WJEC EDUQAS AS CHEMISTRY · COMPONENT 1

# CHEMISTRY



EXAM QUESTION PACK

## 1.3 CHEMICAL CALCULATIONS

| SCORE | PERCENTAGE |
| --- | --- |
| /141 | % |

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1.

*(a)* Ethanol, $C_2H_5OH$, is a liquid at room temperature. It is being increasingly used as a fuel.

(i) Write the equation that represents the standard molar enthalpy change of formation ($\Delta H_f$) of ethanol. [1]

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(ii) Suggest why this enthalpy change cannot be measured directly. [1]

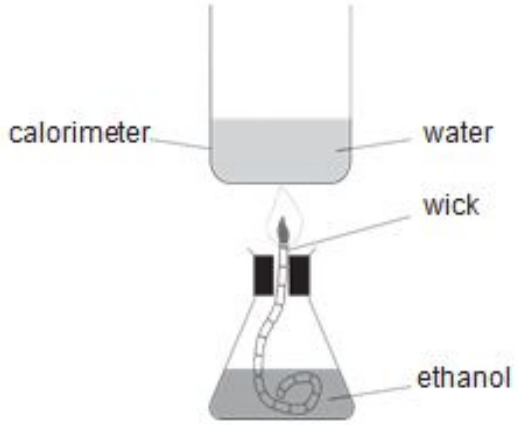
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*(b)* Enthalpy changes of combustion can often be measured directly. The equation for the reaction which represents the enthalpy change of combustion ($\Delta H_c$) of ethanol is as follows.

$$C_2H_5OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(l)$$

A student used the apparatus below to determine the enthalpy change of combustion of ethanol.




The student obtained the following results.

| | |
|---|---|
| Mass of spirit burner + ethanol at start | = 72.27 g |
| Mass of spirit burner + ethanol after combustion | = 71.46 g |
| Temperature of water at start | = 21.5 °C |
| Temperature of water after combustion | = 75.5 °C |
| Volume of water in calorimeter | = 100 $cm^3$ |

The energy released in the experiment can be calculated using the formula

$$\text{energy released} = mc\Delta T$$

where m = mass of the water in grams (assume 1 $cm^3$ has a mass of 1 g)
c = 4.2 $Jg^{-1}°C^{-1}$
$\Delta T$ = change in temperature of the water

(i) Calculate the energy released in the experiment

[1]

*Energy released* = . . . . . J

(ii) The enthalpy change of combustion of ethanol is defined as the energy change per mol of ethanol burned.
Use your answer to (i) to calculate the enthalpy change of combustion of ethanol.

Give your answer in kJ $mol^{-1}$ and correct to **3 significant figures**. Include the sign.

[3]

(c) Another student did not carry out an experiment to find $\Delta H_c$ of ethanol. He looked up the literature value on a respected internet site.

How would you expect the numerical values obtained by the two students to differ? Explain your answer.

You may assume that both values were found under the same conditions of temperature and pressure.

[2]

(d) The students then used the apparatus from (b) to find the enthalpy change of combustion of higher relative molecular mass alcohols. They found that as the number of carbon atoms increased the value of the enthalpy change of combustion became more negative.

(i) Write the equation for the reaction which represents the enthalpy change of combustion of propanol, $C_3H_7OH$. [1]

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(ii) In terms of bond strengths, explain why enthalpy changes of combustion are negative.

[1]

(iii) Explain why the enthalpy change of combustion of propanol is more negative than that of ethanol

[1]

(e) Recent research has been carried out to find economic and environmentally friendly uses for waste straw and wood chippings.
The process of gasification involves the material being partly combusted at a temperature of about 700 °C to give a mixture consisting mainly of hydrogen and carbon monoxide but also some carbon dioxide.

Another approach has been to use enzyme catalysed reactions to change the waste material into glucose and then to ethanol.

Comment on the economic and environmental factors involved in both of these processes.

[4] QWC [2]

**(Total 17)**

2. In an experiment, Aled titrated 25.00 $cm^3$ of potassium hydroxide solution with hydrochloric acid, and obtained the following results.

| | 1 | 2 | 3 | 4 |
|---|---|---|---|---|
| Initial burette reading / $cm^3$ | 0.10 | 0.25 | 1.20 | 21.30 |
| Final burette reading / $cm^3$ | 20.85 | 20.45 | 21.30 | 41.60 |
| Volume used / $cm^3$ | | | | |

(a) Complete the table to show the volume used in each titration.

[1]

(b) Calculate the mean volume that Aled should use for his further calculations.

[1]

.....$cm^3$

3.

Ethanol is an important industrial chemical and can be made by the direct hydration of ethene using a phosphoric acid catalyst.

$$CH_2=CH_2(g) + H_2O(g) \rightleftharpoons CH_3CH_2OH(g) \quad \Delta H = -46\ kJ\ mol^{-1}$$

*(a)* State, giving your reasons, the general conditions of temperature and pressure required to give a high equilibrium yield of ethanol in this process. [4]

*QWC* [1]

*(b)* Using the standard enthalpy change for the reaction above and the standard enthalpy changes of formation ($\Delta H_f^{\ominus}$) given in the table below, calculate the standard enthalpy change of formation of gaseous ethanol. [3]

| Compound | $\Delta H_f^{\ominus}$ / kJ mol$^{-1}$ |
|---|---|
| $CH_2{=}{=}CH_2(g)$ | 52.3 |
| $H_2O(g)$ | –242 |

