain in your own words how this statement and how your answer to question 1 can account for your observations when heating the rubber band.
3. Draw a number of lines to represent chains of rubber molecules showing how they might be arranged in the unstretched and stretched forms. (Hint: the lines of polymer should show less order in the unstretched form than in the stretched form.)

RS•C

Polymer slime

Introduction

A solution of polyvinyl alcohol can be made into a gel (slime) by adding a borax solution, which creates crosslinks between chains. In this activity, some interesting properties of the slime are investigated.



What to record

Results of the tests.

What to do

1. Collect 40 cm^3 of polyvinyl alcohol solution in a disposable cup containing a spatula.
2. If desired add one drop of food colour or fluoroscein dye to the solution. Stir well.
3. Measure 10 cm^3 of borax solution and add this to the polyvinyl alcohol solution. Stir vigorously until gelling is complete.
4. Remove the slime from the cup and pat and knead it thoroughly to completely mix the contents. Roll the slime around in your hand, gently squeezing the material to remove air bubbles at the same time.

Safety

Wear eye protection.

Questions

Test the properties of your slime

1. Pull slowly – what happens?
2. Pull sharply and quickly – what happens?
3. Roll the slime into a ball and drop it on the bench – what happens?
4. Place a small bit on the bench and hit it with your hand – what happens?
5. Write your name on a piece of paper with a water based felt tip pen. Place the slime on top, press firmly, then lift up slime. What happened to the writing? To the slime? Try the same thing using a spirit-based pen. Does this show the same effect?

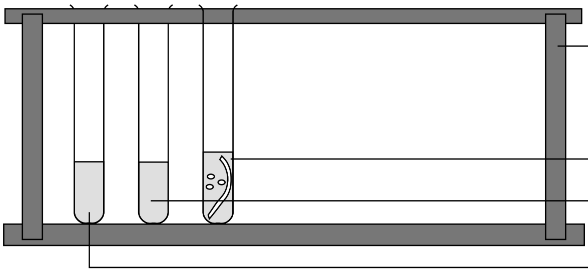
6. Place a small piece of your slime on a watch glass or petri dish. Add dilute hydrochloric acid dropwise, stirring well after each drop. When a change is noticed record the number of drops added and your observations.
7. Now add dilute sodium hydroxide solution dropwise to the same sample used in 6 stirring after each drop. When a change is noticed record your observations. Can the whole process be repeated with tests 6 and 7? Try it!

RS•C

The properties of ethanoic acid

Introduction

Acids are an important group of chemicals. Organic acids are characterised by the presence of a -COOH group attached to a carbon atom. In this experiment, some typical properties of a weak organic acid are observed.



What to record

What was observed.

What to do

1. Use a small quantity (1–2 cm^3 at a time) of dilute ethanoic acid.
2. Observe the effect on full range indicator paper of adding drops of sodium carbonate solution to 2 cm^3 of dilute ethanoic acid.
3. Repeat the experiment with dilute sodium hydroxide solution.
4. Add a small piece of magnesium ribbon to the ethanoic acid. What is observed? Try to confirm the identity of the gas given off.

Safety

Wear eye protection.

Questions

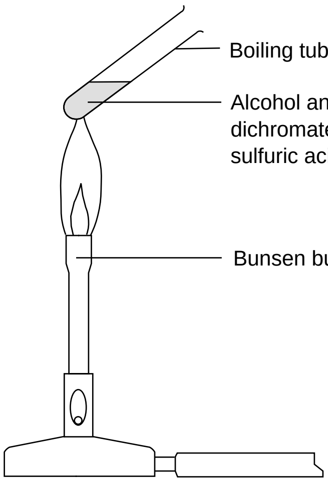
1. Write equations for the reactions of hydrochloric acid (HCl) with: -
   (*a*) sodium carbonate solution (Na_2CO_3)
   (*b*) sodium hydroxide solution (NaOH)
   (*c*) magnesium ribbon (Mg).
2. The formula of ethanoic acid is written as CH_3COOH. Write similar equations for the reactions of ethanoic acid with:
   (*a*) sodium carbonate solution
   (*b*) sodium hydroxide solution
   (*c*) magnesium.

RS•C

Properties of alcohols

Introduction

Alcohols are an important group of organic chemicals. The alcohol people drink is called ethanol and is produced by fermentation. Alcohols are characterised by an -OH group attached to a carbon atom.



What to record

What was observed.

What to do

1. Take a small quantity of ethanol and add the same volume of water. What is the pH of the mixture? Test the mixture with full range indicator solution. Does the water mix with the ethanol?
2. Put a small quantity of ethanol on a tin lid and ignite it with a splint. Does it burn, and if so, describe the flame.
3. Put 5 cm^3 of dilute sulfuric acid in a boiling tube. Add five drops of potassium dichromate(VI) solution. Now add two drops of ethanol and heat the mixture until it just boils. Is there any sign of a reaction? Is there any change of smell that could come from a new compound?

Safety

Ethanol is highly flammable. Potassium dichromate (VI) is toxic. Wear disposable gloves. Avoid skin contact. Wash hands after use.

Questions

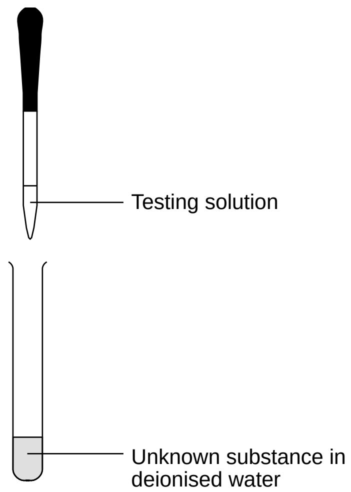
1. What is the name of the process used to produce ethanol on a commercial scale?
2. The reaction of an alcohol to produce an acid is called an oxidation reaction. What is the opposite reaction called that produces an alcohol from an acid?
3. Write a chemical equation for the combustion of ethanol.

RS•C

Testing salts for anions and cations

Introduction

Chemists often have to identify the composition of unknown substances. This experiment involves identifying the cations and anions in various salt solutions.



What to record

Sample	Test performed	Result of test

What to do

1. Dissolve the unknown substance in deionised water. 5–10 cm^3 of solution may be needed.
2. Using the analysis table, test small aliquots (portions).
3. Repeat for the other unknown substances.

Safety

Wear eye protection. Some of the unknowns may be toxic or corrosive.

Questions

1. Write word and ionic equations for those reactions that give a positive result.

RS•C

Testing salts for anions and cations.

