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About AS LEVEL

CHEMISTRY

Paper 2

(TOPICAL)

About Learning Corner

This section reveals the extra/relevant information that are not required as part of the model answers but to expand the student's knowledge in the respective fields. At the same time, it induces eagerness and interest into their learning process.

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form	Topic By Topic
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CONTENTS

AS LEVEL

Topic 1

Atoms, Molecules and Stoichiometry

1.1 Relative masses of atoms and molecules

Learning outcomes

Candidates should be able to:

1 define the unified atomic mass unit as one twelfth of the mass of a carbon-12 atom

2 define relative atomic mass, A_r, relative isotopic mass, relative molecular mass, M_r, and relative formula mass in terms of the unified atomic mass unit

1.2 The mole and the Avogadro constant

Learning outcomes

Candidates should be able to:

1 define and use the term mole in terms of the Avogadro constant

1.3 Formulae

Learning outcomes

Candidates should be able to:

1 write formulae of ionic compounds from ionic charges and oxidation numbers (shown by a Roman numeral), including:
   (a) the prediction of ionic charge from the position of an element in the Periodic Table
   (b) recall of the names and formulae for the following ions:
   NO_3^-, CO_3^{2-}, SO_4^{2-}, OH^-, NH_4^+, Zn^{2+}, Ag^+, HCO_3^-, PO_4^{3-}

2 (a) write and construct equations (which should be balanced), including ionic equations (which should not include spectator ions)
   (b) use appropriate state symbols in equations

3 define and use the terms empirical and molecular formula

4 understand and use the terms anhydrous, hydrated and water of crystallisation

5 calculate empirical and molecular formulae, using given data

1.4 Reacting masses and volumes (of solutions and gases)

Learning outcomes

Candidates should be able to:

1 perform calculations including use of the mole concept, involving:

(a) reacting masses (from formulae and equations) including percentage yield calculations

(b) volumes of gases (e.g. in the burning of hydrocarbons)

(c) volumes and concentrations of solutions

(d) limiting reagent and excess reagent

(When performing calculations, candidates' answers should reflect the number of significant figures given or asked for in the question. When rounding up or down, candidates should ensure that significant figures are neither lost unnecessarily nor used beyond what is justified (see also Mathematical requirements section).)

(e) deduce stoichiometric relationships from calculations such as those in 2.4.1 (a)–(d)

TOPIC *opening ....*

Question 1

In 1814, Sir Humphrey Davy and Michael Faraday collected samples of a flammable gas, **A**, from the ground near Florence in Italy.

They analysed **A** which they found to be a hydrocarbon. Further experiments were then carried out to determine the molecular formula of **A**.

(a) What is meant by the term *molecular formula*? [2]

Davy and Faraday deduced the formula of **A** by exploding it with an excess of oxygen and analysing the products of combustion.

(b) Complete and balance the following equation for the complete combustion of a hydrocarbon with the formula C_xH_y.

$$C_xH_y + \left(x + \frac{y}{4}\right)O_2 \rightarrow \ldots\ldots\ldots\ldots\ldots\ldots + \ldots\ldots\ldots\ldots\ldots\ldots$$

[2]

(c) When 10 cm^3 of **A** was mixed at room temperature with 50 cm^3 of oxygen (an excess) and exploded, 40 cm^3 of gas remained after cooling the apparatus to room temperature and pressure.

When this 40 cm^3 of gas was shaken with an excess of aqueous potassium hydroxide, KOH, 30 cm^3 of gas still remained.

(i) What is the identity of the 30 cm^3 of gas that remained at the end of the experiment?

(ii) The combustion of **A** produced a gas that reacted with the KOH(aq). What is the identity of this gas?

(iii) What volume of the gas you have identified in **(ii)** was produced by the combustion of **A**?

(iv) What volume of oxygen was used up in the combustion of **A**? [4]

(d) Use your equation in **(b)** and your results from **(c)(iii)** and **(c)(iv)** to calculate the molecular formula of **A**.

Show all of your working. [3]

[N10/P2/Q1]

Suggested Solution:

(a) The molecular formula specifies the actual number of atoms of each element in one molecule of a compound.

(b) $$C_xH_y + \left(x + \frac{y}{4}\right)O_2 \longrightarrow x\,CO_2 + \frac{y}{2}H_2O$$

(c) **(i)** Oxygen gas (O_2)

(ii) Carbon dioxide gas (CO_2)

(iii) 10 cm^3

(iv) 20 cm^3

(d) $C_xH_y \;+\; \left(x+\frac{y}{4}\right)O_2 \longrightarrow x\,CO_2 \;+\; \frac{y}{2}H_2O$

$10\ cm^3 \quad 20\ cm^3 \longrightarrow 10\ cm^3$

1 mol of C_xH_y gives 1 mol of CO_2, hence $x = 1$

1 mol of C_xH_y reacts with 2 mol of O_2 or $\left(x+\frac{y}{4}\right) = 2$, hence $y = 4$

$\therefore$ molecular formula of **A** = CH_4 $(CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O)$

Learning CORNER

(d) For gaseous hydrocarbons their number of moles are directly proportional to their volumes $V \propto n$, hence their volumes are directly ratioed as their moles.

Question 2

Compound **A** is an organic compound which contains carbon, hydrogen and oxygen.

When 0.240 g of the vapour of **A** is slowly passed over a large quantity of heated copper(II) oxide, CuO, the organic compound **A** is completely oxidised to carbon dioxide and water. Copper is the only other product of the reaction.

The products are collected and it is found that 0.352 g of CO_2 and 0.144 g of H_2O are formed.

(a) In this section, give your answers to three decimal places.

(i) Calculate the mass of carbon present in 0.352 g of CO_2.
Use this value to calculate the amount, in moles, of carbon atoms present in 0.240 g of **A**.

(ii) Calculate the mass of hydrogen present in 0.144 g of H_2O.
Use this value to calculate the amount, in moles, of hydrogen atoms present in 0.240 g of **A**.

(iii) Use your answers to calculate the mass of oxygen present in 0.240 g of **A**.
Use this value to calculate the amount, in moles, of oxygen atoms present in 0.240 g of **A**. [6]

(b) Use your answers to **(a)** to calculate the empirical formula of **A**. [1]

(c) When a 0.148 g sample of **A** was vapourised at 60 °C, the vapour occupied a volume of 67.7 cm^3 at a pressure of 101 kPa.

(i) Use the general gas equation $pV = nRT$ to calculate M_r of **A**.

(ii) Hence calculate the molecular formula of **A**. [3]

(d) Compound **A** is a liquid which does **not** react with 2,4-dinitrophenylhydrazine reagent or with aqueous bromine.
Suggest **two** structural formulae for **A**.