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*(c)* Another way of calculating the enthalpy change of a reaction is by using average bond enthalpies. Use the values in the table below to calculate the enthalpy change for the direct hydration of ethene. [3]



| Bond | Average bond enthalpy / kJ $mol^{-1}$ |
|---|---|
| C —C | 348 |
| C =C | 612 |
| C —H | 412 |
| C —O | 360 |
| O—H | 463 |

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*(d)* (i) Give a reason why the calculated value in *(c)* is different to the actual value, –46 kJ $mol^{-1}$. [1]

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(ii) Explain whether your answer to part (i) supports the use of average bond enthalpies to calculate the energy change for a reaction. [1]

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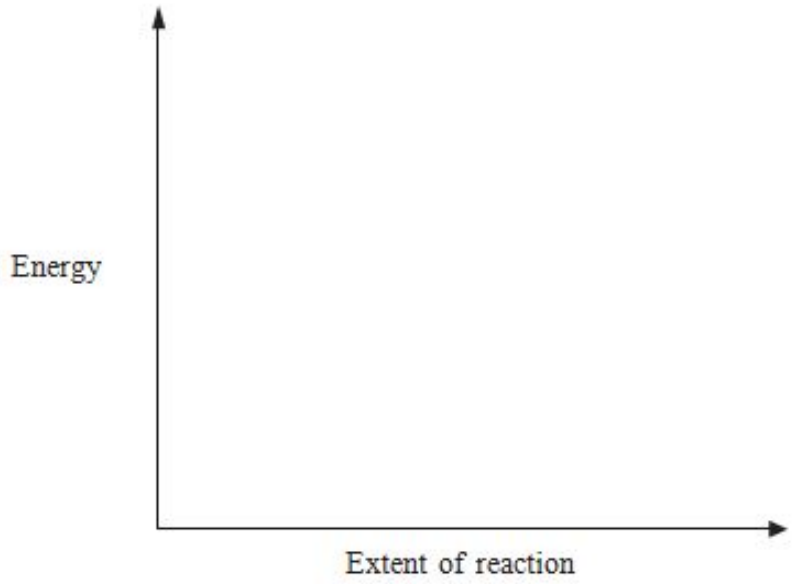
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(e) Phosphoric acid is an example of a heterogeneous catalyst.
Explain the term *heterogeneous* in this context. [1]

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(f) (i) Sketch on the axes below the energy profile for an exothermic reaction. [1]




(ii) On the same axes, sketch and label the energy profile if the same reaction is carried out using a catalyst. [1]

Total [16]

4. Judith carried out three experiments to study the reaction between powdered magnesium and hydrochloric acid.

She used a gas syringe to measure the volume of hydrogen evolved, at room temperature and pressure, at set intervals. In each case, the amount of acid used was sufficient to react with all the magnesium.

$$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$$

The details of each experiment are shown in Table 1 below.

| Experiment | Mass of magnesium / g | Volume of HCl / $cm^3$ | Concentration of HCl / mol $dm^{-3}$ |
|---|---|---|---|
| **A** | 0.061 | 40.0 | 0.50 |
| **B** | 0.101 | 40.0 | 1.00 |
| **C** | 0.101 | 20.0 | 2.00 |

**Table 1**

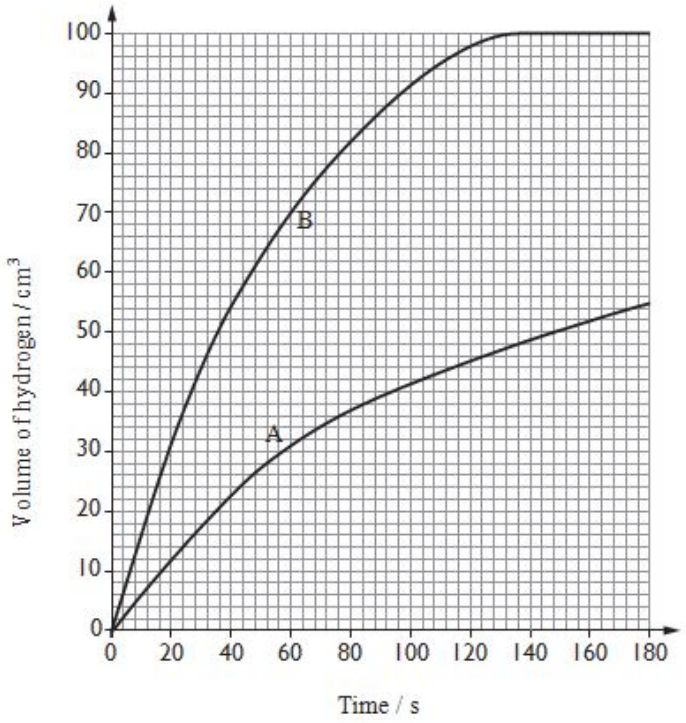
The results obtained in experiment **C** are shown in Table 2 below.

| Time / s | Volume of hydrogen / $cm^3$ |
|---|---|
| 0 | 0 |
| 20 | 50 |
| 40 | 75 |
| 60 | 88 |
| 80 | 92 |
| 100 | 100 |
| 120 | 100 |

**Table 2**

(a) The results for experiments **A** and **B** have already been plotted on the grid below. On the same grid, plot the results for experiment **C** and draw a line of best fit.

[3]




(b)(i) State in which experiment the reaction begins most rapidly and **use the graph** to explain your choice.

[2]

(ii) By referring to Table 1 give an explanation of your answer in part (i).