For anions: carry out the three tests A,B and C below:

Test	Anion	Test and observation
A Silver nitrate followed by ammonia solution	Chloride (Cl^-)	Add a few drops of dilute nitric acid (**Irritant**) followed by a few drops of silver nitrate solution. A white precipitate of silver chloride is formed. The precipitate is soluble in ammonia solution.
	Bromide (Br^-)	Add a few drops of dilute nitric acid (**Irritant**) followed by a few drops of silver nitrate solution. A pale yellow precipitate of silver bromide is formed. The precipitate is slightly soluble in ammonia solution.
	Iodide (I^-)	Add a few drops of dilute nitric acid followed by a few drops of silver nitrate solution. A yellow precipitate of silver iodide is formed. It is insoluble in ammonia solution.
B Barium chloride	Sulfate (SO_4^{2-})	Add a few drops of barium chloride solution (**Toxic**) followed by a few drops of dilute hydrochloric acid. A white precipitate of barium sulfate is formed.
C Hydrochloric acid	Carbonate (CO_3^{2-})	Add dilute hydrochloric acid to the solution (or add it to the solid). Bubbles of carbon dioxide are given off.

For cations: carry out the two tests D and E below:

Cation	**D Add sodium hydroxide solution (Irritant)**	**E Add ammonia solution**
Ammonium (NH_4^+(aq))	Warm carefully. Do not allow to spit. Ammonia (alkali gas) is given off	————————
Copper (Cu^{2+}(aq))	Blue (jelly-like) precipitate of $Cu(OH)_2$(s)	Blue jelly like precipitate dissolves in excess ammonia to form a deep blue solution.
Iron(II) (Fe^{2+}(aq))	Green gelatinous precipitate of $Fe(OH)_2$(s)	Green gelatinous precipitate
Iron(III), (Fe^{3+}(aq))	Rust-brown gelatinous precipitate of $Fe(OH)_3$(s)	Rust brown gelatinous precipitate
Lead(II), (Pb^{2+}(aq))	White precipitate $Pb(OH)_2$(s) dissolves in excess NaOH(aq)	White precipitate, $Pb(OH)_2$
Zinc (Zn^{2+}(aq))	White precipitate, $Zn(OH)_2$(s)	White precipitate, $Zn(OH)_2$(s) dissolves in excess NH_3(aq)
Aluminium (Al^{3+}(aq))	Colourless precipitate, $Al(OH)_3$(s)	Colourless precipitate, $Al(OH)_3$(s)

RS•C

Flame tests.

1. Slightly open the air hole of the Bunsen burner.
2. Heat a piece of nichrome wire in a Bunsen flame until the flame is no longer coloured.
3. Dip the loop at the end of the wire into some water.
4. Dip the loop into an unknown salt.
5. Hold the wire in the edge of the flame.
6. Record the colour and identify the cation using the table below.

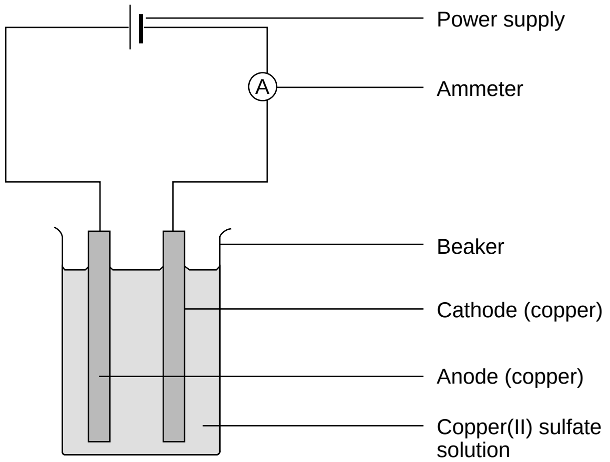
Metal	Colour of flame
Barium	Apple-green
Calcium	Brick-red
Copper	Green with blue streaks
Lithium	Crimson
Potassium	Lilac
Sodium	Yellow

RS•C

Quantitative electrolysis

Introduction

When electrolysis is done on a commercial scale it is important to know how much current is required and for how long. This experiment relates the amount of metal removed from an electrode to the electric current and the time the current flows.



What to record

- The masses of the electrodes before electrolysis (identify the electrodes by writing on them).
- The masses of the electrodes after electrolysis.
- The current flowing.
- The time the current flows.

What to do

1. Clean the electrodes with emery paper (avoid inhaling any dust).
2. Weigh the anode.
3. Immerse the electrodes to a depth of 3–4 cm in the solution of copper(II) sulfate.
4. Allow about 0.4 A to pass for about 30 min.
5. Remove the anode, wash carefully in water and dry gently with a paper towel.
6. Reweigh the anode.

Safety

Wear eye protection.

Questions

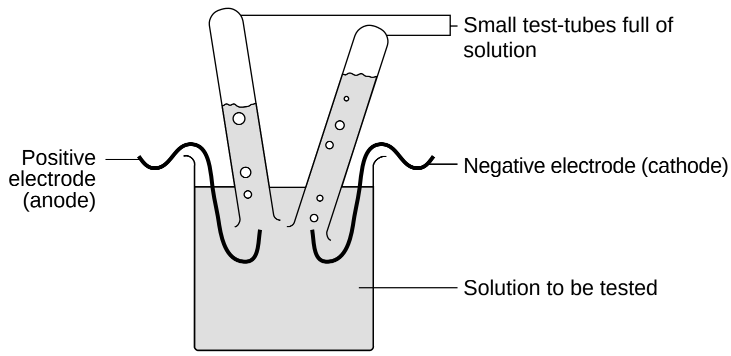
1. Calculate the number of moles of copper that have been removed from the anode.(Mass lost in g / 63.5)
2. Calculate the charge that has flowed through the circuit using the relationship charge (in Coulombs) = current (in amps) x time (in seconds).
3. Using the answers to questions 1 and 2, calculate the number of Coulombs required to remove one mole of copper.
4. 193,000 (2 x 96,500) Coulombs is required to remove one mole of copper. The difference between this and the answer to question 3 is due to errors in the experiment. What are the main sources of error in this experiment?

RS•C

The electrolysis of solutions

Introduction

When electricity passes through molten compounds, like sodium chloride, the ions move towards the electrode of opposite charge. Sodium chloride gives sodium metal and chlorine gas. This experiment illustrates what happens when the system is made more complicated because water is present. Electricity is passed through various solutions and the products are identified.



What to record

Solution	Product at the anode	Product at the cathode

What to do

1. Set up the apparatus as shown.
2. Switch on and observe what happens.
3. Try to identify the gases produced (if any).

Safety

Wear eye protection

The gases produced may be flammable, oxidising, and toxic. Take care not to inhale them. Do not let the current flow for very long. Some of the solutions are toxic.

