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Centres are responsible for their own hazard analysis and risk assessment before beginning this practical work with pupils.

Chemistry

Unit AS1: Practical Manual

Practical 1.1

Determine the formula of a hydrated compound by weighing and heating a hydrated salt to constant mass (spec ref: 1.1.7)

Introduction

Salts which contain water of crystallisation and are known as hydrated salts. If the water of crystallisation is removed then the salts are known as anhydrous salts. The degree of hydration can be determined by heating a hydrated salt to constant mass. In this practical you will determine the degree of hydration in hydrated copper(II) sulfate, $CuSO_4.xH_2O$. This form of copper(II) sulfate is blue. On heating, the hydrated copper sulfate loses its water to become anhydrous copper(II) sulfate which is white. This process can be represented by the equation:

$$CuSO_4.xH_2O(s) \rightarrow CuSO_4(s) + xH_2O(g)$$

By weighing the copper(II) sulfate sample before and after strong heating it is possible to calculate the value of x.

Apparatus and materials

- safety goggles
- crucible and pipe-clay triangle
- spatula
- Bunsen burner
- heat mat
- tripod
- tongs
- hydrated copper(II) sulfate
- access to a mass balance (2 d.p)

Practical 1.1

Determine the formula of a hydrated compound by weighing and heating a hydrated salt to constant mass (spec ref: 1.1.7)

Procedure

1. Weigh and record the mass of the empty crucible.
2. Weigh between 3.5 g and 4.00 g of hydrated copper(II) sulfate into the crucible. Record the mass of the crucible and contents.
3. Set up a tripod, heat mat and Bunsen burner. Place a pipe-clay triangle over the tripod. Securely mount the crucible on top of the pipe-clay triangle.

4. Heat the crucible steadily for five minutes. Use tongs to remove the crucible onto a heat proof mat and leave to cool for a few minutes.
5. Once cooled, use tongs to transfer the crucible to the balance and record the total mass of the crucible and contents.
6. Place the crucible back onto the pipe-clay triangle and heat for a further five minutes. Leave to cool for a few minutes.
7. After cooling, transfer the crucible to the balance and record the mass of the crucible and contents.
8. Repeat steps 6 and 7 if the mass of the crucible has not yet reached a constant mass.

Mass of empty crucible / g	
Mass of crucible and contents / g	
Mass of copper(II) sulfate before heating / g	
Mass of crucible and contents after step 4 / g	
Mass of crucible and contents after step 6 / g	
Mass of crucible and contents after step 8 / g	
Mass of copper(II) sulfate after heating / g	
Mass of water of crystallisation / g	

Practical 1.1

Determine the formula of a hydrated compound by weighing and heating a hydrated salt to constant mass (spec ref: 1.1.7)

Results analysis

1. Calculate the relative formula masses of water and anhydrous copper(II) sulfate.

2. Calculate the number of moles of anhydrous copper(II) sulfate formed.

3. Calculate the number of moles of water of crystallisation lost by evaporation / heating.

4. Calculate how many moles of water would have been driven off if 1 mole of anhydrous copper(II) sulfate had been formed. This value represents the degree of hydration in hydrated copper(II) sulfate. Use your results to suggest the formula for hydrated copper(II) sulfate.

Questions

1. The formula of hydrated copper(II) sulfate is $CuSO_4.5H_2O$. Compare the formula with that calculated using the experimental results and suggests reasons for any difference.

Practical 2.1

Use the deflection of a stream of liquid from a burette to indicate polarity or lack of polarity within a molecule (spec ref: 1.3.9)

Introduction

Molecules are non-polar if they do not contain polar bonds. Non-polar molecules which are liquids will not be deflected by a charged rod. Some molecules which contain polar bonds are non-polar if the polar bonds are arranged symmetrically. Molecules which contain polar bonds which are not arranged symmetrically are polar. Polar molecules which are liquids will be deflected by a charged rod.

Apparatus and materials

- safety goggles
- 5 × burettes
- 5 × funnels
- 5 × 250 cm^3 beakers
- plastic rod / ruler
- cloth
- water
- propanone
- ethanol
- cyclohexane
- methylbenzene

Practical 2.1

Use the deflection of a stream of liquid from a burette to indicate polarity or lack of polarity within a molecule (spec ref: 1.3.9)

Procedure

1. Fill a burette with one of the liquids provided.
2. Place an empty 250 cm^3 beaker under the burette to collect the jet of liquid.
3. Charge the plastic rod/ruler by rubbing it vigorously with a piece of dry cloth.
4. Open the burette tap so that a jet of liquid flows into the beaker. Bring the plastic rod/ruler towards the jet but do not let them touch. Record which liquids are deflected and identify the liquid which is deflected the most.