[2]

(e) Compound **A** contains only carbon, hydrogen and oxygen.
Explain how the information on the opposite page about the reaction of **A** with CuO confirms this statement. [1]

[N11/P2/Q1]

Suggested Solution:

(a) (i) Mass of carbon present in 0.352 g of CO_2

$$= \frac{\text{molecular mass of carbon in } CO_2}{M_r \text{ of } CO_2} \times (\text{mass of } CO_2 \text{ collected})$$

$$= \frac{12}{12+(16\times 2)} \times 0.352 = 0.096 \text{ g}$$

$$\text{moles of carbon present in } \mathbf{A} = \frac{\text{mass of carbon in } \mathbf{A}}{A_r \text{ of carbon}}$$

$$= \frac{0.096}{12} = 0.008 \text{ moles}$$

(ii) Mass of hydrogen present in 0.144 g of H_2O

$$= \frac{\text{molecular mass of hydrogen in } H_2O}{M_r \text{ of } H_2O} \times \text{mass of } H_2O \text{ collected}$$

$$= \frac{2\times 1}{(1\times 2)+16} \times 0.144 = 0.016 \text{ g}$$

$$\text{moles of hydrogen present in } \mathbf{A} = \frac{\text{mass of hydrogen in } \mathbf{A}}{A_r \text{ of hydrogen}}$$

$$= \frac{0.016}{1} = 0.016 \text{ moles}$$

(iii) Mass of **A** = mass of oxygen in **A** + mass of hydrogen in **A** + mass of carbon in **A**

$\Rightarrow$ 0.240 = mass of oxygen in **A** + 0.016 + 0.096

$\Rightarrow$ mass of oxygen in **A** = 0.240 − 0.016 − 0.096 = 0.128 g

$$\text{moles of oxygen present in } \mathbf{A} = \frac{\text{mass of oxygen in } \mathbf{A}}{A_r \text{ of oxygen}}$$

$$= \frac{0.128}{16} = 0.008 \text{ moles}$$

(b) Consider the number of moles of carbon, hydrogen, and oxygen present in **A**

Element	C	H	O
No. of moles	0.008	0.016	0.008
simple ratio	$\frac{0.008}{0.008}$	$\frac{0.016}{0.008}$	$\frac{0.008}{0.008}$
	1	2	1

$\therefore$ Empirical formula of **A** = CH_2O

(c) (i) $pV = nRT$

$$\Rightarrow pV = \left(\frac{m}{M_r}\right)RT \qquad (\text{since } n = \frac{m}{M_r})$$

$$\Rightarrow M_r = \frac{mRT}{pV}$$

$$= \frac{(0.148)(8.31)(273+60)}{(101\times 10^3)(67.7\times 10^{-6})}$$

$$= 59.896 \approx 59.9$$

(ii) $n = \dfrac{\text{molar mass}}{\text{empirical formula mass}}$

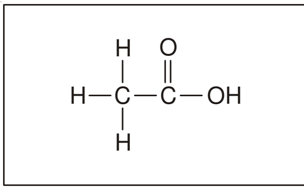
$= \dfrac{59.896}{12+2+16}$

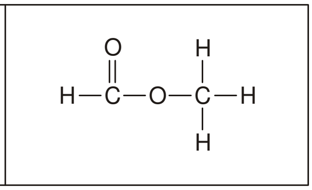
$= \dfrac{59.896}{30} = 1.996 \approx 2$

$\therefore$ Molecular formula of **A** $= C_nH_{2n}O_n$

$= C_2H_4O_2$

(d)





(e) The only products of the reaction are copper and two oxides, H_2O and CO_2. Hence compound **A** contains only carbon, hydrogen and oxygen.

Question 3

Zinc is an essential trace element which is necessary for the healthy growth of animals and plants. Zinc deficiency in humans can be easily treated by using zinc salts as dietary supplements.

(a) One salt which is used as a dietary supplement is a hydrated zinc sulfate, $ZnSO_4.xH_2O$, which is a colourless crystalline solid.

Crystals of zinc sulfate may be prepared in a school or college laboratory by reacting dilute sulfuric acid with a suitable compound of zinc.

Give the formulae of **two** simple compounds of zinc that could each react with dilute sulfuric acid to produce zinc sulfate. [2]

(b) A simple experiment to determine the value of x in the formula $ZnSO_4.xH_2O$ is to heat it carefully to drive off the water.

$$ZnSO_4.xH_2O(s) \rightarrow ZnSO_4(s) + xH_2O(g)$$

A student placed a sample of the hydrated zinc sulfate in a weighed boiling tube and reweighed it. He then heated the tube for a short time, cooled it and reweighed it when cool. This process was repeated four times. The final results are shown below.

mass of empty tube / g	mass of tube + hydrated salt / g	mass of tube + salt after fourth heating / g
74.25	77.97	76.34

(i) Why was the boiling tube heated, cooled and reweighed four times?

(ii) Calculate the amount, **in moles**, of the anhydrous salt produced.

(iii) Calculate the amount, **in moles**, of water driven off by heating.

(iv) Use your results to **(ii)** and **(iii)** to calculate the value of x in $ZnSO_4.xH_2O$. [7]

(c) For many people, an intake of approximately 15 mg per day of zinc will be sufficient to prevent deficiencies.

Zinc ethanoate crystals, $(CH_3CO_2)_2Zn.2H_2O$, may be used in this way.

(i) What mass of pure crystalline zinc ethanoate ($M_r = 219.4$) will need to be taken to obtain a dose of 15 mg of zinc?

(ii) If this dose is taken in solution as 5 cm^3 of aqueous zinc ethanoate, what would be the concentration of the solution used?
Give your answer in mol dm^{-3}. [4]

[N12/P2/Q1]

Suggested Solution:

(a) $ZnCO_3$ and $Zn(OH)_2$

(b) (i) This was done to ensure that all of the water of crystallization was driven off and a constant mass of the tube and salt was obtained.

(ii) Mass of $ZnSO_4 = 76.34 - 74.25 = 2.09\,g$

M_r of $ZnSO_4 = 65.4 + 32.1 + (4 \times 16.0) = 161.5$

$\therefore$ number of moles of anhydrous $ZnSO_4 = \dfrac{2.09}{161.5}$

$= 0.01294 \approx 1.29 \times 10^{-2}$ moles.

(iii) Mass of H_2O driven off $= 77.97 - 76.34 = 1.63$ g

M_r of $H_2O = (1 \times 2) + 16 = 18$

$\therefore$ number of moles of H_2O driven off $= \dfrac{1.63}{18}$

$= 0.09055 \approx 9.1 \times 10^{-2}$ mol.