[1]

(c) State the volume of hydrogen evolved after 30 seconds in experiment **B**.

[1]

(d) Using **only** the values in Table 1, show that the acid is in excess in experiment **C**.

[2]

(e)(i) In experiment **A**, 0.061 g of magnesium produces 60 $cm^3$ of hydrogen. If 0.122 g of magnesium were used, under the same conditions, then 120 $cm^3$ would be produced. Explain why using 0.610 g would not produce 600 $cm^3$ of hydrogen.

[1]

(ii) Calculate the volume of hydrogen produced using 0.610 g of magnesium.
(1 mole of gas molecules occupies 24 $dm^3$ at 25 °C)

[2]

(f) State one method of slowing down the reaction in experiment **C** and use collision theory to explain your choice. Assume that the quantities of magnesium and hydrochloric acid are the same as those in Table 1.

[3] *QWC* [1]

**(Total 16)**

5. Hydromagnesite is a mixture of magnesium carbonate and soluble impurities. A student crushed some hydromagnesite and added a sample of mass 0.889 g to excess dilute hydrochloric acid so that the magnesium carbonate component reacted fully.

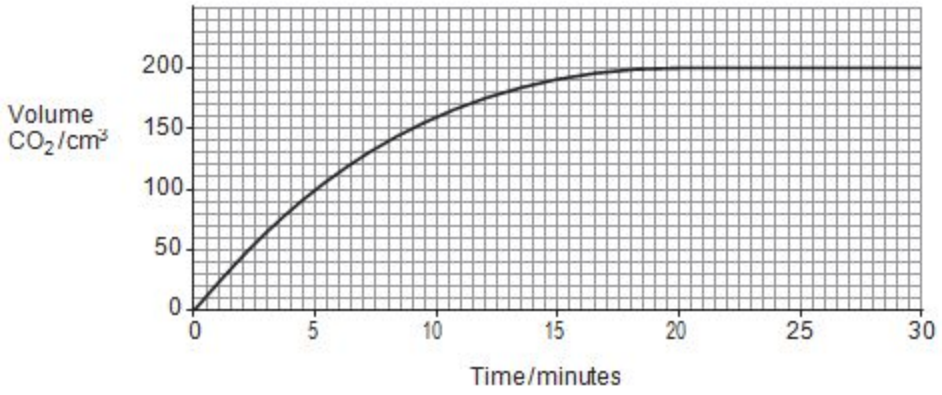
(a) Explain why the rock was crushed before being added to the acid.

[1]

(b) Write the equation for the reaction between magnesium carbonate and dilute hydrochloric acid.

[1]

(c) The gas formed was collected in a gas syringe and its volume was measured over a period of time. The volumes and times were plotted. The volume of 1 mol of gas under these conditions is 24.0 $dm^3$




Describe what happened to the rate of the reaction over the 30 minute period.Explain why any changes in the rate occurred

[3]

(d) Other than by using an indicator, how would the student know that hydrochloric acid was in excess?

[1]

(e) (i) Use the graph to calculate how many moles of magnesium carbonate reacted with the hydrochloric acid. [2]

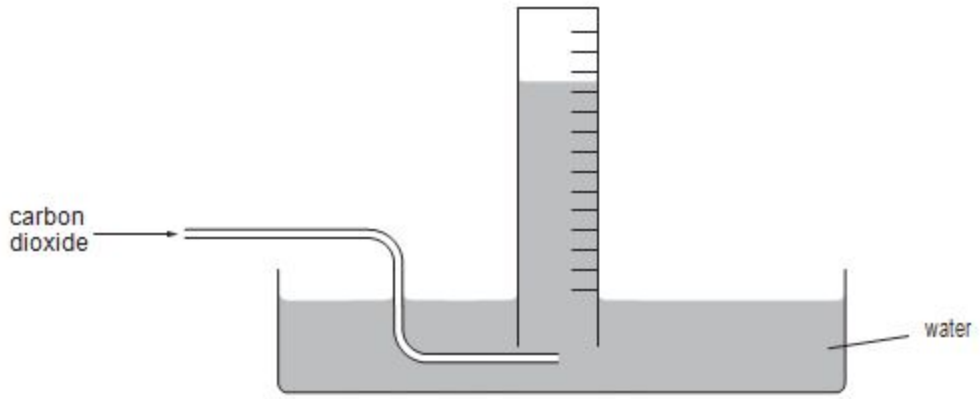
*Number of moles* $MgCO_3$ = ..... mol

(ii) Find the mass of magnesium carbonate that reacted and hence the percentage of magnesium carbonate present in hydromagnesite.

[2]

*Percentage of magnesium carbonate* = . . . . . %

(f) A student wanted to carry out this experiment on another sample of hydromagnesite. He did not have a gas syringe and therefore he decided to collect the carbon dioxide over water in a measuring cylinder.




Explain what effect this would have on the results of the experiment. You should assume that the gas syringe and the measuring cylinder can both be read to the same precision

[2]

(g) When magnesium carbonate is heated it decomposes to make magnesium oxide and carbon dioxide.

$$MgCO_3(s) \longrightarrow MgO(s) + CO_2(g)$$

Magnesium oxide has a very high melting temperature and so can be used to line furnaces.

What is the atom economy for the production of magnesium oxide from magnesium carbonate?

[2]

*Atom economy* = . . . . . %

**(Total 14)**

6.