Liquid	Deflection?	Polar molecule?
Water		
Propanone		
Ethanol		
Cyclohexane		
Methylbenzene		

Questions

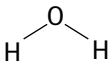
1. Which liquids deflected when the charged rod was brought near to them?

2. How do your results suggest that water is the most polar liquid out of the liquids that were deflected by the charged rod?

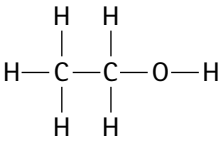
Practical 2.1

Use the deflection of a stream of liquid from a burette to indicate polarity or lack of polarity within a molecule (spec ref: 1.3.9)

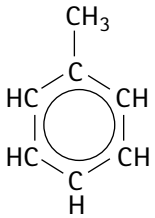
3. The structures of the five compounds are shown below. Mark the polarities on the polar bonds as appropriate.



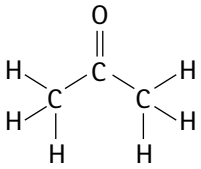
Water



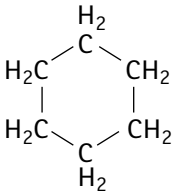
Ethanol



Methylbenzene

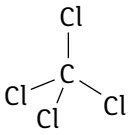


Propanone



Cyclohexane

4. Explain why, despite having polar bonds, carbon tetrachloride is non-polar molecule.



Practical 3.1

Carry out tests of electrical conductivity on solids and liquids and aqueous solutions of ionic and covalent substances (spec ref: 1.5.2)

Introduction

Substances will conduct electricity if they contain charge carriers, such as ions or electrons, that can move. In this practical you will investigate the conductivity of an ionic and covalent substance.

Apparatus and materials

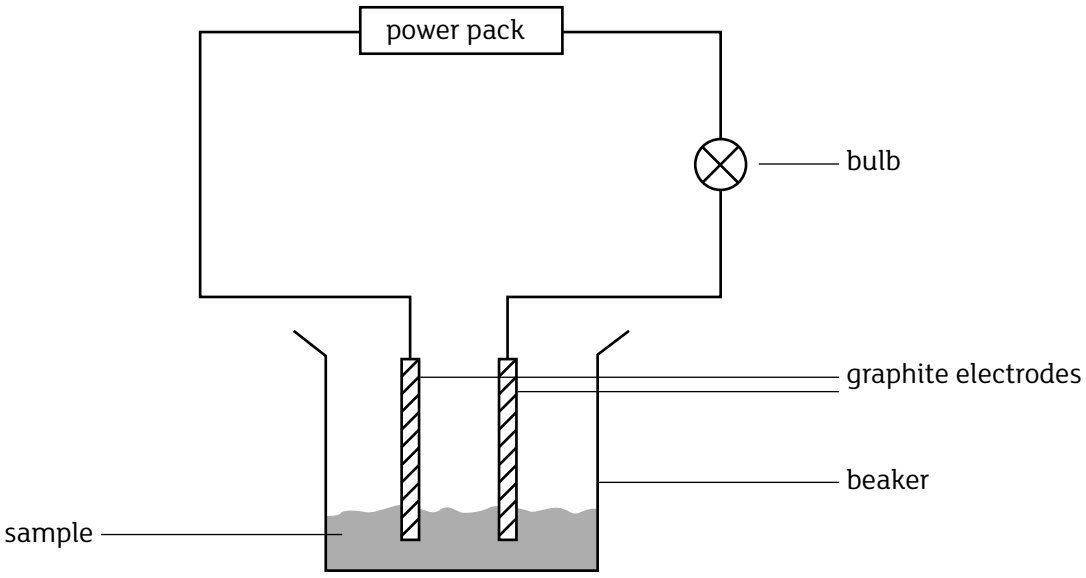
- safety goggles
- direct current 6 V power supply
- bulb
- 4 × crocodile clips
- 2 × graphite rods
- 3 × electrical wires
- 100 cm^3 beaker
- glass rod
- spatula
- 50 cm^3 measuring cylinder
- bottle of deionised water
- Bunsen burner
- tripod
- gauze
- heat mat
- evaporating basin
- sodium chloride
- sucrose

Practical 3.1

Carry out tests of electrical conductivity on solids and liquids and aqueous solutions of ionic and covalent substances (spec ref: 1.5.2)

Procedure

1. Place four spatula measures of sodium chloride into a 100 cm^3 beaker.
2. Set up a circuit as shown below.
3. Dip the electrodes into the sodium chloride, ensuring the electrodes do not touch each other.
4. Switch on the power pack and record observations.