Questions

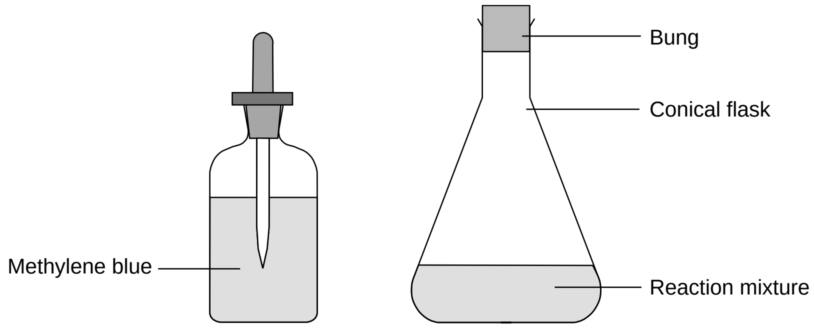
1. What type of element is formed at the negative electrode?
2. What type of element is formed at the positive electrode?
3. Your table of results should show some products, which could not come from the compound itself that was electrolysed. Where could these other products have come from?
4. Write a general rule for the products formed at
   (*a*) the cathode
   (*b*) the anode.

RS•C

An oxidation and reduction reaction

Introduction

A conical flask contains a colourless solution. When shaken, a blue colour forms. After a few seconds, the blue colour fades and the solution again becomes colourless. The process can be repeated. It is an oxidation followed by a reduction process



What to do

1. Put some water in the conical flask. Put in the stopper. Shake vigorously to check for leaks. If there are none, pour the water away and proceed.
2. Put 100 cm^3 of potassium hydroxide solution into a conical flask.
3. Add 3.3 g dextrose.
4. Add 3–4 drops of methylene blue indicator.
5. Put a stopper on the flask.
6. Shake vigorously.
7. When the solution clears, repeat the process.
8. It is necessary periodically to remove the stopper.

Safety

Wear eye protection.

Questions

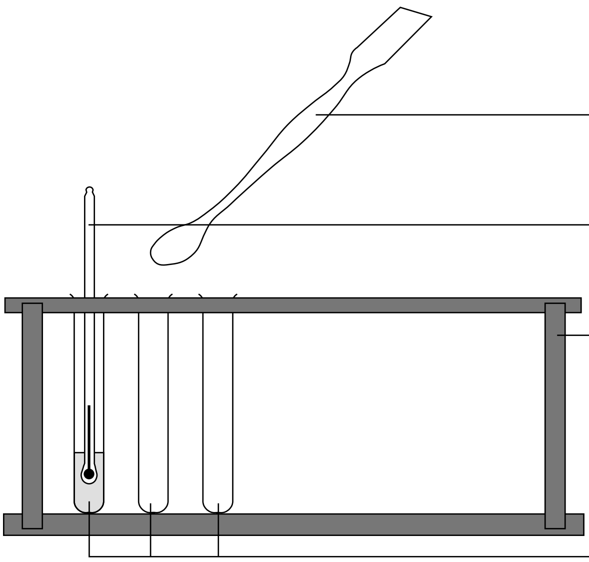
1. Why is it necessary to periodically remove the stopper?

RS•C

Heats of reaction (exothermic or endothermic reactions)

Introduction

Instant hot and cold packs are available for use in first aid. This experiment illustrates the types of chemical reaction that occur in these packs.



What to record

What was done and any changes in temperature from the starting temperature of your reaction. A table may be useful.

Initial solution	Temperature of solution/°C	Solid added	Final temperature of mixture/°C	Temperature change	Type of reaction

What to do

Experiment 1.
1. Put 2 cm^3 of water in a test-tube.
2. Record the temperature of the water.
3. Add a spatula measure of anhydrous (white) copper (II) sulfate.
4. Carefully stir, using the thermometer, and record the temperature again.

Experiment 2.
1. In a dry test-tube mix one spatula measure of citric acid and one spatula measure of sodium hydrogencarbonate.
2. Put 2 cm^3 of water in another test tube.

RS•C

3. Record the temperature of the water.
4. Add the mixture to the water.
5. Watch what happens and take the temperature of the solution.

Experiment 3.

1. Put about 5 cm^3 of copper(II) sulfate solution in a test-tube.
2. Record the temperature.
3. Add a spatula measure of powdered zinc.
4. Record the new temperature.

Safety

Wear eye protection.

Anhydrous copper(II) sulfate is harmful.

Zinc powder is flammable.

Questions

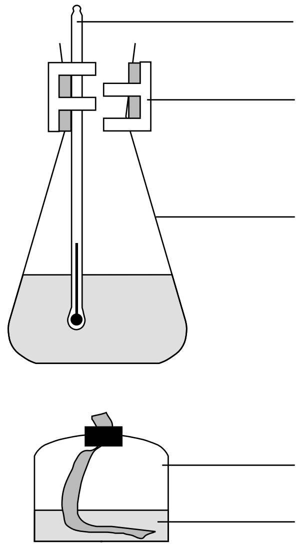
1. Which reactions are exothermic and which are endothermic?
2. Write word and symbol equations for Experiment 3.
3. Which two substances could be put in a cold pack?
4. Golfers need a hand warmer to keep their hands warm on a cold day. Which chemicals could be put in these warmers?

RS•C

Comparing the heat energy produced by combustion of various alcohols

Introduction

The combustion of alcohol produces energy. This experiment compares the amount of heat produced by the combustion of various alcohols.



What to record

Alcohol	Initial temp/°C	Final temp/°C	Temp change/°C	Initial mass/g	Final mass/g	Mass used/g
Methanol						
Ethanol						
Propanol						
Butanol						

What to do

1. Fill the conical flask with 100 cm^3 of water. Clamp the flask at a suitable height so that the spirit burner can be easily placed below.
2. Weigh the spirit burner (and lid) containing the alcohol and record the mass and name of the alcohol.
3. Record the initial temperature of the water using the thermometer.
4. Place the spirit burner under the conical flask and light the wick.

RS•C

5. Allow the alcohol to heat the water so the temperature rises by about 40 °C.
6. Replace the cap to extinguish the flame.
7. Reweigh the spirit burner and cap and work out the mass of alcohol used.

Repeat for different alcohols. Use 100 cm^3 of new cold water each time.

Safety

Wear eye protection. Do not open the spirit burner.

Question

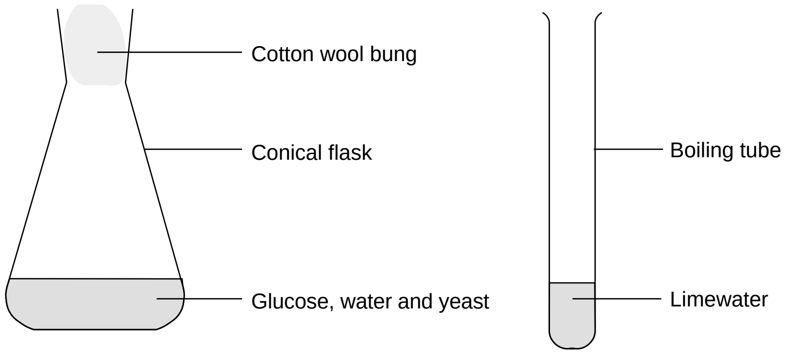
1. Which fuel provides the most energy per gram?