*(a)* Sodium carbonate can be manufactured in a two-stage process as shown by the following equations.

$$NaCl + NH_3 + CO_2 + H_2O \longrightarrow NaHCO_3 + NH_4Cl$$

$$2NaHCO_3 \longrightarrow Na_2CO_3 + H_2O + CO_2$$

Calculate the maximum mass of sodium carbonate which could be obtained from 900 g of sodium chloride. [3]

*Maximum mass of sodium carbonate* = ..... g

*(b)* Sodium carbonate can form a hydrate, $Na_2CO_3.xH_2O$.

When 4.64 g of this hydrate was heated, 2.12 g of anhydrous $Na_2CO_3$ remained.

(i) State the mass of water in 4.64 g of the hydrate. [1]

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(ii) Calculate the number of moles of sodium carbonate and the number of moles of water in 4.64 g of the original hydrate. Use these values to calculate the value of $x$ in $Na_2CO_3.xH_2O$. [2]

$x$ = .....

(c) Hannah is given an impure sample of anhydrous sodium carbonate and she carries out an experiment to determine the percentage of sodium carbonate in the sample.
She finds that she needs 18.0 $cm^3$ of hydrochloric acid of concentration 0.50 mol $dm^{-3}$ to react completely with 0.55 g of the impure sample. The impurity does not react with hydrochloric acid. The equation for the reaction is given below.

$$Na_2CO_3 + 2HCl \longrightarrow 2NaCl + H_2O + CO_2$$

(i) Calculate the number of moles of HCl used in the titration. [1]

*Number of moles of* HCl = ..... mol

(ii) Deduce the number of moles of $Na_2CO_3$ that reacted with the HCl. [1]

---

(iii) Calculate the mass of $Na_2CO_3$ in the sample. [1]

*Mass of* $Na_2CO_3$ *in sample* = ..... g

(iv) Calculate the percentage by mass of $Na_2CO_3$ in the sample. [1]

*Percentage by mass* = ..... %

Total [10]

7. Jewels such as diamonds, rubies and emeralds are highly valued but are all closely related to much less precious materials.

(a) Emeralds are a form of the mineral beryl, with their green colour due to the impurities present.

A sample of beryl contains 10.04 % aluminium, 53.58 % oxygen and 31.35 % silicon by mass, with beryllium making up the remainder. Its molecular formula is $Al_2Be_xSi_6O_{18}$. Find the percentage by mass of beryllium in the compound and hence calculate the value of $x$ in this formula. [3]

$x$ = .....

(b) The most common form of carbon is graphite, however the element also exists in the form of diamond.
We can calculate the standard enthalpy change of reaction for making diamond from graphite using Hess' Law.

| Reaction | Standard enthalpy change of reaction / kJ $mol^{-1}$ |
|---|---|
| $C(diamond) + O_2(g) \rightarrow CO_2(g)$ | –395.4 |
| $C(graphite) + O_2(g) \rightarrow CO_2(g)$ | –393.5 |

(i) State Hess' Law.

[1]

(ii) Use Hess' Law and the data in the table on page 4 to calculate the enthalpy change of the reaction below. [2]

$$C(graphite) \longrightarrow C(diamond)$$

*Enthalpy change of reaction* = ..... kJ $mol^{-1}$

(iii) Kyran states that because diamond is an element, its enthalpy of formation under standard conditions must be zero.State whether Kyran is correct and give a reason to support your answer.

[1]

(iv) Most diamonds used in jewellery come from natural sources, but it is possible to produce diamonds artificially although these are rarely of gemstone quality.

(I) One proposed use of artificial diamond is to protect medical implants. To cover a particular implant, a volume of 2.08 $cm^3$ of diamond is needed. Calculate the mass of diamond required

[1]

[Density of diamond under standard conditions = 3.51 g $cm^{-3}$]

*Mass of diamond* = ..... g

(II) The process of producing diamond from graphite has a yield of 93 %. Calculate the mass of graphite needed to make the diamond required.

[2]

*Mass of graphite* = ..... g

**(Total 10)**

8.

*(a)* Lithium was discovered in 1817 by the Swedish chemist Johan August Arfwedson. Its name derives from the Greek word *lithos*, meaning 'stone', to reflect its discovery in a solid mineral, as opposed to potassium, which had been isolated from plant ashes 10 years earlier. Naturally occurring lithium is composed of two stable isotopes – $^{6}Li$ and $^{7}Li$.

In a mass spectrometer, a sample of lithium must be ionised before it can be analysed.

(i) Describe how vaporised atoms of Li are converted into $Li^{+}$ ions in a mass spectrometer. [2]

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(ii) Suggest why no more than the minimum energy is used to ionise the sample of lithium.

[1]

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(iii) State the difference, if any, between the chemical properties of the isotopes $^{6}Li$ and $^{7}Li$, giving a reason for your answer. [2]

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(b) The mass spectrum of a naturally occurring sample of lithium gave the following results.

| Isotope | % abundance |
|---|---|
| $^{6}Li$ | 7.25 |
| $^{7}Li$ | 92.75 |

These results can be used to determine the relative atomic mass of the lithium sample.

(i) Calculate the relative atomic mass of the sample

[2]

*Relative atomic mass* = . . . . . .

(ii) State and explain which of the $Li^{+}$ ions formed from the isotopes of Li will be deflected more in a mass spectrometer. [1]

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*(c)* Lithium hydroxide reacts with ammonium sulfate to form ammonia, lithium sulfate and water as shown in the equation below.

$$(NH_4)_2SO_4 + 2LiOH \longrightarrow 2NH_3 + Li_2SO_4 + 2H_2O$$

A 2.06 g sample of ammonium sulfate reacted exactly with 29.80 $cm^3$ of a lithium hydroxide solution.

