AH Exercise 3.

Calculations

2001 AH L2

2. Meat fillings in pies can be preserved by addition of sodium sulphite.

The concentration of the sodium sulphite can be determined using the method outlined in steps 1, 2 and 3.

Step 1 Addition of hydrochloric acid to displace sulphur dioxide.

$$Na_2SO_3 + 2HCl \rightarrow 2NaCl + H_2O + SO_2$$

Step 2 Absorption of the sulphur dioxide in a solution of sodium hydroxide to form sodium sulphite solution.

$$SO_2 + 2NaOH \rightarrow Na_2SO_3 + H_2O$$

Step 3 Titration of the sulphite ions against a standard solution of iodine.

$$SO_3^{2-} + H_2O + I_2 \rightarrow 2I^- + 2H^+ + SO_4^{2-}$$

A 10·0 g sample of meat filling, treated as above, released enough sulphur dioxide to react with 15·2 cm^3 of 0·0020 mol l^{-1} iodine solution.

(*a*) Suggest a reason why steps 1 and 2 are carried out rather than adding the meat filling to water and titrating directly with iodine solution. **1**

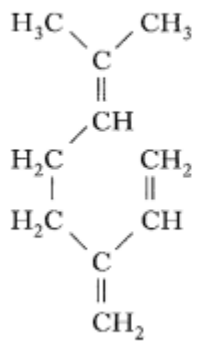
(*b*) Which indicator would be used in the titration? **1**

(*c*) (i) Calculate the number of moles of iodine used in the titration. **1**

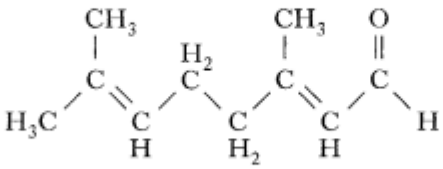
(ii) Calculate the % by mass of sodium sulphite in the sample of meat filling. **2**

2002 AH L12b

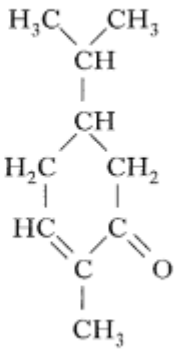
12. Myrcene, citral and carvone belong to a large group of compounds known as terpenes.



myrcene



citral



carvone

(*b*) A 0·019 mol sample of one of the above terpenes required 47·5 cm^3 of 1·20 mol l^{-1} bromine solution for complete reaction.

Identify, by calculation, the terpene used in the reaction. **Show your working.** **2**

2003 AH L2

2. In an experiment 0·25 g of an impure sample of magnesium carbonate was added to 40 cm^3 of 0·16 mol l^{-1} hydrochloric acid.

8·1 cm^3 of 0·11 mol l^{-1} sodium hydroxide solution was required to neutralise the excess hydrochloric acid.

(*a*) Show by calculation that 0·0055 mol of hydrochloric acid reacted with the magnesium carbonate. **2**

(*b*) (i) Write the balanced equation for the reaction of magnesium carbonate with hydrochloric acid. **1**

(ii) Calculate the percentage purity of the magnesium carbonate. **2**

2004 AH L6a+6b

6. A 1·11 g sample of steel containing manganese was dissolved in nitric acid. The manganese(II) ions formed were then oxidised to permanganate ions. The resulting purple solution was made up to 100 cm^3 in a standard flask.

In a titration, a 25·0 cm^3 portion of the permanganate solution was reduced by 30·1 cm^3 of 0·0020 mol l^{-1} iron(II) sulphate solution.

(*a*) Using information from the Data Booklet, write a redox equation for the reaction between the permanganate ions and iron(II) ions. **1**

(*b*) (i) Calculate the number of moles of permanganate ions in the 25·0 cm^3 portion titrated. **2**

(ii) Calculate the percentage by mass of manganese in the original sample of steel. **2**

2006 AH L2

2. 25·0 cm^3 of an acidified solution of potassium oxalate, $K_2C_2O_4$, was heated to 80 °C and titrated with a standard solution of 0·020 mol l^{-1} potassium permanganate, $KMnO_4$.