RS•C

Fermentation

Introduction

Beer and wine are produced by the fermentation of glucose by yeast. In this experiment, a glucose solution is left to ferment. The resulting mixture is then tested for the presence of any fermentation products.



What to record

What was seen to happen to the limewater.

What to do

Lesson 1

1. Put 5 g of glucose in the flask and dissolve it in 50 cm^3 of warm water.
2. Add 1 g of yeast to the glucose solution and plug the top with cotton wool.
3. Wait 20 min while fermentation takes place.
4. Remove the cotton wool and pour the invisible gas from the flask into the boiling tube with limewater. Take care not to pour any of the liquid as well.
5. Gently swirl the limewater round the tube and observe what happens.

Lesson 2

1. Note the smell of the solution.

Safety

Wear eye protection. Do not taste the product.

Questions

1. What gas is present in the flask after fermentation?
2. Suggest a different method for measuring the speed of this reaction.
3. Yeast contains a chemical called zymase, which is an enzyme. Complete the word equation for fermentation.

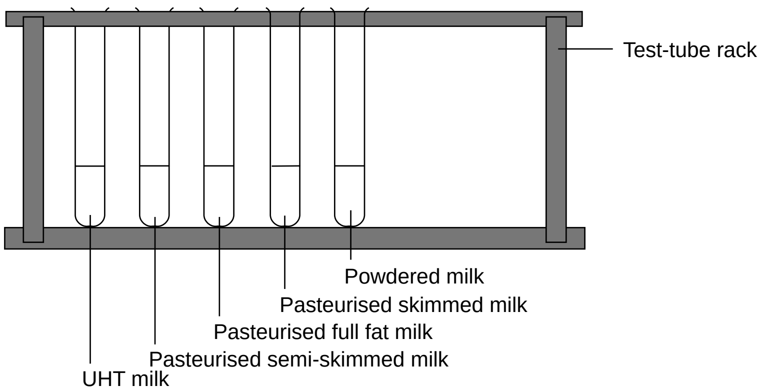
$$\text{Glucose} \xrightarrow{\text{zymase}} \quad\quad + $$

RS•C

Microbes, milk and enzymes

Introduction

Microbes are responsible for the production of some foods, for example cheese and yoghurt. They are also responsible for food decay. The enzymes they contain are catalysts. This experiment shows how these microbes and enzymes effect various types of milk.



What to do

The experiment needs to be done over two lessons.

First lesson.

1. Place five test-tubes in a test-tube rack. Fill each one to a depth of about 3 cm with five different types of milk.
2. Label each type with a sticky label near the top of the test-tube.
3. Leave in a warm place for between 3-5 days.

Second lesson.

1. Fill a beaker with about 100 cm^3 of tap water and stand the test-tubes in the beaker. Heat over a Bunsen burner to about 60 °C.
2. Turn off the Bunsen burner and carefully lift the beaker off the tripod.
3. Put six drops of rezasurin indicator in each test-tube of milk, shake thoroughly with the normal side to side action.
4. Leave to stand for 15 min and note any colour change.

Key

The rezasurin indicates the number of bacteria present.

Colour	Number of bacteria	Condition of milk
Purple	None	Completely sterilised
Blue	Few	Milk still fresh
Pink	Some	Milk on the turn
Colourless	Many	Milk has gone off

Safety

Do not taste any of the milk.

Safety

Wear eye protection.

The transition metal compounds may be harmful or irritant, as may their solutions, depending on the concentration.

Ammonia vapour irritates eyes, lungs and the respiratory system

Answers

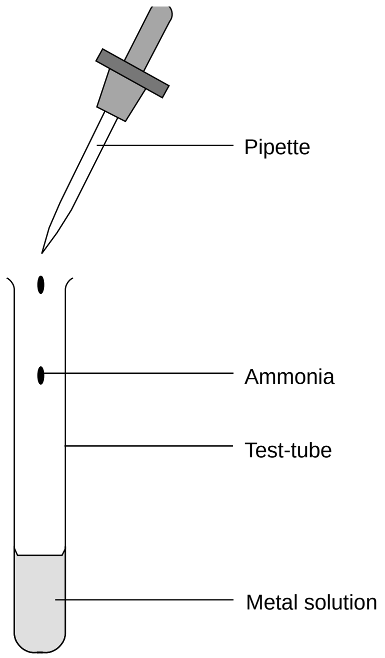
1. They are hard, dense and shiny. They are good conductors of heat and electricity. They are also malleable and ductile.
2. Transition metals react with water very slowly, if at all.
3. As well as the above they also form coloured compounds. They form compounds that can have more than one formula.

RS•C

Properties of the transition metals and their compounds

Introduction

The Periodic Table allows chemists to see similarities and trends in the properties of chemical elements. This experiment illustrates some properties of the common transition elements and their compounds.



What to record

What was observed. A table may be useful.

What to do

1. Test the metal samples for hardness and ability to bend without breaking. Complicated apparatus is not needed for this! Record your answers qualitatively.
2. Find out which samples are magnetic.
3. Set up an experiment to see if the metals react with water. (This may need to be left for some time).
4. Take a small sample of a solution of copper(II) sulfate (approximately 2 cm^3), add ammonia solution to it a few drops at a time. Record your observations. Add ammonia solution until there is no further change.
5. Repeat with the other solutions of transition metal compounds.

Safety

Wear eye protection.

The transition compounds may be harmful or irritant.

Ammonia vapour irritates eyes, lungs and the respiratory system.