5. Switch off the power pack and remove the electrodes from the beaker.
6. Measure 40 cm^3 of deionised water using a measuring cylinder and add to the beaker containing sodium chloride. Stir to dissolve.
7. Dip the electrodes into the aqueous solution of sodium chloride ensuring the electrodes do not touch and then switch on the power pack.
8. Repeat steps 1-6 using sucrose.

Substance	Observation with solid	Observation with aqueous solution
Sodium chloride		
Sucrose		

Practical 3.1

Carry out tests of electrical conductivity on solids and liquids and aqueous solutions of ionic and covalent substances (spec ref: 1.5.2)

Questions

1. Explain the observations made with sodium chloride and sucrose as solids and in aqueous solution.

Practical 4.1

Determine the solubility of chlorine and iodine in aqueous and non-aqueous solvents (spec ref: 1.8.2)

Introduction

Chlorine, bromine and iodine have low solubilities in water; they are much more soluble in non-polar solvents such as hexane. Iodine is the least soluble of the halogens in water; iodine solution is usually iodine dissolved in potassium iodide solution.

Apparatus and materials

- safety goggles
- 3 × test tubes and a test tube rack
- 3 × stoppers
- 4 × plastic dropping pipettes
- chlorine water
- bromine water
- iodine solution
- hexane

Procedure

1. Using plastic dropping pipettes, place 3 cm^3 of each aqueous halogen solution into separate test tubes and record the colour.
2. Using a plastic dropping pipette, place 3 cm^3 of hexane into each test tube.
3. Stopper each tube and shake the mixture by inverting the tube a few times.
4. Allow the two layers to settle; observe and record the colour of the upper layer.

Aqueous solution	Colour of aqueous solution	Colour of halogen in hexane
Chlorine		
Bromine		
Iodine		

Practical 4.1

Determine the solubility of chlorine and iodine in aqueous and non-aqueous solvents (spec ref: 1.8.2)

Questions

1. Explain, without using density values, how you would show that the aqueous layer is the bottom layer in each test tube.

Practical 4.2

Produce a reactivity order of the halogens using the displacement reactions of halogens with other halide ions in solution (spec ref: 1.8.5)

Introduction

The reactivity of the halogens decreases down the group. This can be demonstrated using displacement reactions. The colours of aqueous solutions of bromine and iodine are similar so the presence of iodine can be confirmed using a non-polar solvent such as hexane as iodine is violet/purple in non-polar solvents.

Apparatus and materials

- safety goggles
- 9 × test tubes and a test tube rack
- 6 × plastic dropping pipettes
- chlorine water
- bromine water
- iodine solution
- potassium chloride solution
- potassium bromide solution
- potassium iodide solution
- hexane
- bottled of deionised water

Procedure

1. Using plastic dropping pipettes, add 1 cm^3 of each aqueous halogen solution into separate test tubes.
2. Using a plastic dropping pipette, add 1 cm^3 of potassium chloride solution into each test tube and record any colour change.
3. Repeat the experiment with potassium bromide solution and then with potassium iodide solution. Add 1 cm^3 of hexane, using a plastic dropping pipette, to the test tubes when using potassium iodide solution.

Practical 4.2

Produce a reactivity order of the halogens using the displacement reactions of halogens with other halide ions in solution (spec ref: 1.8.5)

Halogen	Colour of aqueous halogen solution	Colour of solution after addition of KCl(aq)	Colour of solution after addition of KBr(aq)	Colour of solution after addition of KI(aq)	Colour of upper layer
Chlorine					
Bromine					
Iodine					

Questions

1. Write ionic equations for each reaction that occurs.

2. Explain, using oxidation numbers, why halogen displacement reactions are classed as redox reactions.

Practical 4.3

Carry out the reactions of the halides with concentrated sulfuric and phosphoric acids and perform chemical tests for the products (spec ref: 1.8.6)

Introduction

The halogens can act as oxidising agents. They will be themselves reduced, gaining electrons to form their corresponding halide ion. It follows that the halide ions can therefore act as reducing agents and can be oxidised back to the halogen. The halogens oxidising ability decreases down the group so the converse must be true: the halides reducing ability increases down the group. This can be demonstrated using concentrated sulfuric acid, which can act as an oxidising agent with some halides. In addition, concentrated sulfuric acid and concentrated phosphoric acid react similarly with halides in a reaction which is not classed as a redox reaction.