The end-point was reached when 22·5 cm^3 of $KMnO_4$ solution had been added.

The ion-electron equations for the reactions involved are:

$$C_2O_4^{2-}(aq) \longrightarrow 2CO_2(g) + 2e^-$$

$$MnO_4^-(aq) + 8H^+(aq) + 5e^- \longrightarrow Mn^{2+}(aq) + 4H_2O(\ell)$$

(*a*) How would the end-point of the titration be determined? **1**

(*b*) Write the redox equation for the reaction. **1**

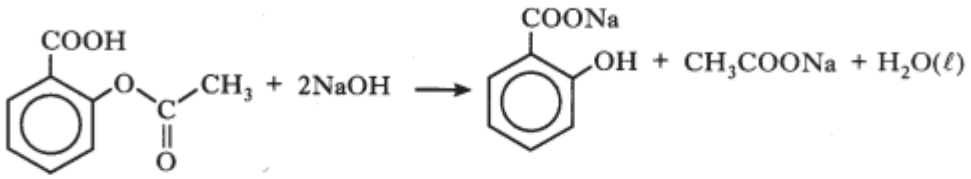
(*c*) Calculate the concentration of the potassium oxalate solution used in this titration. **3**

2005 AH L6a+6b

6. As well as the active ingredient, aspirin tablets contain other substances.

In a PPA experiment the aspirin content was determined by the method indicated in steps 1 and 2 below.

Step 1 Crushed tablets were simmered in excess sodium hydroxide solution.



Aspirin
(1 mole = 180 g)

Step 2 The excess sodium hydroxide was determined by back titration with a standard solution of sulphuric acid.

$$2NaOH + H_2SO_4 \longrightarrow Na_2SO_4 + 2H_2O$$

Three aspirin tablets were added to 25·0 cm^3 of 1·00 $mol\,l^{-1}$ sodium hydroxide solution and simmered for 30 minutes. When cooled the reaction mixture was diluted to exactly 250 cm^3 in a standard flask.

25·0 cm^3 samples were then titrated with 0·0500 $mol\,l^{-1}$ sulphuric acid until concordant results were obtained. The average titre was 15·2 cm^3.

(*a*) (i) Calculate the number of moles of sulphuric acid in the average titre. **1**

(ii) Calculate the number of moles of excess sodium hydroxide in the **standard flask**. **1**

(iii) Calculate the number of moles of sodium hydroxide which reacted with the aspirin. **1**

(iv) Calculate the average mass of pure aspirin in each tablet. **2**

(*b*) Why does a back titration technique have to be used to determine the mass of aspirin? **1**

2008 AH L6

6. Sodium hypochlorite, NaClO, is the active ingredient in household bleach. The concentration of the hypochlorite ion, ClO^-, can be determined in two stages.

In stage 1, an acidified iodide solution is added to a solution of the bleach and iodine is formed.

$$ClO^-(aq) + 2I^-(aq) + 2H^+(aq) \rightarrow I_2(aq) + Cl^-(aq) + H_2O(\ell)$$

In stage 2, the iodine formed is titrated with sodium thiosulphate solution.

$$2S_2O_3^{2-}(aq) + I_2(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$$

10·0 cm^3 of a household bleach was diluted to 250 cm^3 in a standard flask.

25·0 cm^3 of this solution was added to excess acidified potassium iodide solution.

The solution was then titrated with 0·10 $mol\,l^{-1}$ sodium thiosulphate using an appropriate indicator.

The volume of thiosulphate solution required to reach the end point of the titration was 20·5 cm^3.

(*a*) Calculate the number of moles of **iodine** which reacted in the titration. **1**

(*b*) Calculate the concentration, in $mol\,l^{-1}$, of the ClO^- in the original household bleach. **2**

2007 AH L6

6. In a PPA, 3·43 g of hydrated nickel(II) sulphate, $NiSO_4.6H_2O$, was dissolved in water and made up to 100 cm^3 in a standard flask.

20·0 cm^3 of this solution was titrated against a 0·101 $mol\,l^{-1}$ solution of EDTA using murexide as an indicator. The results are shown below.