Questions

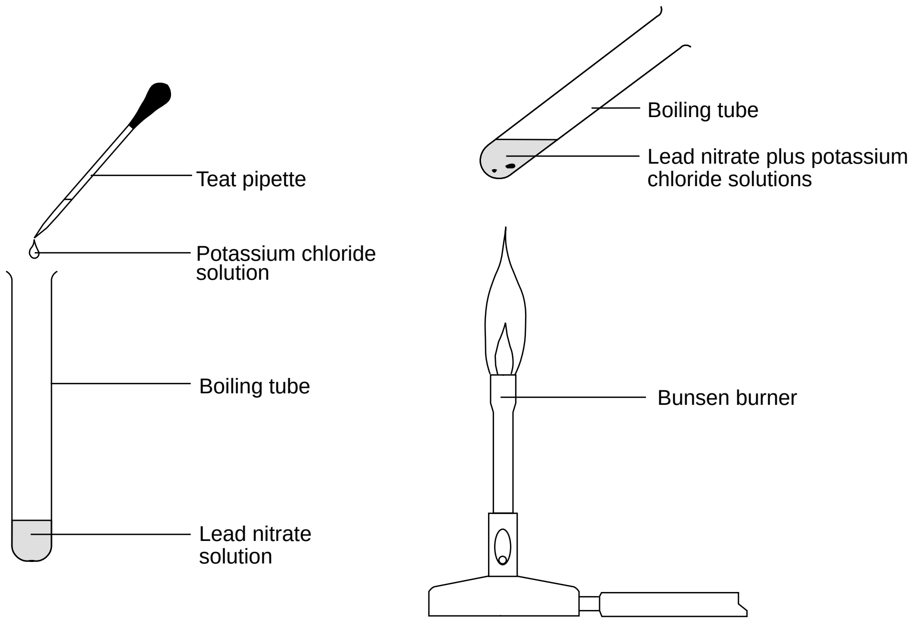
1. Describe the physical properties of transition metals.
2. How do transition metals react with water?
3. What properties do the compounds of transition metals have in common?

RS•C

Halogen compounds

Introduction

The halogens are elements of Group 7 of the Periodic table. This experiment illustrates some of the trends and similarities within the compounds of this group.



Safety

Wear eye protection.

What to record

What was observed. The table may be used.

What to do

Test 1 Silver nitrate	**Observations**
Put a little sodium chloride solution (Cl^- ions) in a test-tube and add five drops of silver nitrate solution. Leave in the light for a few minutes.	
Put a little sodium bromide solution (Br^- ions) in a test-tube and add five drops of silver nitrate solution. Leave in the light for a few minutes.	
Put a little sodium iodide solution (I^- ions) in a test-tube and add five drops of silver nitrate solution. Leave in the light for a few minutes.	

RS•C

Test 2 Chlorine water (Toxic vapour)	Observations
Add a few drops of a solution of chlorine (**Toxic vapour**) to sodium bromide solution.	
Add a few drops of a solution of chlorine (**Toxic vapour**) to sodium iodide solution.	

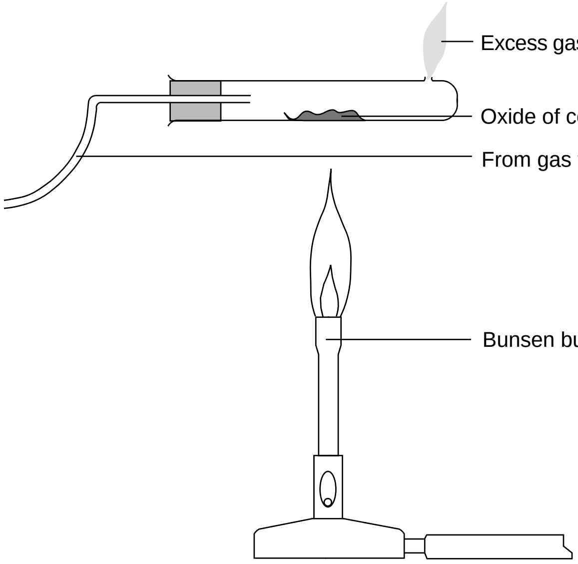
Test 3 Lead nitrate (Toxic)		Observations
Put approximately 4 cm^3 of lead nitrate solution (**Toxic**) into a boiling tube, then add five drops of potassium chloride solution to a boiling tube and heat till it boils.	Place all three tubes in a beaker of cold water to cool.	
Put approximately 4 cm^3 of lead nitrate solution (**Toxic**) into a boiling tube, then add five drops of potassium bromide solution to the boiling tube and heat till it boils.		
Put about 4 cm^3 of lead nitrate solution (**Toxic**) into a boiling tube, then add five drops of potassium iodide solution to a boiling tube and heat till it boils.		

RS•C

Finding the formula of an oxide of copper

Introduction

The chemical formula gives the types of atom in the substance. It also gives the relative number of each type. From the mass of each element in a sample, the number of moles can be calculated. The lowest whole number ratio provides the simplest chemical formula.



What to record

- ▼ Weight of test-tube + bung.
- ▼ Weight of test-tube + bung + copper(II) oxide.
- ▼ Weight of test-tube + bung + copper.

What to do

1. Weigh the test-tube + bung.
2. Place two spatulas of dry black copper(II) oxide in the centre of the tube. Try to spread it out.
3. Weigh the tube + bung + copper(II) oxide.
4. Assemble the apparatus as shown in the diagram.
5. Pass a gentle stream of gas through the tube without lighting it. This will flush out the air. After a few seconds set light to the gas and adjust the height of the flame coming out of the test tube to about 3 cm. Keep your head well back.
6. Heat the copper(II) oxide strongly and move the flame slowly to and fro. Continue to heat for five min after the solid has turned a brownish pink colour.
7. Stop heating the tube but keep the gas flowing through the test-tube and burning at the end. This prevents re-oxidation of the copper.
8. Let the test-tube cool, turn off the gas and reweigh the tube + bung + copper.

Safety

Wear eye protection

Copper(II) oxide is harmful.

Questions

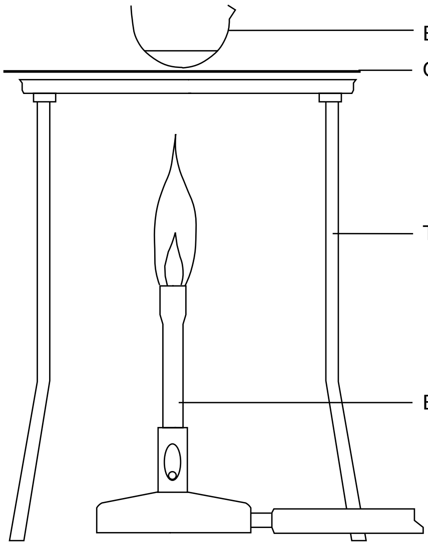
1. What is the mass of copper(II) oxide used?
2. What is the mass of copper formed?
3. What is the mass of oxygen lost?
4. How many moles of copper were formed?
5. How many moles of oxygen were combined with this number of moles of copper?
6. What is the simplest whole number ratio of moles of copper to moles of oxygen?
7. What is the formula of copper(II) oxide?

RS•C

Making a fertiliser

Introduction

Producting fertilisers is very important. This experiment involves preparing ammonium sulfate. Ammonium sulfate is a popular and effective fertiliser.



What to record

What was done and what was observed.

What to do

1. Put 20 cm^3 sulfuric acid into an evaporating basin.
2. Add the ammonia solution a little at a time, with stirring, until a definite smell of ammonia is obtained.
3. Check the pH is 7 or above with indicator paper.
4. Evaporate the solution to about one-fifth of its original volume (**Care** – do not let the solution spit), and cool.
5. Filter off the crystals and dry.