(i) Calculate the amount, in moles, of $(NH_4)_2SO_4$ in 2.06 g of ammonium sulfate. Give your answer to **three** significant figures. [2]

*Number of moles* = ..... mol

(ii) Calculate the concentration, in mol $dm^{-3}$, of the lithium hydroxide solution used. [2]

*Concentration* = ..... mol $dm^{-3}$

(iii) Calculate the percentage atom economy for the production of ammonia in the reaction between ammonium sulfate and lithium hydroxide. [2]

*Atom economy* = ..... %

Total [14]

9. (a) An aqueous solution of methanoic acid can be used to dissolve 'limescale' in kettles. The concentration of a methanoic acid solution used for this purpose can be found by a titration using sodium hydroxide solution. For this purpose a 25.0 $cm^3$ sample of aqueous methanoic acid was diluted to 250 $cm^3$.

(i) State the name of the piece of apparatus used to:

(I) measure out 25.0 $cm^3$ of aqueous methanoic acid, [1]

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(II) contain exactly 250 $cm^3$ of the diluted solution. [1]

---

(ii) A 25.0 $cm^3$ sample of the diluted methanoic acid was titrated with sodium hydroxide solution of concentration 0.200 mol $dm^{-3}$. A volume of 32.00 $cm^3$ was needed to react with all the methanoic acid present.

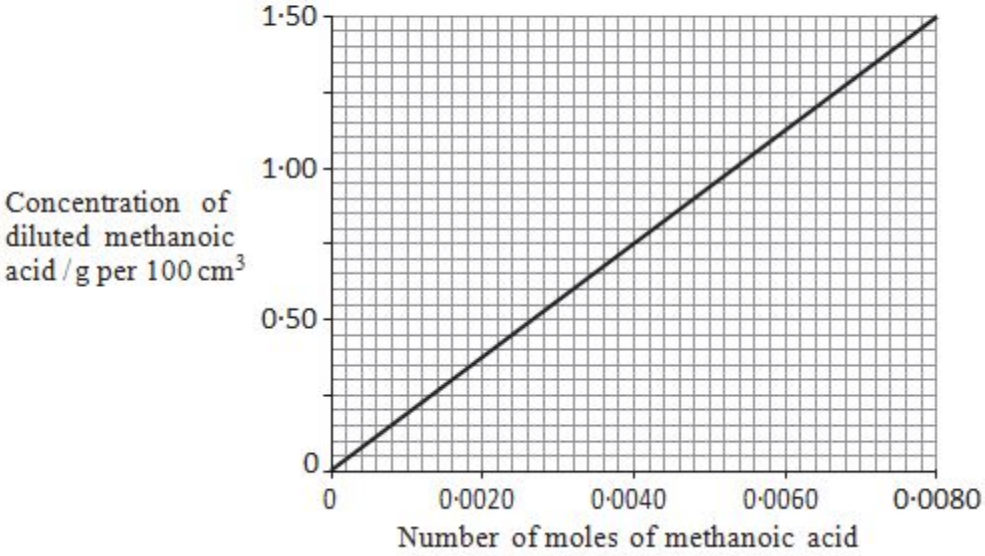
Calculate the number of moles of sodium hydroxide used.

[1

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*Moles of sodium hydroxide* = .....mol

(iii) Methanoic acid and sodium hydroxide react together in a 1:1 molar ratio.
Use the graph below and your result from (ii) to find the concentration of methanoic acid present in the diluted solution in g per 100 $cm^3$ of solution. [1]




*Concentration* = ..... g per 100 $cm^3$

(iv) State the concentration of the original methanoic acid in g per 100 $cm^3$ solution. [1]

*Original concentration* = ..... g per 100 $cm^3$

*(b)* Methanoic acid, HCOOH, can be reduced to methanol, $CH_3OH$, in a gas phase reaction, by using hydrogen in the presence of a solid ruthenium metal catalyst.

(i) Ruthenium is acting as a heterogeneous catalyst.
State the meaning of the word *heterogeneous*. [1]

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(ii) The equation for the reduction of methanoic acid is shown below.



Use the table of bond enthalpies to find the enthalpy change for this reaction. [3]

| Bond | Average bond enthalpy / kJ mol$^{-1}$ |
|---|---|
| C—H | 412 |
| C—O | 360 |
| C=O | 743 |
| H—H | 436 |
| O—H | 463 |

*Enthalpy change* = ..... kJ mol$^{-1}$

*(c)* The relative molecular mass of methanoic acid is 46.02.

State why this quantity does not have units. [1]

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*(d)* Methanoic acid reacts with propan-1-ol to give 1-propyl methanoate.

$$HCOOH + CH_3CH_2CH_2OH \rightleftharpoons HCOOCH_2CH_2CH_3 + H_2O$$

1-propyl methanoate

(i) This reaction eventually reaches dynamic equilibrium.
State what is meant by *dynamic equilibrium*. [1]

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(ii) Give the empirical formula of 1-propyl methanoate. [1]

*Empirical formula* .....

Total [12]

10. Berian was asked to find the identity of a Group 1 metal hydroxide by titration.

He was told to use the following method:

• Fill a burette with hydrochloric acid solution.

• Accurately weigh about 1.14 g of the metal hydroxide.