	Rough titre	1st titre	2nd titre	3rd titre
Initial burette reading/cm^3	0·0	0·0	24·6	0·0
Final burette reading/cm^3	24·8	24·6	48·8	24·3
Volume of EDTA added/cm^3	24·8	24·6	24·2	24·3

(*a*) Give a reason why murexide is a suitable indicator in this titration reaction. **1**

(*b*) Calculate the percentage of nickel present in the hydrated salt from these experimental results. **3**

(*c*) The theoretical yield for the percentage of nickel in the hydrated salt is 22·3%. Suggest a reason for the difference between the theoretical percentage of nickel and the answer calculated in part (*b*). **1**

2010 AH L6

6. The formula of potassium hydrogen oxalate can be written as $K_xH_y(C_2O_4)_z$.

In an experiment to determine the values of **x, y** and **z**, 4·49 g of this compound was dissolved in water and the solution made up to one litre.

(*a*) 20·0 cm^3 of the solution was pipetted into a conical flask and then titrated with 0·0200 $mol\,l^{-1}$ acidified potassium permanganate at 60 °C. The average titre volume was 16·5 cm^3.

The equation for the reaction taking place in the conical flask is

$$5C_2O_4^{2-} + 16H^+ + 2MnO_4^- \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$$

(i) What colour change would indicate the end point of the titration? **1**

(ii) From the titration result, calculate the number of moles of oxalate ions, $C_2O_4^{2-}$, in 20·0 cm^3 of the solution. **1**

(iii) Calculate the mass of oxalate ions in one litre of the solution. **1**

(iv) Using another analytical procedure, 4·49 g of potassium hydrogen oxalate was found to contain 0·060 g of hydrogen.

Use this information with the answer to (*a*)(iii) to calculate the mass of potassium in this sample. **1**

(*b*) Calculate the values of **x**, **y** and **z**. **2**

2011 AH L4b

4. Iron and manganese are transition metals which have many uses in industry.

(*b*) The transition metal titanium is the seventh most abundant element in the Earth's crust.

Two of the reactions involved in the conversion of the ore ilmenite, $FeTiO_3$, into metallic titanium are shown below.

Step 1—Ilmenite is reacted with concentrated sulphuric acid.

$$FeTiO_3(s) + 3H_2SO_4(\ell) \longrightarrow FeSO_4(aq) + Ti(SO_4)_2(aq) + 3H_2O(\ell)$$

Step 2—After separation the titanium sulphate is reacted with sodium hydroxide.

$$Ti(SO_4)_2(aq) + 4NaOH(aq) \longrightarrow TiO_2(s) + 2H_2O(\ell) + 2Na_2SO_4(aq)$$

How many kilograms of titanium oxide can theoretically be produced from 3·25 kg of ilmenite? **2**

2011 AH L5b(iii)

5. The PPA "Complexometric Determination of Nickel using EDTA" has two main stages.

Stage 1 Preparation of nickel(II) sulphate solution.

Stage 2 Titration of the nickel(II) sulphate solution with EDTA.

(*b*) In **Stage 2**, 25·0 cm^3 of the nickel(II) sulphate solution were titrated against 0·110 $mol\,l^{-1}$ EDTA solution.

The results of the titrations are shown below

	Rough titre	**1st titre**	**2nd titre**
Initial burette reading/cm^3	2·00	25·90	10·00
Final burette reading/cm^3	25·90	49·40	33·60
Volume of EDTA added/cm^3	23·90	23·50	23·60

The equation for the reaction is represented by

$$Ni^{2+}(aq) + [EDTA]^{4-}(aq) \longrightarrow Ni[EDTA]^{2-}(aq)$$

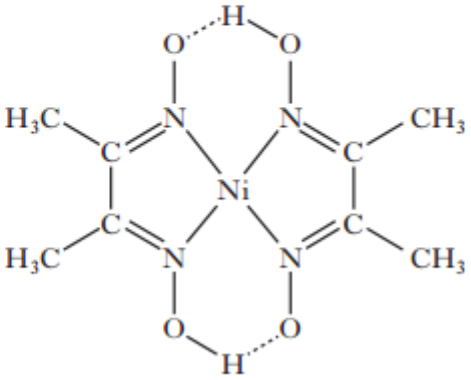
(iii) The accurate mass of the nickel(II) sulphate used was 2·656 g.