Apparatus and materials

- safety goggles
- 3 × test tubes and a test tube rack
- plastic dropping pipette
- spatula
- glass rod
- filter paper
- concentrated sulfuric acid
- concentrated phosphoric acid
- potassium chloride
- potassium bromide
- potassium iodide
- concentrated ammonia solution
- acidified potassium dichromate(VI) solution
- universal indicator paper
- bottle of deionised water

Practical 4.3

Carry out the reactions of the halides with concentrated sulfuric and phosphoric acids and perform chemical tests for the products (spec ref: 1.8.6)

Procedure

1. Place a spatula measure of potassium chloride in a test tube.
2. Using a plastic dropping pipette, add a few drops of concentrated sulfuric acid and record observations.
3. Test the gas evolved with damp universal indicator paper and a rod dipped in concentrated ammonia solution.
4. Repeat the experiment with potassium bromide. In addition to the test described in 3, also test the gases evolved using a piece of filter paper dipped in acidified potassium dichromate(VI) solution.
5. Repeat the experiment with potassium iodide, testing for gases using the tests described in 3 and 4.
6. Repeat the experiment using concentrated phosphoric acid instead of concentrated sulfuric acid. Test the gas evolved in each case with damp universal indicator paper and a rod dipped in concentrated ammonia solution.

Halide	Observations with concentrated sulfuric acid	Observations with concentrated phosphoric acid
Potassium chloride		
Potassium bromide		
Potassium iodide		

Practical 4.3

Carry out the reactions of the halides with concentrated sulfuric and phosphoric acids and perform chemical tests for the products (spec ref: 1.8.6)

Questions

1. Explain, for each salt, the observations made when concentrated sulfuric acid is added.

2. Write equations for the reaction of each salt with concentrated sulfuric acid which causes formation of steamy/white fumes. Explain why this reaction is not classed as a redox reaction.

Practical 4.3

Carry out the reactions of the halides with concentrated sulfuric and phosphoric acids and perform chemical tests for the products (spec ref: 1.8.6)

3. Write ionic redox equations for the reactions of concentrated sulfuric acid with potassium bromide and potassium iodide. Explain, using oxidation numbers, how the observations made demonstrates that iodide ions are a stronger reducing agent than bromide ions.

4. Write equations for the reaction of each salt with concentrated phosphoric acid which causes formation of steamy/white fumes. Explain why this reaction is not classed as a redox reaction.

5. Explain why redox products are not obtained when each salt reacts with concentrated phosphoric acid.

Practical 5.1

Prepare solutions of known concentration
(spec ref: 1.9.8)

Introduction

A solution of known concentration is known as a standard solution. In this practical you will prepare a standard solution of hydrochloric acid, by dilution, for use in the titration practical 6.1.

Apparatus and materials

- safety goggles
- hydrochloric acid
- graduated pipette
- pipette filler
- 250 cm^3 volumetric flask & stopper
- bottle of deionised water
- 2 × 250 cm^3 beakers
- 100 cm^3 beaker
- plastic dropping pipette
- sticky label

Procedure

You have access to hydrochloric acid which has a known concentration of 2.0 mol dm^{-3}. You will use this to prepare 250 cm^3 of a solution of hydrochloric acid with a concentration of 0.1 mol dm^{-3}. This is done by diluting a known volume of the original hydrochloric acid solution and making it up to a volume of 250 cm^3.

Calculate what volume of the original hydrochloric acid solution you will need to prepare the diluted solution.

Practical 5.1

Prepare solutions of known concentration (spec ref: 1.9.8)

1. Using a pipette filler, rinse the pipette with hydrochloric acid and then fill the pipette with the calculated volume of hydrochloric acid.
2. Transfer the solution into the 250 cm^3 volumetric flask. Make up to the graduation mark with deionised water.
3. Stopper the flask and invert it to mix the contents thoroughly.

Questions

1. Assuming the pipette has an error of +/- 0.10 cm^3, calculate the percentage error in your measurement.

2. Explain the purpose of inverting the volumetric flask.

Practical 5.2

Prepare solutions of known concentration
(spec ref: 1.9.8)

Introduction

Vinegar is a solution that consists mainly of ethanoic acid. It has a variety of uses in the food industry. The consistency of vinegar, like any product is very important and this is why during manufacture it undergoes batch testing. This ensures good quality control during the manufacturing process.