Safety

Wear eye protection.

The ammonia solution gives off ammonia which irritates eyes, lungs and respiratory system.

Sulfuric acid causes burns.

Questions

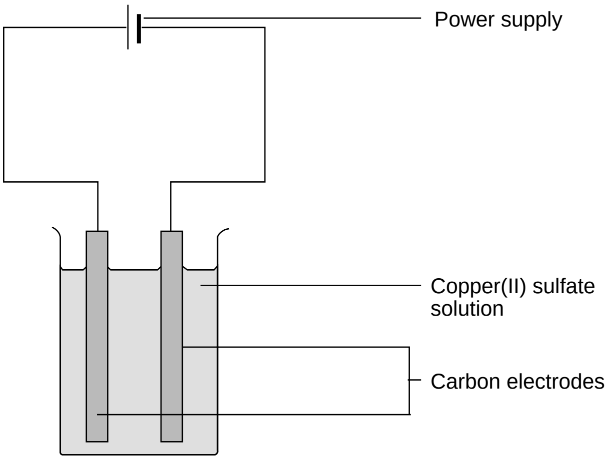
1. Write a word equation for this preparation.
2. Write a balanced symbol equation for this preparation.
3. Calculate the percentage of nitrogen in this fertiliser.

RS•C

Electrolysing copper(II) sulfate solution

Introduction

When a solution contains positive and negative ions, it conducts electricity. When electricity passes through a solution, chemical reactions may occur at the electrodes. This process is called electrolysis.



What to record

Which electrode has a fresh deposit of copper?

What to do

1. Set up the circuit shown in the diagram. Do not switch on the power supply until everything is ready.
2. Switch on and leave for about 5 min. Watch what happens and record your results.
3. Exchange the copper plated electrode for a key or key ring, or other article of your choice. Leave for 5 min.
4. Attempt to remove the copper plate by reversing the polarity.

Safety

Wear eye protection.

Copper(II) sulfate is harmful.

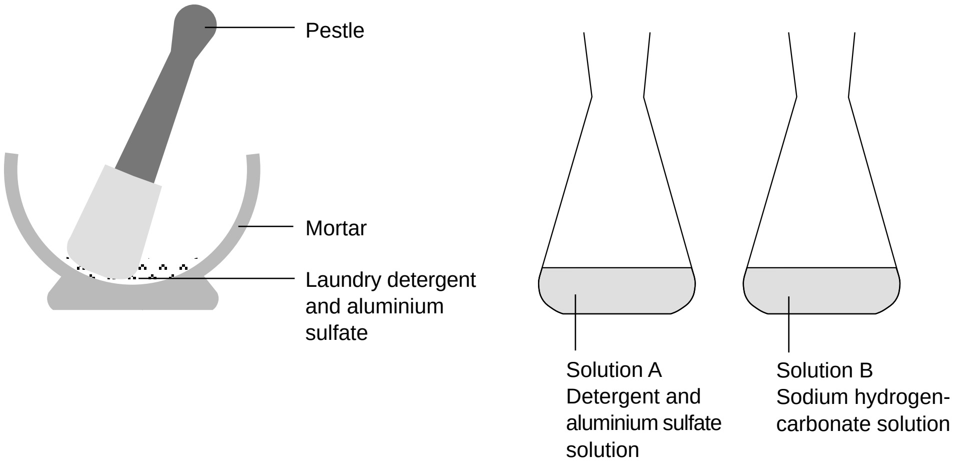
Questions

1. To which electrode has the copper gone?
2. What does this indicate about the charge on the copper in solution?
3. Suggest a way to use this process to purify impure or scrap copper.

Producing a Foam

Introduction

This experiment examines how a stable foam is produced and what reactions are involved.



What to record

What happened (observations).

What to do

1. Put 1 g of laundry detergent and 7 g of aluminium sulfate in a mortar and grind into a fine powder with a pestle.
2. Dissolve this powder in approximately 50 cm^3 of water in a conical flask (A).
3. Dissolve 5 g of sodium hydrogen carbonate in 50 cm^3 water in another conical flask (B).
4. Pour the contents of flask A into flask B and mix quickly.

Safety

Wear eye protection.

Detergent may irritate the skin.

Questions

1. What reactions lead to the production of the foam?
2. How is this reaction similar to that involving the production of carbon dioxide (CO_2) during the baking process?
3. Name some other examples of foams.

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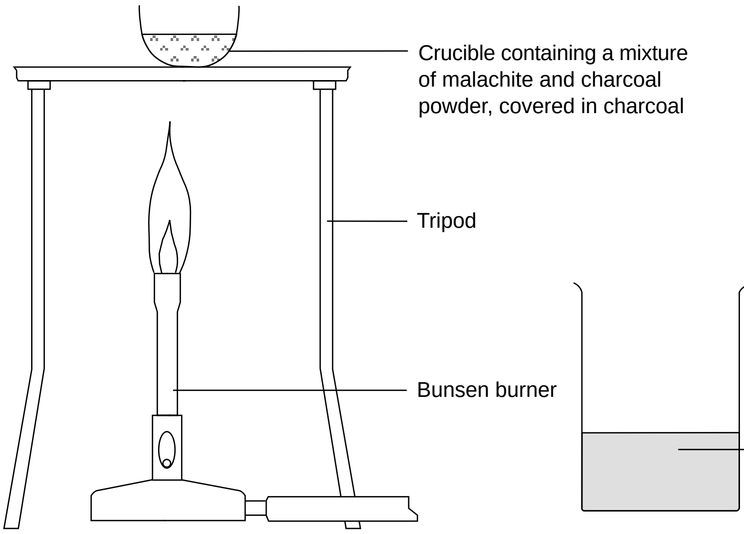
Getting metals from rocks

Introduction

This experiment involves producing copper from copper ore (malachite). The composition of malachite is mainly copper carbonate ($CuCO_3$). This experiment involves heating the copper carbonate with carbon.



What to do

1. Mix a spatula of crushed malachite with two spatulas of charcoal powder. (This can be done in a dry test-tube or on a piece of paper).
2. Make a pile of the mixture in a crucible and then cover the pile with a layer of charcoal in small pieces.
3. Put the crucible on a tripod and pipe clay triangle and heat very strongly. (**Beware sparks**)
4. Half fill the beaker with water and then use tongs to tip the powder from the crucible into the water.
5. Swirl the beaker round so any copper falls to the bottom and then pour off the water and charcoal.
6. Add more water and keep on pouring and swirling so only the heavy material is left at the bottom of the beaker.

Safety

Wear eye protection.

Copper compounds are harmful.

Beware sparks.