(iv) In 1.29×10^{-2} mol of ZnSO4, number of mol of H_2O driven off $= 9.1 \times 10^{-2}$

In 1 mol of $ZnSO_4$, number of mol of H_2O driven off $= \dfrac{9.1 \times 10^{-2}}{1.29 \times 10^{-2}}$

$= 7.054 \approx 7\,mol$

Thus, the value of $x = 7$.

(c) (i) Number of moles of zinc in 15 mg of zinc $= \dfrac{0.015}{65.4}$

$= 2.294 \times 10^{-4} \approx 2.29 \times 10^{-4}$ mol.

As the number of moles of zinc = number of moles of zinc ethanoate, hence, the mass of pure crystalline zinc ethanoate required

$= 2.29 \times 10^{-4} \times 219.4$

$= 0.05024$ g

≈ 0.05 g $= 50$ mg

(ii) Volume of aqueous zinc ehanoate = 5 cm^3 = 0.005 dm^3

concentration of solution $= \dfrac{2.29 \times 10^{-4}}{0.005}$

$= 0.0458 = 4.58 \times 10^{-2}$ mol / dm^3

(a) $Zn(NO_3)_2$, ZnO and other compounds of Zn which react with sulfuric acid could also have been used.

(b) (i) If the boiling tube had been heated only once, there could have been the possibility that not all of the H_2O was driven off from the hydrated zinc sulfate. Hence the experiment would have produced wrong results.

Question 4

Chile saltpetre is a mineral found in Chile and Peru, and which mainly consists of sodium nitrate, $NaNO_3$. The mineral is purified to concentrate the $NaNO_3$ which is used as a fertiliser and in some fireworks.

In order to find the purity of a sample of sodium nitrate, the compound is heated in NaOH(aq) with Devarda's alloy which contains aluminium. This reduces the sodium nitrate to ammonia which is boiled off and then dissolved in acid.

$$3NaNO_3(aq) + 8Al(s) + 5NaOH(aq) + 18H_2O(l) \longrightarrow 3NH_3(g) + 8NaAl(OH)_4(aq)$$

The ammonia gas produced is dissolved in an excess of H_2SO_4 of known concentration.

$$2NH_3 + H_2SO_4 \longrightarrow (NH_4)_2SO_4$$

The amount of unreacted H_2SO_4 is then determined by back-titration with NaOH of known concentration.

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

(a) A 1.64 g sample of impure $NaNO_3$ was reacted with an excess of Devarda's alloy.
The NH_3 produced was dissolved in 25.0 cm^3 of 1.00 mol dm^{-3} H_2SO_4.
When all of the NH_3 had dissolved, the resulting solution was titrated with NaOH(aq).
For neutralisation, 16.2 cm^3 of 2.00 mol dm^{-3} NaOH were required.

(i) Calculate the amount, in moles, of H_2SO_4 present in the 25.0 cm^3 of 1.00 mol dm^{-3} H_2SO_4.

(ii) Calculate the amount, in moles, of NaOH present in 16.2 cm^3 of 2.00 mol dm^{-3} NaOH.

(iii) Use your answer to **(ii)** to calculate the amount, in moles, of H_2SO_4 that reacted with 16.2 cm^3 of 2.00 mol dm^{-3} NaOH.

(iv) Use your answers to **(i)** and **(iii)** to calculate the amount, in moles, of H_2SO_4 that reacted with the NH_3.

(v) Use your answer to **(iv)** to calculate the amount, in moles, of NH_3 that reacted with the H_2SO_4.

(vi) Use your answer to **(v)** to calculate the amount, in moles, of $NaNO_3$ that reacted with the Devarda's alloy.

(vii) Hence calculate the mass of $NaNO_3$ that reacted.

(viii) Use your answer to **(vii)** to calculate the percentage by mass of $NaNO_3$ present in the impure sample.
Write your answer to a suitable number of significant figures. [9]

(b) The above reaction is an example of a redox reaction.
What are the oxidation numbers of nitrogen in $NaNO_3$ and in NH_3?

$NaNO_3$ .... NH_3 .... [1]

[J13/P2/Q2]

Suggested Solution:

(a) (i) Number of moles of $H_2SO_4 = \frac{25}{1000} \times 1.00 = 0.025$ moles.

(ii) Number of moles of $NaOH = \frac{16.2}{1000} \times 2.00 = 0.0324$ moles.

(iii) NaOH and H_2SO_4 react in the molar ratio 2 : 1

$\therefore$ Number of moles of H_2SO_4 that reacted $= \frac{0.0324}{2} = 0.0162$ moles

(iv) Moles of H_2SO_4 that reacted with $NH_3 = 0.025 - 0.0162$
$= 0.0088$ moles

(v) NH_3 and H_2SO_4 react in the molar ratio 2 : 1
$\therefore$ Moles of NH_3 that reacted with $H_2SO_4 = 2 \times 0.0088$
$= 0.0176$ moles

(vi) From given equation,
3 moles of $NaNO_3$ produce 3 moles of NH_3
Hence, number of moles of $NaNO_3 = 0.0176$ moles.

(vii) Mass of $NaNO_3$ = number of moles of $NaNO_3 \times$ Mr of $NaNO_3$
$= 0.0176 \times \left(23 + 14 + (16 \times 3)\right)$
$= 0.0176 \times 85 = 1.496$ g

(viii) Percentage by mass of $NaNO_3 = \frac{1.496}{1.64} \times 100$
$= 91.2\%$

(b) $NaNO_3$: $+5$ and NH_3 : -3

(b) In $NaNO_3$ let the oxidation number of N be n
$(+1) + n + (3 \times (-2)) = 0$
$\therefore n = +5$
In NH_3, let the oxidation number of N be y
$y + (3 \times (+1)) = 0$
$\therefore y = -3$

Question 5

(a) Explain what is meant by the term *nucleon number*. [1]

(b) Bromine exists naturally as a mixture of two stable isotopes, ^{79}Br and ^{81}Br, with relative isotopic masses of 78.92 and 80.92 respectively.

(i) Define the term *relative isotopic mass*. [2]

(ii) Using the relative atomic mass of bromine, 79.90, calculate the relative isotopic abundances of ^{79}Br and ^{81}Br. [3]

(c) Bromine reacts with the element **A** to form a compound with empirical formula **A**Br_3. The percentage composition by mass of **A**Br_3 is **A**, 4.31; Br, 95.69.
Calculate the relative atomic mass, A_r, of **A**.
Give your answer to **three** significant figures. [3]

[J14/P2/Q1 (a,b,c)]

Suggested Solution:

(a) The nucleon number is the total number of protons and neutrons in the nucleus of an atom.

(b) (i) The relative isotopic mass of an isotope is the mass of an atom of the isotope compared to $\frac{1}{12}$ of the mass of a carbon-12 atom.