Calculate the percentage by mass of nickel present in the hydrated salt from these experimental results. **3**

2013 AH L9b and 2013 revAH L5b

9. Nickel(II) ions react quantitatively with dimethylglyoxime ($C_4H_8O_2N_2$) forming a complex which precipitates out as a red solid. The equation for the reaction and the structure of the complex are shown below.

$$Ni^{2+} + 2C_4H_8O_2N_2 \rightarrow Ni(C_4H_7O_2N_2)_2 + 2H^+$$



Relative formula mass = 288·7

(*b*) When 0·968 g of an impure sample of nickel(II) sulphate, $NiSO_4.7H_2O$, was dissolved in water and reacted with dimethylglyoxime, 0·942 g of the red precipitate was formed.

Calculate the percentage, by mass, of nickel in the impure sample of nickel(II) sulphate. **2**

2013 revAH L9

9. (*a*) State one of the characteristics of a primary standard. **1**

(*b*) As part of an AH Chemistry investigation, a student had to prepare a standard solution of sodium carbonate.

Outline how the student would prepare this standard solution from pure sodium carbonate. **2**

(*c*) Outline how 250 cm^3 of 0·20 mol l^{-1} sodium carbonate solution would be prepared from a standard 1·00 mol l^{-1} sodium carbonate solution. **2**

2014 AH L4f

4. A chromium compound is known to exist in the following three isomeric forms. The co-ordination number of chromium is the same in each isomer.

A $[Cr(H_2O)_6]^{3+}(Cl^-)_3$

B $[Cr(H_2O)_5Cl]^{2+}(Cl^-)_2.H_2O$

C $[Cr(H_2O)_4Cl_2]^+(Cl^-).2H_2O$

(*f*) 2·565 g of one of the above forms of the compound was dissolved in water to give 100 cm^3 of solution.

Silver(I) nitrate solution was added until no more precipitate was formed.

The mass of silver(I) chloride produced was 2·748 g.

(i) Calculate the number of moles of chromium compound dissolved. **1**

(ii) How many moles of silver(I) chloride were produced? **1**

(iii) Which one of the three isomers had been dissolved? **1**

2. The dicarboxylic acid, oxalic acid, has molecular formula $H_2C_2O_4$.

It can be prepared by reacting calcium oxalate with sulfuric acid.

$$H_2SO_4(aq) + CaC_2O_4(s) + xH_2O(\ell) \rightarrow CaSO_4.xH_2O(s) + H_2C_2O_4(aq)$$

(*a*) Draw a structural formula for oxalic acid. **1**

(*b*) 4·94 g of $CaSO_4.xH_2O$ was dehydrated to produce 3·89 g of $CaSO_4$.

Determine the value of x. **2**

(*c*) The equation for the reaction between oxalic acid solution and sodium hydroxide solution is

$$H_2C_2O_4(aq) + 2NaOH(aq) \rightarrow Na_2C_2O_4(aq) + 2H_2O(\ell)$$

A student used a standard solution of 0·0563 mol l^{-1} oxalic acid to standardise 20·0 cm^3 of approximately 0·1 mol l^{-1} sodium hydroxide solution.

The raw results for the titration are given in the table.

	1st attempt	**2nd attempt**	**3rd attempt**
Final burette reading/cm^3	17·2	33·8	16·6
Initial burette reading/cm^3	0·0	17·2	0·1
Titre/cm^3	17·2	16·6	16·5

Calculate the accurate concentration of the sodium hydroxide solution. **2**

(*d*) Oxalic acid is a primary standard but sodium hydroxide is not. State **one** characteristic of sodium hydroxide that makes it unsuitable as a primary standard. **1**

12. To determine the composition of an old coin containing silver, copper and nickel, a student dissolved the coin of mass 10·04 g in nitric acid. The resulting solution was diluted with deionised water to 1000 cm^3 in a standard flask.