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Questions

1. Malachite ore is mainly copper carbonate ($CuCO_3$). When heated it produces carbon dioxide.
   Write word and symbol equations for the decomposition of copper carbonate.
2. Carbon reduces copper oxide to copper and carbon dioxide.
   Write word and symbol equations for this reaction.
3. An alternative method of extracting copper from malachite is by reaction with sulfuric acid, followed by electrolysis.
   Write formula equations for this reaction scheme.

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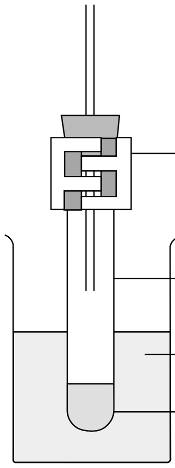
Addition polymerisation

Introduction

Alkenes (carbon compounds containing double bonds) undergo addition reactions. In this experiment molecules of phenylethene (styrene) – the monomer – add on to each other to form polyphenylethene (polystyrene) - the polymer.



What to record

What was done. Compare the appearance of your product with the starting material.

What to do

1. Prepare a beaker of boiling water to act as a water bath. Keep your Bunsen burner away from all other chemicals.
2. Add 0.1 g of di(dodecanoyl) peroxide (**Irritant**) to 5 cm^3 of phenylethene (**Flammable**) in a boiling tube.
3. Put a bung containing a 20 cm length of glass tubing in the top and clamp the tube in the boiling water bath.
4. Heat for about 30 min and leave to cool. Extinguish all flames.
5. Pour the contents of the tube into 50 cm^3 of ethanol (**Highly flammable**). Use a glass rod to push the polyphenylethene into a lump and pour off the ethanol.
6. Dry the solid on a filter paper.

Safety

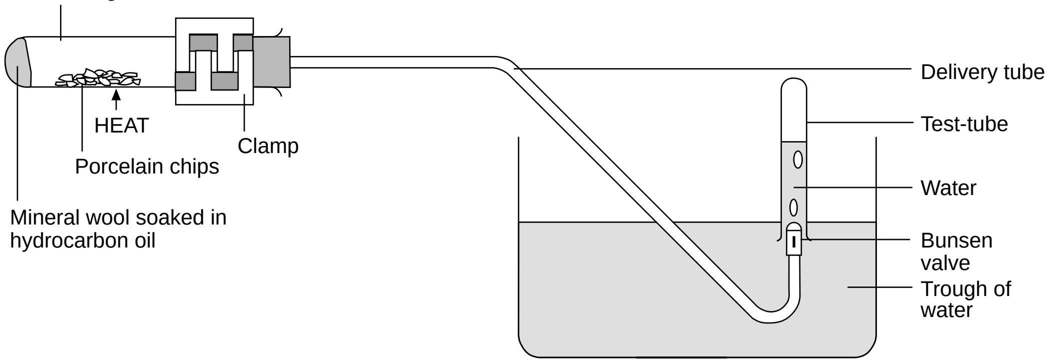
Wear eye protection. Work in a fume cupboard or ensure good ventilation.

RS•C

Cracking hydrocarbons

Introduction

The demand for petrol is greater than the amount produced by distilling crude oil. The cracking of hydrocarbons also produces molecules which can be converted into petrol. This experiment models the industrial cracking process.



What to record

Record:

1. The appearance of the original oil and the gas.
2. Whether the oil and gas burn.
3. Whether the oil and gas smell.
4. The effect of the oil and the gas on aqueous bromine water.
5. A table may be helpful.

What to do

1. Set up the apparatus as shown in the diagram. Fill four test-tubes with water and invert them in the trough.
2. Strongly heat the catalyst (porcelain chips) for a few minutes.
3. Now flick the flame onto the end of the tube containing the mineral wool and the oil. Try to produce a steady stream of bubbles.
4. Collect tubes of the gas. Discard the first one, which will be mainly air.
5. Stopper the three tubes of gas and test them as follows
   (*a*) What do they look like?
   (*b*) What do they smell like (care)?
   (c) Use a lighted splint to see if they burn.
   (*d*) Add two drops of aqueous bromine and shake.

Safety

Wear eye protection.

Be careful to avoid the water in the trough sucking back. When heating stops lift the apparatus out of the water using the stand.

Make sure the apparatus is not blocked. If no gas appears check the bung has not melted and that the catalyst has not blocked the delivery tube.

Aqueous bromine water is harmful and irritant.

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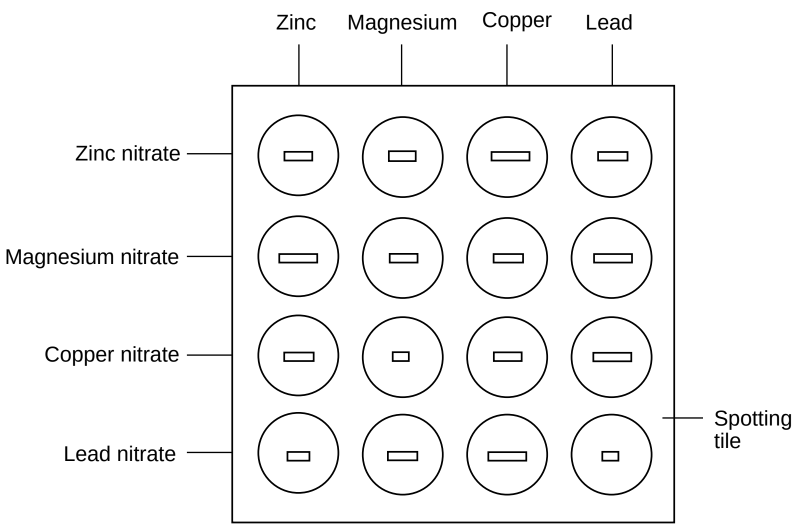
Displacement reactions between metals and their salts

Introduction

Some metals are more reactive than others. In this experiment, a strip of metal is added to a solution of another. If the metal is more reactive than the metal in solution, this metal displaces (pushes out) the less reactive metal from the solution.



What to record

Record which metals react with the solutions.

A table may be useful. Use a ✔ to show reactivity and an ✘ to show no reaction.

Solution / Metal	Zinc nitrate solution	Magnesium nitrate solution	Copper nitrate solution	Lead nitrate solution
Zinc				
Magnesium				
Copper				
Lead				

What to do

1. Clean each of the metal strips with emery paper.
2. Using a teat pipette put some of the solution of a metal compound in four of the holes in the spotting tile. (Label this row with the name of the solution).
3. Do this for each solution of a metal compound.
4. Put a piece of each metal in each of the solutions.
5. Put a tick or a cross in your table to show if they have reacted.

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Safety

Wear eye protection. Some solutions are toxic.