(ii) Let the percentage abundance of ^{79}Br be $x\%$

$\Rightarrow$ percentage abundance of ^{81}Br $= (100 - x)\%$

$$\therefore \quad (78.92 \times \frac{x}{100}) + (80.92 \times \frac{100 - x}{100}) = 79.9$$

$$0.7892x + 0.8092(100 - x) = 79.9$$

$$0.7892x + 80.92 - 0.8092x = 79.9$$

$$x = 51$$

$\therefore$ relative isotopic abundance of ^{79}Br $= 51\%$

relative isotopic abundance of ^{81}Br $= 49\%$

(c)

Element	**A**		Br
Mass	4.31		95.69
Molar Ratio	$\frac{4.31}{A_r}$	:	$\frac{95.69}{79.9}$

As the empirical formula of the compound is **A**Br_3, the molar ratio must be equal to 1 : 3.

$$\frac{4.31}{A_r} : \frac{95.69}{79.9} = 1 : 3$$

$$\Rightarrow \quad \frac{4.31}{A_r} = \frac{1}{3} \times \frac{95.69}{79.9} \quad \Rightarrow \quad \frac{4.31}{A_r} = 0.39921 \quad \Rightarrow \quad A_r = 10.796 \approx 10.8$$

Question 6

A 6.30 g sample of hydrated ethanedioic acid, $H_2C_2O_4.xH_2O$, was dissolved in water and the solution made up to 250 cm³.

A 25.0 cm³ sample of this solution was acidified and titrated with 0.100 mol dm⁻³ potassium manganate(VII) solution. 20.0 cm³ of this potassium manganate(VII) solution was required to react fully with the ethanedioate ions, $C_2O_4^{2-}$, present in the sample.

(a) The MnO_4^- ions in the potassium manganate(VII) *oxidise* the ethanedioate ions.

(i) Explain, in terms of electron transfer, the meaning of the term *oxidise* in the sentence above. [1]

(ii) Complete and balance the ionic equation for the reaction between the manganate(VII) ions and the ethanedioate ions.

$$2MnO_4^-(aq) + 5C_2O_4^{2-}(aq) + \text{.........}H^+(aq) \rightarrow \text{.........}(aq) + 10CO_2(aq) + \text{.........}H_2O(l)$$

[3]

(b) (i) Calculate the number of moles of manganate(VII) used in the titration. [1]

(ii) Use the equation in **(a)(ii)** and your answer to **(b)(i)** to calculate the number of moles of $C_2O_4^{2-}$ present in the 25.0 cm³ sample of solution used. [1]

(iii) Calculate the number of moles of $H_2C_2O_4.xH_2O$ in 6.30 g of the compound. [1]

(iv) Calculate the relative formula mass of $H_2C_2O_4.xH_2O$. [1]

(v) The relative formula mass of anhydrous ethanedioic acid, $H_2C_2O_4$, is 90. Calculate the value of x in $H_2C_2O_4.xH_2O$. [1]

[J14/P2/Q2]

Suggested Solution:

(a) (i) Oxidation is the loss of electrons. In the given sentence, the term oxidise means that ethanedioate ions lose electrons.

(ii) $2MnO_4^-(aq) + 5C_2O_4^{2-}(aq) + \ldots\mathbf{16}\ldots H^+(aq) \rightarrow \mathbf{2Mgn^{2+}}\ldots(aq) + 10CO_2(aq) + \ldots\mathbf{8}\ldots H_2O(l)$

(b) (i) Moles of manganate (VII) $= 0.1 \times \frac{20}{1000} = 2 \times 10^{-3}$ moles.

(ii) MnO_4^- and $C_2O_4^{2-}$ are in the molar ratio $2:5$

$\therefore$ moles of $C_2O_4^{2-} = \frac{5}{2} \times 2 \times 10^{-3} = 5 \times 10^{-3}$ moles.

(iii) Number of moles $= 5 \times 10^{-3} \times \frac{250}{25} = 0.05$ moles

(iv) Relative formula mass, $M_r = \frac{\text{mass in grams}}{\text{number of moles}} = \frac{6.30}{0.05} = 126.$

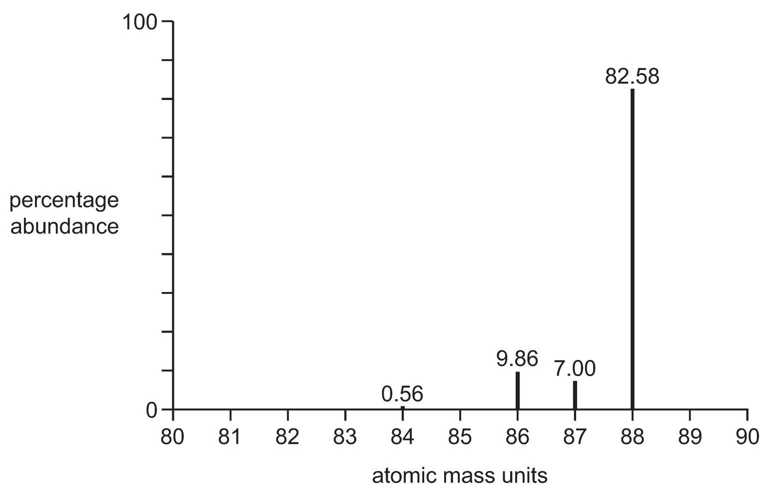
(v) M_r of $H_2C_2O_4 = 90$, M_r of $H_2O = 18$

$\therefore \quad 90 + 18x = 126$

$18x = 36 \quad \Rightarrow \quad x = 2$

Question 7

(b) A sample of strontium, atomic number 38, gave the mass spectrum shown. The percentage abundances are given above each peak.



(i) Complete the full electronic configuration of strontium.
$1s^2\ 2s^2\ 2p^6$ [1]

(ii) Explain why there are four different peaks in the mass spectrum of strontium. [1]

(iii) Calculate the atomic mass, A_r, of this sample of strontium.
Give your answer to **three** significant figures. [2]

(c) A compound of barium, **A**, is used in fireworks as an oxidising agent and to produce a green colour.

(i) Explain, in terms of electron transfer, what is meant by the term *oxidising agent*. [1]

(ii) **A** has the following percentage composition by mass:
Ba, 45.1; C*l*, 23.4; O, 31.5. Calculate the empirical formula of **A**. [3]

[N14/P2/Q1 (b,c)]

Suggested Solution:

(b) (i) $1s^2\ 2s^2\ 2p^6$ $\mathbf{3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6\ 5s^2}$

(ii) Because strontium has four isotopes.