You will prepare a diluted vinegar solution from household white vinegar. You will determine the concentration of ethanoic acid present by titration in practical 6.2.

Apparatus and materials

- safety goggles
- access to white vinegar
- 25.0 cm^3 pipette
- pipette filler
- 250 cm^3 volumetric flask & stopper
- bottle of deionised water
- 100 cm^3 beaker
- plastic dropping pipette
- sticky label

Procedure

1. Using a pipette filler, rinse the pipette with white vinegar and then fill the pipette with 25.0 cm^3 of white vinegar.
2. Transfer the solution into the 250 cm^3 volumetric flask. Make up to the graduation mark with deionised water.
3. Stopper the flask and invert it to mix the contents thoroughly.

Practical 5.3

Prepare solutions of known concentration (spec ref: 1.9.8)

Introduction

A solution of known concentration is known as a standard solution. In this practical you will prepare a solution of sodium carbonate for use in the titration practical 6.3.

Apparatus and materials

- safety goggles
- weighing boat
- spatula
- glass rod
- funnel
- 250 cm^3 volumetric flask & stopper
- bottle of deionised water
- 100 cm^3 beaker
- plastic dropping pipette
- sticky label
- approximately 4.00 g hydrated sodium carbonate
- access to a mass balance (2 d.p)

Procedure

1. Weigh between 3.57 and 3.58 g of hydrated sodium carbonate in a weighing boat.
2. Dissolve the solid in a suitable volume of deionised water in a beaker, stirring with a glass rod. Rinse the weighing boat contents and glass rod into the beaker.
3. Transfer the contents of the beaker into a 250 cm^3 volumetric flask, using a funnel.
4. Rinse the beaker into the volumetric flask, using deionised water, and then rinse the funnel into the flask.
5. Remove the funnel and make the solution up to the mark with deionised water, using a pipette to ensure the meniscus is on the line.
6. Stopper the flask and invert it to mix the contents thoroughly.

Practical 5.3

Prepare solutions of known concentration (spec ref: 1.9.8)

Questions

1. Explain how the loss of solution is minimised when it is transferred to the volumetric flask.

Practical 6.1

Carry out an acid-base titration to determine the concentration of acid/base, the degree of hydration in a hydrated metal carbonate and the percentage of ethanoic acid in vinegar (spec ref: 1.9.2)

Introduction

A titration is a method of volumetric analysis. It involves reacting two solutions together in a highly controlled way and using an indicator to show when end point has been reached. One solution is placed in a burette and the other is placed in a conical flask. The indicator is added to the solution in the conical flask. The solution in the burette is then added to the solution in the conical flask. The indicator changes colour when the end point of the titration has been reached. At AS titrations will involve simple acid-base reactions used to determine the concentration (or volume of a solution) and titrations involving a hydrated metal carbonate to determine the degree of hydration.

Apparatus and materials

- safety goggles
- 50.0 cm^3 burette
- clamp and stand
- 25.0 cm^3 pipette
- pipette filler
- 2 × 250 cm^3 conical flasks
- 2 × 100 cm^3 beakers
- funnel
- bottle of deionised water
- plastic dropper
- white tile
- bottle of phenolphthalein indicator
- sodium hydroxide solution
- hydrochloric acid solution from practical 5.1

Practical 6.1

Carry out an acid-base titration to determine the concentration of acid/base, the degree of hydration in a hydrated metal carbonate and the percentage of ethanoic acid in vinegar (spec ref: 1.9.2)

Procedure

A standard solution of sodium hydroxide can be used to exactly determine the concentration of the hydrochloric acid solution prepared in Practical 5.1. These two chemicals react together according to the following equation:

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

You will be using the indicator phenolphthalein and in this titration the end point is when the indicator turns from pink (the alkali in the flask) to colourless (acid added from burette).

1. Rinse and fill the burette with hydrochloric acid.
2. Rinse the pipette with sodium hydroxide solution, using a pipette filler.
3. Transfer 25.0 cm^3 of sodium hydroxide solution to a conical flask using the pipette and pipette filler.
4. Add 3–5 drops of phenolphthalein indicator to the solution in the conical flask. Place a white tile underneath the conical flask. Note the colour of the indicator in the flask.
5. Add the hydrochloric acid from the burette, with swirling, to the conical flask until the indicator changes colour. Record the reading on the burette.
6. Repeat the titration to achieve two concordant results and calculate the mean titre.