(*a*) 0·2 $mol\,l^{-1}$ hydrocholoric acid was added to 100 cm^3 of this solution until precipitation of silver(I) chloride was complete. After filtration, the precipitate was washed and dried and found to have a mass of 0·620 g.

(i) Calculate the percentage, by mass, of silver in the coin. **2**

(ii) Suggest how the student would test that no silver(I) ions remained in the solution. **1**

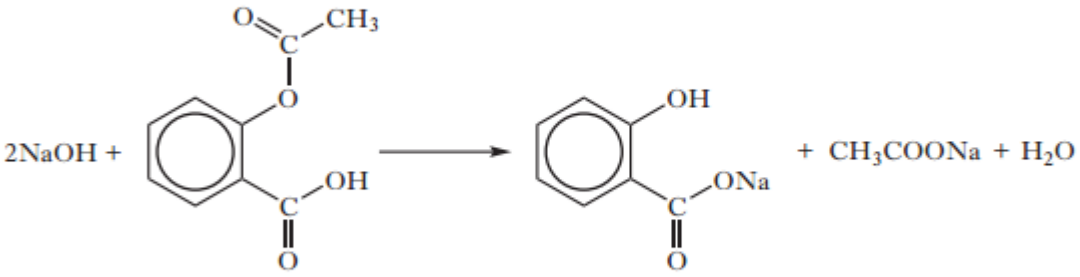
(*b*) The filtrate was treated to reduce the copper(II) ions to copper(I) ions. Ammonium thiocyanate solution was added to precipitate the copper as copper(I) thiocyanate.

$$Cu^+(aq) + CNS^-(aq) \rightarrow Cu^+CNS^-(s)$$

After filtration, drying and weighing, the precipitate was found to weigh 0·320 g. Calculate the percentage, by mass, of copper in the coin. **2**

7. In a **PPA**, the acetylsalicylic acid ($C_9H_8O_4$) content of an aspirin tablet was determined using a back titration.

Five aspirin tablets were crushed and added to 25·0 cm^3 of 1·00 mol l^{-1} sodium hydroxide solution. The mixture was heated and allowed to simmer for 30 minutes.



The resulting mixture was allowed to cool before being transferred to a 250 cm^3 standard flask and made up to the mark with deionised water.

25·0 cm^3 samples of this solution were titrated with 0·050 mol l^{-1} sulphuric acid.

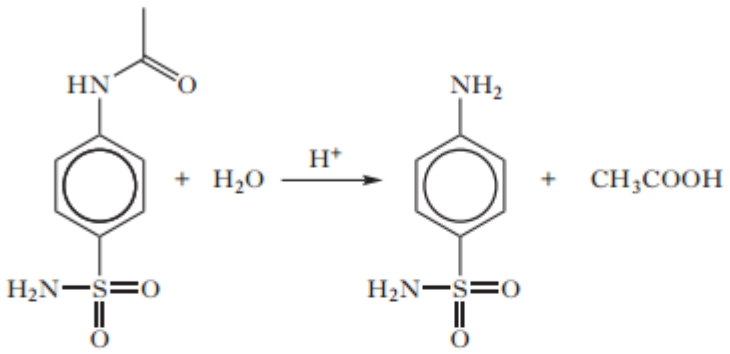
$$2NaOH + H_2SO_4 \longrightarrow Na_2SO_4 + 2H_2O$$

The results of the titration are shown in the table.

	Rough titration	**1st titration**	**2nd titration**
Initial burette reading/cm^3	0·0	9·0	17·7
Final burette reading/cm^3	9·0	17·7	26·3
Volume used/cm^3	9·0	8·7	8·6

4. Sulfa drugs are compounds with antibiotic properties. Sulfa drugs can be prepared from a solid compound called sulfanilamide.

Sulfanilamide is prepared in a six stage synthesis. The equation for the final step in the synthesis is shown.



$C_8H_{10}N_2SO_3$ $C_6H_8N_2SO_2$

4-acetamidobenzenesulfonamide sulfanilamide

(*a*) What type of reaction is this? **1**

(*b*) The sulfanilamide is separated from the reaction mixture and recrystallised from boiling water.