(iii) $A_r = (84\times\frac{0.56}{100}) + (86\times\frac{9.86}{100}) + (87\times\frac{7}{100}) + (88\times\frac{82.58}{100})$

$= 0.4704 + 8.4796 + 6.09 + 72.6704 = 87.7104 \approx 87.7$

(c) (i) An oxidizing agent is a substance that gains electrons in a chemical reaction.

(ii)

element	Ba	C*l*	O
% by mass	45.1	23.4	31.5
Ar	137	35.5	16
no. of moles =	$\frac{45.1}{137}$	$\frac{23.4}{35.5}$	$\frac{31.5}{16}$
=	0.329	0.659	1.968
simple ratio =	$\frac{0.329}{0.329}$	$\frac{0.659}{0.329}$	$\frac{1.968}{0.329}$
=	1	2	6

empirical formula of **A** $= BaCl_2O_6$

Question 8

Acidified potassium dichromate(VI) can oxidise ethanedioic acid, $H_2C_2O_4$. The relevant half-equations are shown.

$$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$$

$$H_2C_2O_4 \rightarrow 2CO_2 + 2H^+ + 2e^-$$

(a) State the overall equation for the reaction between acidified dichromate(VI) ions and ethanedioic acid. [2]

(b) In an experiment a 0.242 g sample of hydrated ethanedioic acid, $H_2C_2O_4 \cdot xH_2O$, was reacted with a 0.0200 mol dm^{-3} solution of acidified potassium dichromate(VI).

32.0 cm^3 of the acidified potassium dichromate(VI) solution was required for complete oxidation of the ethanedioic acid.

(i) Calculate the amount, in moles, of dichromate(VI) ions used to react with the sample of ethanedioic acid. [1]

(ii) Calculate the amount, in moles, of ethanedioic acid in the sample. [1]

(iii) Calculate the relative molecular mass, *Mr*, of the hydrated ethanedioic acid. [1]

(iv) Calculate the value of ***x*** in $H_2C_2O_4 \cdot xH_2O$. [1]

[J16/P2/Q3]

Suggested Solution:

(a) $Cr_2O_7^{2-} + 8H^+ + 3H_2C_2O_4 \longrightarrow 2Cr^{3+} + 6CO_2 + 7H_2O$

(b) (i) No. of moles of dichromate(VI) ions, $n = c \times v$

$$= 0.02 \times \frac{32}{1000} = 6.40 \times 10^{-4} \text{ mol.}$$

(ii) $Cr_2O_7^{2-}$ reacts with $H_2C_2O_4$ in the ratio $1:3$

$\therefore$ number of moles of $H_2C_2O_4 = 3 \times 6.4 \times 10^{-4}$

$$= 1.92 \times 10^{-3} \text{ mol}$$

(iii) $M_r = \dfrac{\text{mass}}{\text{moles}}$

$$= \frac{0.242}{0.00192} = 126$$

(iv) $2 + (12 \times 2) + (16 \times 4) + x(2 + 16) = 126$

$$2 + 24 + 64 + 18x = 126$$

$$18x = 36$$

$$x = 2$$

Question 9

A 0.50 g sample of a Group 2 metal, **M**, was added to 40.0 cm^3 of 1.00 mol dm^{-3} hydrochloric acid (an excess).

equation 1 $\quad \mathbf{M}(s) + 2HCl(aq) \rightarrow \mathbf{M}Cl_2(aq) + H_2(g)$

(a) Calculate the amount, in moles, of hydrochloric acid present in 40.0 cm^3 of 1.00 mol dm^{-3} HCl. [1]

(b) When the reaction had finished, the resulting solution was made up to 100 cm^3 in a volumetric flask.

A 10.0 cm^3 sample of the solution from the volumetric flask required 15.0 cm^3 of 0.050 mol dm^{-3} sodium carbonate solution, Na_2CO_3, for complete neutralisation of the remaining hydrochloric acid.

(i) Write the equation for the complete reaction of sodium carbonate with hydrochloric acid. [1]

(ii) Calculate the amount, in moles, of sodium carbonate needed to react with the hydrochloric acid in the 10.0 cm^3 sample from the volumetric flask. [1]

(iii) Calculate the amount, in moles, of hydrochloric acid in the 10.0 cm^3 sample. [1]

(iv) Calculate the total amount, in moles, of hydrochloric acid remaining after the reaction shown in equation 1. [1]

(v) Use your answers to **(a)** and **(b)(iv)** to calculate the amount, in moles, of hydrochloric acid that reacted with the 0.50 g sample of **M**. [1]

(vi) Use your answer to **(v)** and equation 1 to calculate the amount, in moles, of **M** in the 0.50 g sample. [1]

(vii) Calculate the relative atomic mass, A_r, of **M** and identify **M**. [2]

[N16/P2/Q1]

Suggested Solution:

(a) Number of moles $= \frac{40}{1000} \times 1.0 = 0.04$ mol.

(b) (i) $Na_2CO_3 + 2HCl \longrightarrow NaCl + H_2O + CO_2$

(ii) Number of moles of sodium carbonate $= \frac{15}{1000} \times 0.05 = 0.00075$ mol

(iii) 1 mole of Na_2CO_3 reacts with 2 moles of HCl
0.00075 moles of Na_2CO_3 reacts with $2 \times 0.00075 = 0.0015$ moles of HCl

(iv) By ratio method,

volume	:	moles
10.0 cm^3	:	0.0015
100.0 cm^3	:	x

$\Rightarrow \quad x = 0.0015 \times \frac{100.0}{10.0} = 0.015$

$\therefore$ number of moles $= 0.015$ mol

(v) Amount $= 0.04 - 0.015$
$= 0.025$ mol

(vi) Using ratio method,

M	:	HCl
1	:	2
x	:	0.025

$\Rightarrow \quad x = \frac{0.025}{2} = 0.0125$ mol

(vii) A_r of **M** $= \frac{\text{mass}}{\text{moles}}$
$= \frac{0.50}{0.0125} = 40$

Identity of **M** = Calcium

Question 10

Ammonium iron(II) sulfate, $(NH_4)_2Fe(SO_4)_2$, has a relative formula mass, M_r, of 284.

(a) Define the term *relative formula mass*. [2]

(b) One of the cations in ammonium iron(II) sulfate is the ammonium ion, NH_4^+.

(i) Draw a 'dot-and-cross' diagram of an ammonium ion. Show outer shell electrons only.
Use × to show electrons from nitrogen.
Use • to show electrons from hydrogen. [2]

(ii) Suggest the shape of an ammonium ion and predict the bond angle. [2]

(c) In aqueous solution the ammonium ion acts as a weak Brønsted-Lowry acid.