	Initial burette reading /cm^3	Final burette reading /cm^3	Titre /cm^3
Rough			
1st accurate			
2nd accurate			

Mean titre: _____

Practical 6.1

Carry out an acid-base titration to determine the concentration of acid/base, the degree of hydration in a hydrated metal carbonate and the percentage of ethanoic acid in vinegar (spec ref: 1.9.2)

Results analysis

1. Calculate number of moles of sodium hydroxide present in the 25.0 cm^3 sample in the conical flask.

2. Use the reacting ratio from the balanced symbol equation, and your answer from 1, to calculate the number of moles of hydrochloric acid needed to neutralise all the sodium hydroxide in the flask.

3. Use your answer to 2 and your mean titre to calculate the molarity of the hydrochloric acid. Compare this value with what was calculated in practical 5.1.

4. Express the concentration of the hydrochloric acid in g dm^{-3}.

Questions

1. The burette has an error of +/- 0.05 cm^3. Calculate the percentage error in the mean titre.

Practical 6.2

Carry out an acid-base titration to determine the concentration of acid/base, the degree of hydration in a hydrated metal carbonate and the percentage of ethanoic acid in vinegar (spec ref: 1.9.2)

Introduction

In this practical you will titrate the solution of white vinegar you prepared in practical 5.2 against a standard solution of sodium hydroxide. This will allow you to calculate the concentration of the white vinegar solution. The equation for this reaction is:

$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$

Apparatus and materials

- safety goggles
- 50.0 cm^3 burette
- clamp and stand
- 25.0 cm^3 pipette
- pipette filler
- 2 × 250 cm^3 conical flasks
- 2 × 100 cm^3 beakers
- funnel
- bottle of deionised water
- plastic dropper
- white tile
- bottle of phenolphthalein indicator
- sodium hydroxide solution
- white vinegar solution from practical 5.2

Practical 6.2

Carry out an acid-base titration to determine the concentration of acid/base, the degree of hydration in a hydrated metal carbonate and the percentage of ethanoic acid in vinegar (spec ref: 1.9.2)

Procedure

1. Rinse and fill the burette with sodium hydroxide solution.
2. Rinse the pipette with white vinegar solution, using a pipette filler.
3. Transfer 25.0 cm^3 of white vinegar solution to a conical flask using the pipette and pipette filler.
4. Add 3–5 drops of phenolphthalein indicator to the solution in the conical flask. Place a white tile underneath the conical flask. Note the colour of the indicator in the flask.
5. Add the sodium hydroxide solution from the burette, with swirling, to the conical flask until the indicator changes colour. Record the reading on the burette.
6. Repeat the titration to achieve two concordant results and calculate the mean titre.

	Initial burette reading /cm^3	Final burette reading /cm^3	Titre /cm^3
Rough			
1st accurate			
2nd accurate			

Mean titre: _____

Practical 6.2

Carry out an acid-base titration to determine the concentration of acid/base, the degree of hydration in a hydrated metal carbonate and the percentage of ethanoic acid in vinegar (spec ref: 1.9.2)

Results analysis

1. Calculate the moles of sodium hydroxide required to neutralise the diluted vinegar in the conical flask.

2. Use the reacting ratio from the equation above, and your answer from 1, to calculate the concentration of ethanoic acid present in the diluted vinegar.

3. The diluted vinegar was made by pipetting 25.0 cm^3 of the original white vinegar solution into a 250 cm^3 volumetric flask and making it up to the mark with deionised water. Use your answer to 2 to calculate the concentration of ethanoic acid in the original vinegar in mol dm^{-3}.

4. Determine the relative formula mass of ethanoic acid, CH_3COOH.

5. Use your answers to 3 and 4 to calculate the mass of ethanoic acid that is present in 1000 cm^3 of the original vinegar solution.

6. Use your answer from 5 to calculate the mass of ethanoic acid in 100 cm^3 of the original vinegar solution. This corresponds to the percentage of ethanoic acid present in the original vinegar solution.