Why is the recrystallisation necessary? **1**

(*c*) Calculate the percentage yield of sulfanilamide if 4·282 g of 4-acetamidobenzenesulfonamide produced 2·237 g of sulfanilamide. **3**

(*d*) Describe how a mixed melting point experiment would be carried out and the result used to confirm that the product was pure. **2**

(*e*) Suggest another analytical technique which could be used to indicate whether the final sample is pure. **1**

4. As part of an Advanced Higher Chemistry project, a student determined the chloride ion concentration of seawater by two different methods.

Volumetric method
A sample of seawater was titrated with standard silver nitrate solution.

Gravimetric method
A sample of seawater was reacted with standard silver nitrate solution to form a precipitate. The precipitate was collected by filtration and weighed.

(a) For the volumetric method, a $0{\cdot}1\ \text{mol}\ l^{-1}$ standard solution of silver nitrate was prepared by following the instructions below.

1. Dry 5 g of silver nitrate for 2 hours at 100 °C and allow to cool.
2. Weigh accurately approximately 4·25 g of solid silver nitrate.
3. Use this sample to prepare 250 cm^3 of standard silver nitrate solution.

(i) State what is meant by "weigh accurately approximately"

4·25 g of solid silver nitrate. **1**

(ii) Outline how the student would have prepared the standard silver nitrate solution. **2**

(iii) Samples of the diluted seawater were titrated and the average titre was found to be 3·9 cm^3.

Suggest an improvement the student could make to reduce the uncertainty in the titre value. **1**

(b) For the gravimetric method, standard silver nitrate solution was added to a seawater sample to form a precipitate of silver chloride.

(i) Describe how the filtration should have been carried out to ensure a fast means of separating the precipitate from the reaction mixture. **1**

(ii) After the precipitate was filtered, the filtrate was tested with a few drops of silver nitrate solution.

Suggest why the student tested the filtrate in this way. **1**

(c) The student also planned to carry out an analysis of chloride ion concentration in fresh river water.

Explain why the volumetric method, rather than the gravimetric method, would be more appropriate for the analysis of chloride ion concentration in fresh river water. **1**

2014 AH L9c and 2014 revAH L7b(ii)

9. In a PPA, aspirin, $C_9H_8O_4$, was prepared by reacting 2-hydroxybenzoic acid with ethanoic anhydride. A catalyst was added and when the crude aspirin had formed, impurities were removed by dissolving them in a suitable solvent. The crystals of aspirin were dried in an oven before determining the purity and the percentage yield.

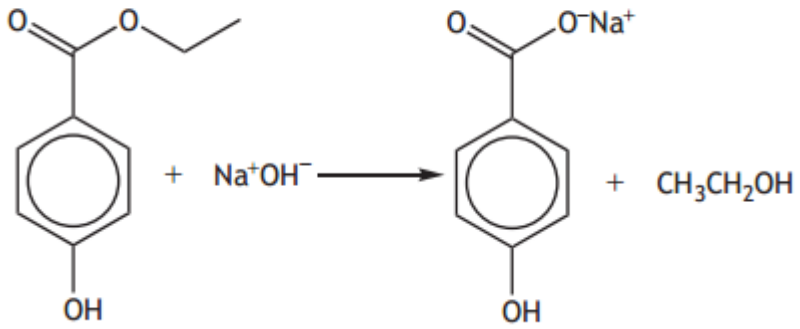
(*c*) The percentage yield for the reaction was 67%. Calculate the minimum mass of 2-hydroxybenzoic acid required to produce 5·00 g of aspirin. **3**

2016 AH L9

9. Parabens are used as preservatives in cosmetics, pharmaceutical products and foods. Parabens are esters of 4-hydroxybenzoic acid.

One common paraben used as a food preservative is ethylparaben.

(b) Another preservative is sodium 4-hydroxybenzoate. It can be prepared by refluxing ethylparaben with sodium hydroxide solution.



(v) In this experiment, the percentage yield of 4-hydroxybenzoic acid was 77·5%.

Calculate the mass of ethylparaben (GFM = 166 g) required to produce 2·48 g of 4-hydroxybenzoic acid (GFM = 138 g). **2**