(i) Explain the meaning of the term *weak Brønsted-Lowry acid.* [2]

(ii) Write an equation to show this behaviour of the ammonium ion in water. Include state symbols. [2]

(d) Mohr's salt, $(NH_4)_2Fe(SO_4)_2 \cdot xH_2O$, is the hydrated form of ammonium iron(II) sulfate.

x represents the number of moles of water in 1 mole of the salt.

A student wanted to determine the value of x. 0.784 g of the hydrated salt was dissolved in water and this solution was acidified.

All of the solution was titrated with 0.0200 mol dm^{-3} potassium manganate (VII). 20.0 cm^3 of this potassium manganate(VII) solution was required for complete reaction with the Fe^{2+} ions.

(i) Use changes in oxidation numbers to balance the equation for the reaction taking place.

$MnO_4^-(aq) + ...Fe^{2+}(aq) + ...H^+(aq) \rightarrow ....Mn^{2+}(aq) + ...Fe^{3+}(aq) + ...H_2O(l)$

[1]

(ii) State the role of the Fe^{2+} ions in this reaction. Explain your answer. [2]

(iii) Calculate the amount, in moles, of manganate(VII) ions that reacted. [1]

(iv) Calculate the amount, in moles, of Fe^{2+} ions in the sample of the salt. [1]

(v) Calculate the relative formula mass of $(NH_4)_2Fe(SO_4)_2 \cdot xH_2O$. [1]

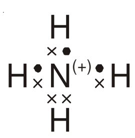
(vi) Calculate the value of x. [1]

[J18/P2/Q2]

Suggested Solution:

(a) The relative formula mass is the mass of 1 mol of a compound, which accounts for the relative masses of all the elements it consists of. These masses are, however, relative to the mass of 1/12th of the carbon-12 atom.

(b) (i)



(ii) Shape: Tetrahedral
Bond angle: 109 – 109.5°

(c) (i) A weak Bronsted-Lowry acid only partially dissociates, donating a hydrogen ion.

(ii) $NH_4^+(aq) + H_2O(l) \rightleftharpoons NH_3\,(aq\ or\ g) + H_3O^+\,(aq)$

(d) (i) $MnO_4^-(aq) + \underline{5}\ Fe^{2+}(aq) + \underline{8}\ H^+(aq) \rightarrow \underline{1}\ Mn^{2+}(aq) + \underline{5}\ Fe^{3+}(aq) + \underline{4}\ H_2O(l)$

(ii) Since Fe^{2+} loses an electron, it acts as a reducing agent.

(iii) Amount in moles = $0.0200\ mol/dm^3 \times \frac{20}{1000}\ dm^3 = 0.0004$ moles

(b) (ii) The tetrahedral shape of ammonium ion is due to sp3 hybridization.

(c) (i) Examples of Bronsted-Lowry Acids include:

HCl, H_2SO_4, HNO_3

(d) (iii) Number of moles = concentration (mol / dm^3) × volume (dm^3).

(iv) Ratio between MnO_4^- and $Fe^{2+} = 1:5$

$\therefore$ Amount in moles of $Fe^{2+} = 0.0004 \times 5 = 0.002$ moles.

(v) Relative formula mass, $M_r = \dfrac{0.784}{0.002} = 392$

(d) (v) $M_r = \dfrac{\text{Given mass}}{\text{Moles}}$

(vi) M_r of salt = 392

M_r of $(NH_4)_2Fe(SO_4)_2 = 2[14 + (1 \times 4)] + 56 + 2[32 + (16 \times 4)]$
$= 36 + 56 + 192 = 284$

Difference = 392 − 284 = 108

M_r of $H_2O = 18$

$\therefore \quad x = \dfrac{108}{18} = 6$

Question 11

The model of the nuclear atom was first proposed by Ernest Rutherford. He developed this model on the basis of results obtained from an experiment using gold metal foil.

(a) Complete the table with information for two of the particles in an atom of ^{197}Au.

particle	relative mass	relative charge	location within atom	total number in an atom of ^{197}Au
electron	0.0005	–1		79
neutron			nucleus	

[4]

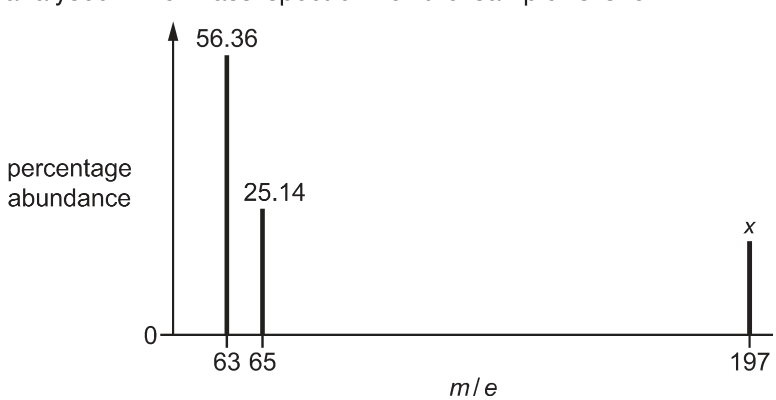
(b) State the type of bonding in gold. [1]

(c) A sample of gold found in the earth consists of only one isotope.

(i) Explain what is meant by the term *isotopes*. [2]

(ii) A different sample of gold contains more than one isotope.
Suggest why this different sample of gold has the same **chemical** properties as the sample found in the earth. [1]

(d) *Tumbaga* is an alloy of copper and gold. A sample of tumbaga was analysed. The mass spectrum of the sample is shown.



(i) Calculate the percentage abundance of gold, x, in the sample of tumbaga. [1]

(ii) Calculate the relative atomic mass, A_r, of the copper present in this sample.

Give your answer to **two** decimal places. [2]

[N18/P2/Q1]

Suggested Solution:

(a)

particle	relative mass	relative charge	location within atom	total number in an atom of ^{197}Au
electron	0.0005	–1	**Shell**	79
neutron	**1**	**0**	nucleus	**118**

(b) Metallic bonding.

(c) (i) Isotopes are defined as atoms of the same element with same number of protons but different number of neutrons.

(ii) Chemical properties are controlled by the electronic structure of an atom. Since, isotopes have equal number of electrons, they have similar chemical properties.

(d) (i) $56.36\% + 25.14\% + x\% = 100\%$

$x = 100 - 56.36 - 25.14$

$x = 18.5\%$

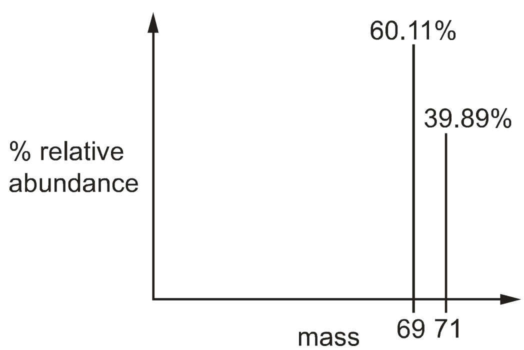
(ii) Relative atomic mass, $A_r = \dfrac{(56.36\% \times 63) + (25.14 \times 65)}{56.36 + 25.14}$

$= \dfrac{3550.68 + 1634.1}{81.5} = 63.62$

Question 12

Gallium is an element in Group 13.