Practical 6.3

Carry out an acid-base titration to determine the concentration of acid/base, the degree of hydration in a hydrated metal carbonate and the percentage of ethanoic acid in vinegar (spec ref: 1.9.2)

Introduction

In this practical you will titrate a solution of sodium carbonate, prepared in practical 5.3, against a standard solution of hydrochloric acid. This will allow you to calculate the degree of hydration in hydrated sodium carbonate, $Na_2CO_3.xH_2O$. The equation for this reaction is:

$$Na_2CO_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$$

Apparatus and materials

- safety goggles
- 50.0 cm^3 burette
- clamp and stand
- 25.0 cm^3 pipette
- pipette filler
- 2 × 250 cm^3 conical flasks
- 2 × 100 cm^3 beakers
- funnel
- bottle of deionised water
- plastic dropper
- white tile
- bottle of methyl orange indicator
- hydrochloric acid
- sodium carbonate solution from practical 5.3

Practical 6.3

Carry out an acid-base titration to determine the concentration of acid/base, the degree of hydration in a hydrated metal carbonate and the percentage of ethanoic acid in vinegar (spec ref: 1.9.2)

Procedure

1. Rinse and fill the burette with hydrochloric acid.
2. Rinse the pipette with sodium carbonate solution, using a pipette filler.
3. Transfer 25.0 cm^3 of sodium carbonate solution to a conical flask using the pipette and pipette filler.
4. Add 3–5 drops of methyl orange indicator to the solution in the conical flask. Place a white tile underneath the conical flask. Note the colour of the indicator in the flask.
5. Add the hydrochloric acid from the burette, with swirling, to the conical flask until the indicator changes colour. Record the reading on the burette.
6. Repeat the titration to achieve two concordant results and calculate the mean titre.

	Initial burette reading /cm^3	Final burette reading /cm^3	Titre /cm^3
Rough			
1st accurate			
2nd accurate			

Mean titre: _____

Practical 6.3

Carry out an acid-base titration to determine the concentration of acid/base, the degree of hydration in a hydrated metal carbonate and the percentage of ethanoic acid in vinegar (spec ref: 1.9.2)

Results analysis

1. Calculate the moles of hydrochloric acid required to neutralise the sodium carbonate solution in the conical flask.

2. Use the reacting ratio from the equation above, and your answer from 1, to calculate the number of moles of sodium carbonate in 25.0 cm^3 of the solution in the conical flask.

3. Use your answer to 2 to calculate the number of moles of sodium carbonate in 250 cm^3 of solution in the volumetric flask.

4. Use the mass of hydrated sodium carbonate weighed in practical 5.3 and your answer to 3 to calculate the relative formula mass of hydrated sodium carbonate, $Na_2CO_3.xH_2O$.

5. Calculate the relative formula mass of anhydrous sodium carbonate, Na_2CO_3.

6. Use your answers to 4 and 5 to determine the value of x in the sample of hydrated sodium carbonate, $Na_2CO_3.xH_2O$.

Practical 7.1

Use chemical tests listed in 'Qualitative tests' to identify unknown substances (spec ref: 1.10)

Introduction

Qualitative tests can be used to identify elements, compounds and ions. You are expected to be able to describe the tests, including reagents required, the observations made and explain the chemistry behind each test.

Apparatus and materials

- safety goggles
- test tubes
- plastic dropping pipettes
- delivery tubes with bungs attached
- flask and thistle funnel
- water trough
- beehive shelf
- gas collection jar
- glass rod
- nichrome wire
- Bunsen burner
- spatula(s)
- wooden splint(s)
- bottle of deionised water
- sodium carbonate
- hydrochloric acid
- limewater
- magnesium ribbon
- manganese dioxide
- hydrogen peroxide
- chlorine water
- universal indicator paper
- concentrated hydrochloric acid
- concentrated ammonia solution

Practical 7.1

Use chemical tests listed in ‘Qualitative tests’ to identify unknown substances (spec ref: 1.10)

- lithium chloride
- sodium chloride
- potassium chloride
- calcium chloride
- barium chloride
- copper(II) chloride
- potassium bromide
- potassium iodide
- silver nitrate solution
- ammonia solution
- sodium sulfate
- barium chloride solution
- sodium hydrogencarbonate
- sodium hydroxide solution
- ammonium chloride
- iodine solution
- starch solution

Practical 7.1

Use chemical tests listed in 'Qualitative tests' to identify unknown substances (spec ref: 1.10)