Questions

1. What is the order of reactivity of these metals (from most to least reactive)?

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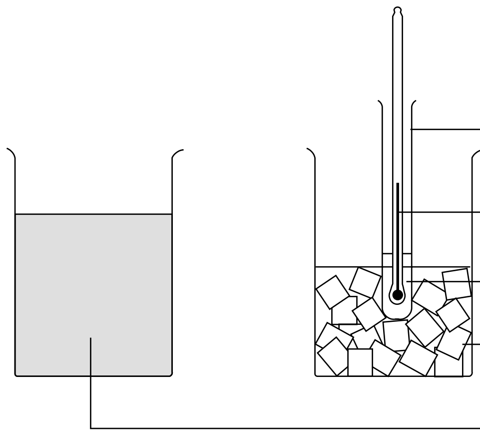
The effect of temperature on solubility

Introduction

Most solid substances that are soluble in water are more soluble in hot water than in cold water. This experiment examines solubility at various temperatures.



What to record

Fill in the temperatures.

Volume of water/cm^3	Solubility/g dm^3	Crystallisation temperature /° C
4	650	
5	520	
6	433	
7	371	
8	325	
9	289	
10	260	

(The crystallisation temperature is the temperature at which crystals appear).

What to do

1. Set up a hot water bath and an ice bath. Put 2.6 g of ammonium chloride (**Harmful**) into the boiling tube. Add 4 cm^3 water.
2. Warm the boiling tube in the hot water bath until the solid dissolves.
3. Put the boiling tube in the ice bath and stir with the thermometer. Use wooden tongs to hold it if necessary.
4. Note the temperature at which crystals first appear and record it in the table
5. Add 1 cm^3 water. Warm the solution again, stirring until all the crystals dissolve.

6. Then repeat the cooling and note the new temperature at which crystals appear.
7. Repeat steps 5, 6 and 7 until 10 cm^3 water has been used.

Safety

Wear eye protection.

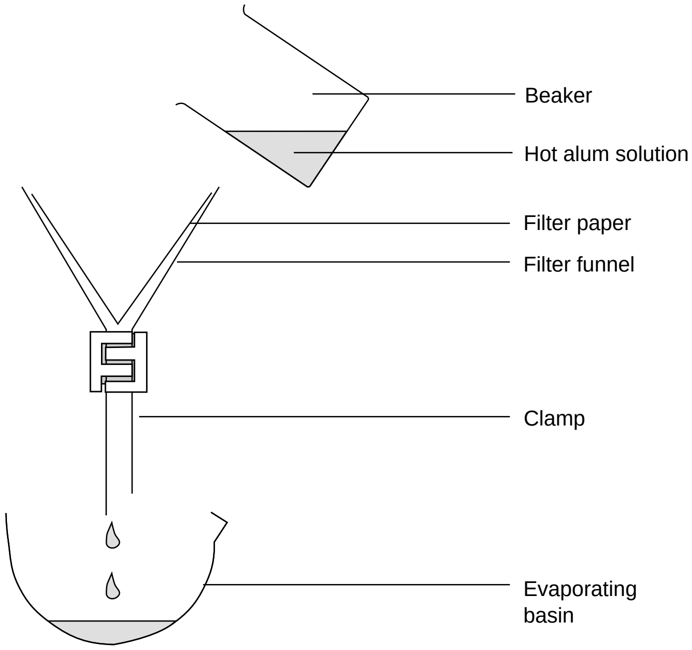
Questions

1. Plot a graph showing solubility on the vertical axis and temperature on the horizontal axis.

Purifying an impure solid

Introduction

It is often necessary to obtain a pure chemical from an impure sample. This experiment involves the purification of a chemical called alum.



What to record

What was done.

What to do

1. Cover the bottom of the beaker with alum.
2. Add enough water to cover it.
3. Heat and stir the mixture with a glass rod.
4. Wait until the mixture boils. If the alum has not dissolved, (ignore the dirt) add some more water and boil.
5. Filter the hot solution into the evaporating basin.
6. Leave the solution to cool.
7. Pour off the liquid and dry the crystals on filter paper.

Safety

Wear eye protection. Take care with hot apparatus.

Questions

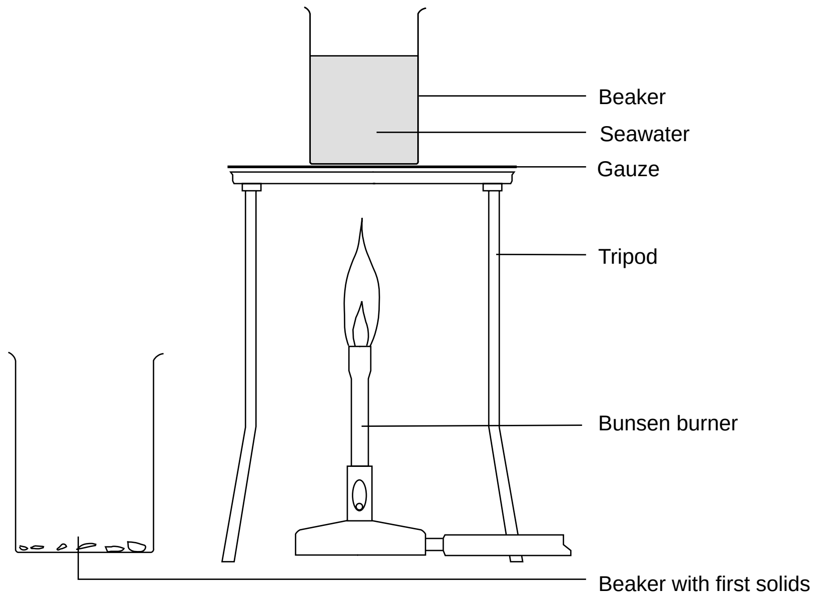
1. Write a couple of sentences to describe this experiment. Make sure all of the following words are used:
   dissolve, solvent, filter, filtrate, evaporate, solute, crystallise, crystal.

RS•C

Chemicals from seawater

Introduction

Seawater is often called salt water. Seawater contains various different salts. This experiment involves separating some of these salts from the mixture.



What to record

What was done and what was seen.

What to do

1. Put 200 cm^3 of seawater into a 250 cm^3 beaker.
2. Heat the seawater to boiling point and boil the liquid until 60–70 cm^3 remains. Stop heating when solid is observed and let the solid settle.
3. Pour (decant) the clear liquid into a 100 cm^3 beaker.
4. Add a few drops of hydrochloric acid to the solid left behind. What happens?
5. Put the beaker back on the heat and boil it again until another solid appears (probably when the liquid level is between 20 cm^3 and 40 cm^3.)
6. Filter the liquid off, wash out the beaker, and boil the liquid again until there is almost none left. Let it cool. What is observed?

Safety

Wear eye protection. Take care with hot apparatus and solutions.

Questions

1. Is seawater a single substance?
2. What is the difference between these separated salts?