A sample of gallium is analysed using a mass spectrometer. The mass spectrum produced is shown.



(a) Explain what is meant by the term *relative atomic mass*. [2]

(b) Calculate the relative atomic mass of gallium in this sample. Give your answer to 4 significant figures.

Show your working. [2]

(c) Complete the table which describes a gaseous atom of gallium.

isotope	nucleon number	total number of electrons in **lowest** energy level	type of orbital which contains the electron in the **highest** energy level
^{71}Ga			

[3]

(d) When gallium is heated in excess chlorine, gallium trichloride, $GaCl_3$, is made.

Draw the shape of the gallium trichloride molecule and suggest the Cl–Ga–Cl bond angle. [2]

(e) Gallium oxide, Ga_2O_3, and aluminium oxide react in the same way with HCl(aq) and with NaOH(aq).

(i) Suggest the equation for the reaction between Ga_2O_3 and HCl(aq). [1]

(ii) Suggest an equation for the reaction between gallium oxide and NaOH(aq). [2]

[J20/P2/Q1]

Suggested Solution:

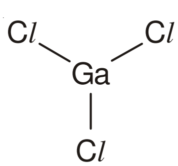
(a) Relative atomic mass is the average mass of all the isotopes of an atom relative to 1/12th of the mass of an atom of ^{12}C, which has a mass of exactly 12.

(b) Relative atomic mass of Gallium $= \left(\frac{60.11}{100} \times 69\right) + \left(\frac{39.89}{100} \times 71\right)$

$= 69.80$

(c)

isotope	nucleon number	total number of electrons in **lowest** energy level	type of orbital which contains the electron in the **highest** energy level
^{71}Ga	**71**	**2**	**p-orbital**

(d) 

Bond angle = 120°.

(e) **(i)** $Ga_2O_3 + 6HCl \rightarrow 2GaCl_3 + 3H_2O$

(ii) $Ga_2O_3 + 2NaOH + 3H_2O \rightarrow 2NaGa(OH)_4$

(c) Gallium lies in Group III and period 4.

(d) The molecule has 24 valence electrons with no lone pairs so the shape of the compound will be Trigonal planar.

(e) (ii) With NaOH, an alternate reaction can be:

$Ga_2O_3 + 2NaOH \rightarrow 2NaGaO_2 + H_2O$

Topic 5

Chemical Energetics

5.1 Enthalpy change, ΔH

Learning outcomes

Candidates should be able to:

1 understand that chemical reactions are accompanied by enthalpy changes and these changes can be exothermic (ΔH is negative) or endothermic (ΔH is positive)

2 construct and interpret a reaction pathway diagram, in terms of the enthalpy change of the reaction and of the activation energy

3 define and use the terms:

(a) standard conditions (this syllabus assumes that these are 298 K and 101 kPa) shown by $^{\ominus}$.

(b) enthalpy change with particular reference to: reaction, ΔH_r , formation, ΔH_f , combustion, ΔH_c , neutralisation, ΔH_{neut}

4 understand that energy transfers occur during chemical reactions because of the breaking and making of chemical bonds

5 use bond energies ((ΔH positive, i.e. bond breaking) to calculate enthalpy change of reaction, ΔH_r

6 understand that some bond energies are exact and some bond energies are averages

7 calculate enthalpy changes from appropriate experimental results, including the use of the relationships q = $mc\Delta T$ and $\Delta H = -mc\Delta T/n$

5.2 Hess's Law

Learning outcomes

Candidates should be able to:

1 apply Hess's Law to construct simple energy cycles

2 carry out calculations using cycles and relevant energy terms, including:

(a) determining enthalpy changes that cannot be found by direct experiment

(b) use of bond energy data

TOPIC *opening ....*

Question 1

Halogenoalkanes have been widely used as aerosol propellants, refrigerants and solvents for many years.

Fluoroethane, CH_3CH_2F, has been used as a refrigerant. It may be made by reacting ethene with hydrogen fluoride.

You are to calculate a value for the C—F bond energy in fluoroethane.

(a) Use relevant bond energies from the *Data Booklet*, and the equation below to calculate a value for the bond energy of the C—F bond.

$$CH_2 = CH_2(g) \quad + \quad HF(g) \quad \longrightarrow \quad CH_3CH_2F(g) \qquad \Delta H^{\ominus} = -73\,kJmol^{-1}$$

[4]

(b) Another halogenoalkane which was used as a refrigerant, and also as an aerosol propellant, is dichlorodifluoromethane, CCl_2F_2.

State **two** reasons why compounds such as CH_3CH_2F and CCl_2F_2 have been used as aerosol propellants and refrigerants. [2]

CCl_2F_2 is one of many chlorofluorocarbon compounds responsible for damage to the ozone layer in the stratosphere.

(c) By using relevant data from the *Data Booklet*, and your answer to **(a)** suggest why CCl_2F_2 is responsible for damage to the ozone layer in the stratosphere whereas CH_3CH_2F is not. [2]

Both CH_3CH_2F and CCl_2F_2 are greenhouse gases.

The 'enhanced greenhouse effect' is of great concern to the international community.

(d) (i) What is meant by the term *enhanced greenhouse effect*?

(ii) Water vapour is the most abundant greenhouse gas.
What is the second most abundant greenhouse gas? [3]

A greenhouse gas which is present in very small amounts in the atmosphere is sulfur hexafluoride, SF_6, which is used in high voltage electrical switchgear.

(e) What shape is the SF_6 molecule? [1]

[J11/P2/Q2]

Suggested Solution:

(a) Bonds broken / Bonds formed

Bonds broken	Bonds formed
4 C—H	5 C—H
1 C═C	1 C—F
1 HF	1 C—C

Let bond energy of C—F bond be x.

$\therefore$ Energy required to break bonds $= (4\times410)+(1\times610)+(1\times562) = 2812$
Energy required to form bonds $= (5\times410)+(1\times x)+(1\times350) = 2400 + x$

$$\Delta H^{\ominus}_{\text{reaction}} = 2812-(2400+x)$$
$$-73 = 2812-2400-x$$
$$x = 2812-2400+73$$
$$= +485\ \text{kJmol}^{-1}$$

(b) 1. They are non-flammable.
2. They are non-toxic.