• Dissolve all the metal hydroxide in water, transfer the solution to a volumetric flask then add more water to make exactly 250 $cm^3$ of solution.

• Accurately transfer 25.0 $cm^3$ of this solution into a conical flask.

• Add 2-3 drops of a suitable indicator to this solution.

• Carry out a rough titration of this solution with the hydrochloric acid.

• Accurately repeat the titration several times and calculate a mean titre.

Berian's results are shown below:

Mass of metal hydroxide = 1.14 g

Concentration of acid solution = 0.730 g HCl in 100 $cm^3$ of water

Mean titre = 23.80 $cm^3$

a) Give a reason why Berian does not simply add 1.14 g of metal hydroxide to 250 $cm^3$ of water.

[1]

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(b) Name a suitable piece of apparatus for transferring 25.0 $cm^3$ of the metal hydroxide solution to a conical flask.

[1]

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(c) State why he adds an indicator to this solution.

[1]

(d) Suggest why Berian was told to carry out a rough titration first.

[1]

(e) Explain why he carried out several titrations and calculated a mean value.

[1]

(f) The equation for the reaction between the metal hydroxide and hydrochloric acid is given below.
M represents the symbol of the Group 1 metal.

$$MOH + HCl \longrightarrow MCl + H_2O$$

(i) Calculate the concentration, in mol $dm^{-3}$, of the HCl in the burette.

[2]

(ii) Calculate the number of moles of HCl used in the titration.

[1]

(iii) Deduce the number of moles of MOH in 25.0 $cm^3$ of the solution.

[1]

(iv) Calculate the total number of moles of MOH in the original solution.

[1]

(v) Calculate the relative molecular mass of MOH.

[1]

(vi) Deduce the Group 1 metal in the hydroxide.

[1]

**(Total 12)**

11. The leaves of the rhubarb plant are rich in ethanedioic acid (oxalic acid) which is a poisonous compound. A solution containing ethanedioate ions can be formed by boiling rhubarb leaves with water. It can be separated and samples titrated against acidified potassium manganate(VII) to find the concentration of the ethanedioate solution.

(a) Suggest how the ethanedioate solution could be separated from the rhubarb leaves.

[1]

*(b)* Write an ion-electron half-equation for the reduction of acidified manganate(VII) ions, $MnO_4^-$. [1]

*(c)* The ion-electron half-equation for the oxidation of ethanedioate ions is given below.

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

(i) Give the oxidation states for carbon at the start and end of this reaction. [1]

(ii) Write an equation for the reaction of acidified manganate(VII) ions with ethanedioate ions. [1]

(d) Give a reason why an indicator is not needed in this titration.

[1]

*(e)* Four samples of 25.00 $cm^3$ of the ethanedioate solution were titrated against acidified potassium manganate(VII) solution of concentration 0.0200 $mol\,dm^{-3}$. The volumes of potassium manganate(VII) solution required for complete reaction are listed below.

| | 1 | 2 | 3 | 4 |
|---|---|---|---|---|
| Volume of $KMnO_4(aq)/cm^3$ | 28.80 | 27.95 | 28.00 | 27.80 |

Use the information given to calculate the concentration of the ethanedioate solution. [4]

(f) Heating ethanedioic acid in glycerol produces methanoic acid, HCOOH.

(i) Write the expression for the acid dissociation constant, $K_a$, for methanoic acid. [1]

(ii) The value of $K_a$ for methanoic acid is $1.8 \times 10^{-4}$ mol $dm^{-3}$.
Calculate the pH of a solution of methanoic acid of concentration 0.2 mol $dm^{-3}$. [3]

(iii) A mixture of methanoic acid and sodium methanoate can be used as a buffer solution. State what is meant by a *buffer solution* and explain how a mixture of methanoic acid and sodium methanoate acts as a buffer.

[3] QWC [1]

*(g)* Acidified potassium dichromate, $K_2Cr_2O_7$, is also an oxidising agent.

(i) Give the colour change that occurs when acidified potassium dichromate acts as an oxidising agent.

[1]

(ii) When sodium hydroxide is added to a solution of potassium dichromate, a colour change occurs without a redox reaction occurring. Give the formula of the new chromium-containing ion and the colour of the solution formed.

[2]

**(Total 20)**

2.

(a)

| | 1 | 2 | 3 | 4 |
|---|---|---|---|---|
| *Volume used / cm$^3$* | 20.75 | 20.20 | 20.10 | 20.30 |

[1]

(b) 20.20 cm$^3$ [1]

3. (a) Low temperature (1)
As temperature is decreased equilibrium moves in exothermic direction. (1)

High pressure (1)

As pressure is increased equilibrium moves towards side with smaller number of gas moles (1) [4]

*QWC* The information is organised clearly and coherently, using specialist vocabulary where appropriate [1]

(b) $\Delta H$reaction = $\Delta H_f$ products – $\Delta H_f$ reactants (1)

$-46 = \Delta H_f$ ethanol – (52.3 – 242)

$\Delta H_f$ ethanol = -46 – 189.7 (1)

$\Delta H_f$ ethanol = -235.7 kJ $mol^{-1}$ (1) [3]

(c) Bonds broken = 1648 + 612 + 926 = 3186 kJ $mol^{-1}$ (1)

Bonds formed = 2060 + 348 + 360 + 463 = 3231 kJ $mol^{-1}$ (1)