Procedure – tests for gases

Gas	Description of test	Observations
Carbon dioxide	Add a spatula measure of sodium carbonate to a few cm^3 of hydrochloric acid in a test tube. Connect a bung and delivery tube and pass the gas into a test tube containing 1–2 cm^3 of limewater.	
Hydrogen	Add 1 cm of magnesium to a few cm^3 of hydrochloric acid in a test tube. Place a lighted splint at the mouth of the test tube.	
Oxygen	Add 2–3 spatula measures of manganese dioxide to a flask and connect a thistle funnel. Add hydrogen peroxide using a plastic dropping pipette and collect the gas evolved over water in a gas collection jar. Place a glowing splint into the gas collection jar.	
Chlorine	Place a piece of damp universal indicator paper at the mouth of a bottle of chlorine water.	
Ammonia	Dip a glass rod in concentrated hydrochloric acid. Place the glass rod near the mouth of a bottle of concentrated ammonia solution.	
Hydrogen chloride	Dip a glass rod in concentrated ammonia. Place the glass rod near the mouth of a bottle of concentrated hydrochloric acid.	

Practical 7.1

Use chemical tests listed in 'Qualitative tests' to identify unknown substances (spec ref: 1.10)

Procedure – test for metal ions

1. Dip a nichrome wire in concentrated hydrochloric acid.
2. Dip the wire into the solid metal chloride.
3. Hold the wire in the blue Bunsen flame and note the colour of the flame.

Metal chloride	Observations
Lithium chloride	
Sodium chloride	
Potassium chloride	
Calcium chloride	
Barium chloride	
Copper(II) chloride	

Practical 7.1

Use chemical tests listed in 'Qualitative tests' to identify unknown substances (spec ref: 1.10)

Procedure – test for halide ions

1. Dissolve a spatula measure of potassium chloride in 2–3 cm^3 of dilute nitric acid.
2. Add 1 cm^3 of silver nitrate solution and record observations.
3. Add 1 cm^3 of dilute ammonia solution and record observations.
4. Add excess dilute ammonia solution and record observations.
5. If the precipitate remains, repeat steps 2 & 3 using concentrated ammonia solution in place of dilute ammonia solution.
6. Repeat steps 1–5 using potassium bromide in place of potassium chloride.
7. Repeat steps 1–5 using potassium iodide in place of potassium chloride.

Potassium halide	Observations
Potassium chloride	
Potassium bromide	
Potassium iodide	

Procedure – test for sulfate ions

1. Dissolve a spatula measure of sodium sulfate in 2–3 cm^3 of deionised water.
2. Add 1–2 cm^3 of barium chloride solution and record observations.

	Observations
Addition of barium chloride solution	

Practical 7.1

Use chemical tests listed in 'Qualitative tests' to identify unknown substances (spec ref: 1.10)

Procedure – test for carbonate/hydrogencarbonate ions

1. Add 1–2 cm^3 of dilute nitric acid to a test tube.
2. Add a spatula measure of sodium hydrogencarbonate to the test tube. Bubble any gas produced through limewater.

	Observations
Addition of sodium hydrogencarbonate to dilute nitric acid	

Procedure – test for ammonium ions

1. Add 1–2 cm^3 of sodium hydroxide solution to a test tube.
2. Add a spatula measure of ammonium chloride to the test tube.
3. Warm the test tube gently and test any gas evolved with damp universal indicator paper and a glass rod dipped in concentrated hydrochloric acid. Record observations.

	Observations
Addition of ammonium chloride to sodium hydroxide solution and heat	

Procedure – test for iodine

1. Add 1–2 cm^3 of iodine solution to a test tube.
2. Add a few drops of starch solution and shake gently.

	Observations
Addition of starch solution to iodine solution	

Practical 7.1

Use chemical tests listed in 'Qualitative tests' to identify unknown substances (spec ref: 1.10)

Questions

1. Write equations for the formation of carbon dioxide, hydrogen and oxygen in the tests performed on page 35.

2. Write equations for the observations made with carbon dioxide, hydrogen, chlorine, ammonia and hydrogen chloride.

3. Explain the use of nichrome wire and concentrated hydrochloric acid in a flame test.

4. Write equations, including state symbols, for the formation of the halide precipitates.

Practical 7.1

Use chemical tests listed in 'Qualitative tests' to identify unknown substances (spec ref: 1.10)

5. Write ionic equations, including state symbols, for the formation of the halide precipitates.

6. Write an ionic equation, including state symbols, for the equation between barium chloride solution and sodium sulfate solution.

7. Write an equation, including state symbols, for the reaction between sodium hydrogencarbonate and dilute nitric acid.

8. Write an equation, including state symbols, for the reaction between ammonium chloride and sodium hydroxide solution.