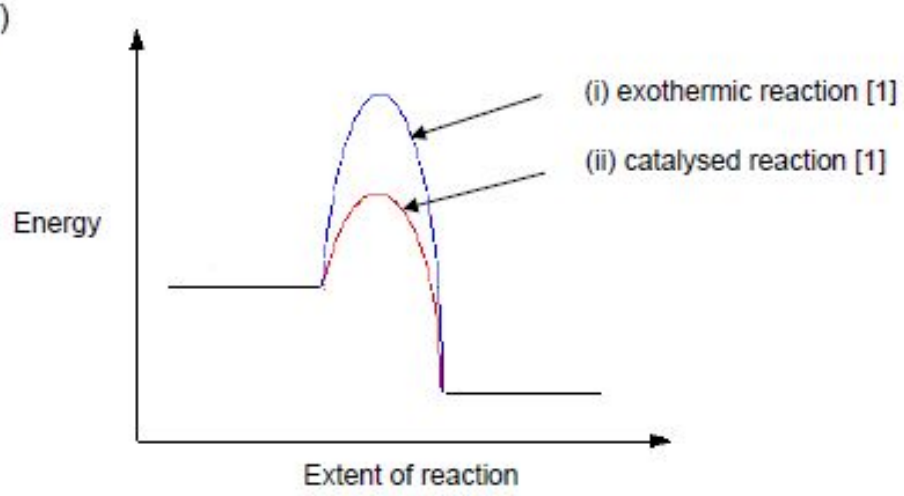
$\Delta H$ reaction = 3186 – 3231 = -45 kJ $mol^{-1}$ (1)
[3]

(d) (i) Average bond enthalpies used (not actual ones) [1]

(ii) Yes, since answers are close to each other [1]

(e) Catalyst is in different (physical) state to reactants [1]

(f)




Total [16]

4.

(a) Plotting (2)

Best fit line (1) [3]

(b) (i) C (1)

Curve steeper (1) [2]

(ii) Concentration of acid is greatest [1]

(c) 44 $cm^3$ (±1 $cm^3$) [1]

(d) Moles Mg = 0.101/24.3 = 0.00416 (1)

Moles HCl = 2 x 0.02 = 0.04 (1) [2]

(e) (i) Mg is not the limiting factor /

Mg now in excess / HCl not in excess [1]

(ii) Moles acid = 0.5 x 0.04 = 0.02 (1)

Volume $H_2$ = 0.01 x 24 = 0.24 $dm^3$

- correct unit needed (1) [2]

(f) Lower the temperature of the acid (1)

Reactants collide with less energy (1)

Fewer molecules that have the required activation energy (1)[3]

or Use pieces of magnesium (1) less surface area (1) less chance of successful collisions (1)

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Total [16]

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5.

(a) to increase rate of reaction / to increase surface area [1]

(b) $MgCO_3 + 2HCl \rightarrow MgCl_2 + CO_2 + H_2O$ (ignore state symbols) [1]

(c) rate starts fast and gradually slows (1)

because concentration becomes less so fewer collisions (per unit time) / less frequent collisions / lower probability of collisions (1)

at time = 17/18 min rate = 0 (1) [3]

(d) all the solid would all have disappeared / if more carbonate is added further effervescence is seen [1]

(e) (i) volume $CO_2 = 200\ cm^3$ (1)

moles $CO_2 = 200 / 24000 = 0.008333 =$ moles $MgCO_3$ (1)
[minimum 2 sf] [2]

(ii) mass $MgCO_3 = 0.008333 \times 84.3 = 0.702$ g (1)

$\%\ MgCO_3 = \frac{0.702}{0.889} \times 100 = 79.0\% / 79\%$ [2]

(e) carbon dioxide is soluble in water / reacts with water (1)

volume collected less therefore % / moles of $MgCO_3$ less (1) [2]

(f) use of 40.3 and 84.3 (1)

atom economy = $40.3 / 84.3 \times 100 = 47.8\%$ (1) [2]

**Total [14]**

6.

| | | | | |
|---|---|---|---|---|
| (a) | | Moles NaCl = $\frac{900}{58.5}$ = 15.38 | (1) | |
| | | Moles $Na_2CO_3$ = 7.69 | (1) | |
| | | Mass $Na_2CO_3$ = 7.69 × 106 = 815(.4) g | (1) | [3] |
| (b) | (i) | 2.52 g | | [1] |
| | (ii) | Moles $Na_2CO_3$ = 0.02<br>Moles $H_2O$ = 0.14 (1) $x$ = 7 | (1)<br>(1) | [2] |
| (c) | (i) | Moles = 0.5 × 0.018 = 0.009 | | [1] |
| | (ii) | 0.0045 | | [1] |
| | (iii) | 0.0045 × 106 = 0.477 | | [1] |
| | (iv) | % = 0.477/0.55 = 86.7 % | | [1] |
| | | | | **Total [10]** |

7.

(a) percentage Be by mass = 5.03% (1)

division of percentage by $A_r$ for Be and at least one other element as shown below (1)

Al 10.04 ÷ 27 = 0.3719 → 1.00

Be 5.03 ÷ 9.01 = 0.5583 → 1.50

O 53.58 ÷ 16 = 3.3488 → 9.00

Si 31.35 ÷ 28.1 = 1.1566 → 3.10

molecular formula = $Al_2Be_3O_{18}Si_6$ or x=3 (1) **[3]**

(b) (i) Hess' Law states that where a reaction can occur by more than one route the total enthalpy **change** for each route will be the same **[1]**

(ii) $\Delta H = -393.5 - (-395.4)$ (1) $= +1.9$ kJ mol$^{-1}$ (1) **[2]**

(iii) Kyran is **incorrect** as diamond is not the **standard state** of carbon **[1]**

(iv) I mass of diamond = 7.30 g **[1]**

II mass of graphite = 7.30 ÷ (93/100) (1) = 7.85 g (1) **[2]**

**Total [10]**

8.

(a) (i) Sample is bombarded by high energy electrons / electron gun used on sample (1)
Electron knocked out (to form ions) (1) [2]

(ii) So no more than / only 1 electron is knocked out [1]

(iii) No difference (1)
Same number of electrons (in the outer shell) (1) [2]

(b) (i) $\frac{(7.25 \times 6) + (92.75 \times 7)}{100}$ (1)

6.928 (1) (accept 6.93) [2]

(ii) $^{6}Li^{+}$ since lower mass / lower m/z / lighter
do not accept 'smaller' [1]

(c) (i) $M_r(NH_4)_2SO_4 = 132.18$ (1)

Moles = 0.0156 (1) [2]

(ii) Moles LiOH = 0.0312 (1)

Concentration $= \frac{0.0312}{0.0298} = 1.05$ mol $dm^{-3}$ (1) [2]

(iii) Atom economy $= \frac{34.06}{180.08} \times 100$ (1)

$= 18.9\,\%$ (1) [2]

**Total [14]**

9.

(a) (i) I burette / (graduated) pipette [1]

II volumetric / graduated / standard flask [1]

(ii) 0.0064 [1]

(iii) 1.20 g / 100 $cm^3$ solution [1]

(iv) 12.0 g / 100 $cm^3$ solution [1]

(b) (i) The catalyst is in a different physical state to the reactants. [1]

(ii) Bonds broken 2 H-H → 872
1 C-H → 412
1 C=O → 743
1 C-O → 360
1 O-H → 463

Total +2850 kJ (1)

Bonds made 3 C-H → 1236
1 C-O → 360
3 O-H → 1389

Total -2985 kJ (1)

$\Delta H = 2850 - 2985 = -135$ kJ $mol^{-1}$ (1) [3]

(c) Relative molecular mass is a relative quantity (based on $^1/_{12}$th of the $^{12}C$ atom as one unit). [1]

(d) (i) The rate of the forward reaction is equal to the rate of the backward reaction. [1]

(ii) $C_2H_4O$ [1]

**Total [12]**

10.

(a) Weighing bottle would not have been washed / difficult to dissolve solid in volumetric flask / final volume would not necessarily be 250 $cm^3$ [1]

(b) Pipette [1]

(c) To show the end point / when to stop adding acid / when it's neutralised [1]

(d) So that a certain volume of acid can be added quickly before adding drop by drop / to save time before doing accurate titrations / to give a rough idea of the end point [1]

(e) To obtain a more reliable value [1]

(f) (i) Moles = 0.730/36.5 = 0.0200 (1)

Concentration = 0.02/0.1 = 0.200 mol $dm^{-3}$ (1) [2]

(ii) Moles = 0.2 x 0.0238 = 0.00476 [1]

(iii) 0.00476 [1]

(iv) 0.00476 x 10 = 0.0476 [1]

(v) $M_r$ = 1.14/0.0476 = 23.95 [1]

(vi) Lithium [1]

- mark consequentially throughout (f)

Total [12]

11.

(a) Filtration [1]

(b) $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ [1]

(c) (i) Carbon O.S. at start = +3; Carbon O. S. at end = +4 [1]

(ii) $2MnO_4^- + 16H^+ + 5C_2O_4^{2-} \rightarrow 2Mn^{2+} + 8H_2O + 10CO_2$ [1]

(d) Colour change of manganate(VII) is used to indicate the change [1]

(e) Volume of manganate(VII) = 27.92 $cm^3$ (1)

Moles manganate = $27.92 \times 0.020 / 1000 = 5.584 \times 10^{-4}$ mol (1)

Moles oxalate = $5.584 \times 10^{-4} \times 5/2 = 1.396 \times 10^{-3}$ mol (1)

Concentration = $1.396 \times 10^{-3} / 25 \times 10^{-3} = 0.0558$ mol $dm^{-3}$ (1) [4]

(f) (i) $K_a = \frac{[H^+][HCOO^-]}{[HCOOH]}$ [1]

(ii) $[H^+]^2 = K_a \times [HCOOH] = 1.8 \times 10^{-4} \times 0.2 = 0.36 \times 10^{-4}$ (1)

$[H^+] = 6.0 \times 10^{-3}$ mol $dm^{-3}$ (1)

$pH = -\log [H^+] = 2.22$ (1) [3]

(iii) A buffer keeps the pH almost constant when **small amounts** of acid or base are added (1)

$HCOOH \rightleftharpoons HCOO^- + H^+$ (1)

Adding acid shifts the equilibrium to the left which removes $H^+$ /
Adding base removes $H^+$ shifts equilibrium to right which replaces $H^+$ (1)
OR answer in terms of $H^+$ reacting with methanoate from
sodium methanoate when acid added (1) and methanoic acid replacing $H^+$ when base removes $H^+$ (1)

MAX 3 [3]

*QWC Selection of a form and style of writing appropriate to purpose and to complexity of subject matter* [1]

(g) (i) Orange to green [1]

(ii) $CrO_4^{2-}$ (1) Yellow (1) [2]

**Total [20]**