(b) *Alternative reasons:*
These compounds are
— unreactive
— volatile
— easily liquified.

(c) CCl_2F_2 contains C—C*l* bonds which have a bond energy of 340 kJ/mol. From part **(a)**, the bond energy of C—F bond is 485 kJ/mol. Hence, C—F bond is stronger than the C—C*l* bond. The C—H bond, having a bond energy of 410 kJ/mol, is also stronger than the C—C*l* bond.
Therefore C—C*l* bonds are broken by uv light, producing Cl^{-} free radicals, which are responsible for damaging the ozone.

(d) **(i)** The term *enhanced greenhouse effect* refers to the trapping of reflected heat from the earth in the lower atmosphere, which contributes to global warming.

(ii) Carbon dioxide.

(e) Octahedral

Question 2

For some chemical reactions, such as the thermal decomposition of potassium hydrogencarbonate, $KHCO_3$, the enthalpy change of reaction cannot be measured directly.
In such cases, the use of Hess' Law enables the enthalpy change of reaction to be calculated from the enthalpy changes of other reactions.

(a) State Hess' Law. [2]

In order to determine the enthalpy change for the thermal decomposition of potassium hydrogencarbonate, two separate experiments were carried out.

experiment 1

30.0 cm^3 of 2.00 mol dm^{-3} hydrochloric acid (an excess) was placed in a conical flask and the temperature recorded as 21.0 °C.
When 0.0200 mol of potassium carbonate, K_2CO_3, was added to the acid and the mixture stirred with a thermometer, the maximum temperature recorded was 26.2 °C.

(b) **(i)** Construct a balanced equation for this reaction.

(ii) Calculate the quantity of heat produced in **experiment 1**, stating your units.
Use relevant data from the *Data Booklet* and assume that all solutions have the same specific heat capacity as water.

(iii) Use your answer to **(ii)** to calculate the enthalpy change per mole of K_2CO_3.
Give your answer in $kJ\ mol^{-1}$ and include a sign in your answer.

(iv) Explain why the hydrochloric acid must be in an excess. [4]

experiment 2

The experiment was repeated with 0.0200 mol of potassium hydrogencarbonate, $KHCO_3$.
All other conditions were the same.
In the second experiment, the temperature fell from 21.0 °C to 17.3 °C.

(c) **(i)** Construct a balanced equation for this reaction.

(ii) Calculate the quantity of heat absorbed in **experiment 2**.

(iii) Use your answer to **(ii)** to calculate the enthalpy change per mole of $KHCO_3$.
Give your answer in $kJ\ mol^{-1}$ and include a sign in your answer. [3]

(d) When $KHCO_3$ is heated, it decomposes into K_2CO_3, CO_2 and H_2O.

$$2KHCO_3 \rightarrow K_2CO_3 + CO_2 + H_2O$$

Use Hess' Law and your answers to **(b)(iii)** and **(c)(iii)** to calculate the enthalpy change for this reaction.
Give your answer in $kJ\ mol^{-1}$ and include a sign in your answer. [2]

[N11/P2/Q3]

Suggested Solution:

(a) The overall enthalpy change for a reaction is independent of the number of steps that are involved in the reaction, provided that the initial and final conditions are kept the same.

(b) **(i)** $K_2CO_3 + 2HCl \rightarrow 2KCl + H_2O + CO_2$

(ii) $V_{HCl} = 30.0\ cm^3$

$\therefore$ Heat produced, $Q = mc\Delta T$

$= 30 \times 4.18 \times (26.2 - 21)$

$= 652.08$ J per 0.0200 moles of K_2CO_3

(iii) Heat produced by 0.02 moles of $K_2CO_3 = 652.08$ J

heat produced by 1 mole of $K_2CO_3 = \dfrac{652.08}{0.02}$

$= 32604\ J \approx 32.60\ kJ$

$\therefore$ enthalpy change per mole of $K_2CO_3 = -32.60\ kJ/mol$

(b) (iii) Negative sign indicates that reaction is exothermic and heat is released into the environment.

(iv) This is done to ensure that all of the K_2CO_3 is completely neutralised.

(c) **(i)** $KHCO_3 + HCl \rightarrow KCl + H_2O + CO_2$

(ii) $V_{HCl} = 30.0\ cm^3$

$\therefore$ Heat absorbed, $Q = mc\Delta T$

$= 30 \times 4.18 \times (21 - 17.3)$

$= 30 \times 4.18 \times 3.7$

$= 463.98$ J per 0.0200 moles of $KHCO_3$

Learning CORNER

(iii) 0.02 moles of $KHCO_3$ cause the absorption of 463.98 J of heat

$\therefore$ Heat absorbed per mole of $KHCO_3 = \dfrac{1}{0.02} \times 463.98$

$= 23199\,J \approx 23.20\ kJ$

Enthalpy change per mole of $KHCO_3 = +23.20\,kJ/mol$

(c) (iii) Positive sign indicates that reaction is endothermic and heat is absorbed during the reaction.

(d)

$$\begin{array}{ccc} 2KHCO_3 & \xrightarrow{\Delta H_{reaction}} & K_2CO_3 + CO_2 + H_2O \\ & \mathbf{A}\searrow \qquad \swarrow\mathbf{B} & \\ & KCl + H_2O + CO_2 & \end{array}$$

Using Hess's law, $\Delta H_{reaction} + \Delta H_B = \Delta H_A$

$\Rightarrow \Delta H_{reaction} = \Delta H_A - \Delta H_B$

$\therefore = (2 \times +23.20) - (-32.60)$

$= 46.40 + 32.60 = +79.00\ kJ/mol$

Question 3

Alcohols such as methanol, CH_3OH, are considered to be possible replacements for fossil fuels because they can be used in car engines.

(a) Define, with the aid of an equation which includes state symbols, the standard enthalpy change of combustion, $\Delta H_c^{\ominus}$, for methanol at 298 K. [3]

Methanol may be synthesised from carbon monoxide and hydrogen. Relevant $\Delta H_c^{\ominus}$ values for this reaction are given in the table below.

compound	$\Delta H_c^{\ominus}$ / KJ mol^{-1}
CO(g)	–283
H_2(g)	–286
CH_3OH(g)	–726

(b) Use these values to calculate $\Delta H^{\ominus}_{reaction}$ for the synthesis of methanol, using the following equation. Include a sign in your answer.

$$CO(g) + 2H_2(g) \longrightarrow CH_3OH(g) \qquad [3]$$

(c) The operating conditions for this reaction are as follows.

pressure 200 atmospheres (2×10^7 Pa)

temperature 600 K

catalyst oxides of Cr, Cu, and Zn

Explain how **each** of these conditions affects the **rate of formation** of methanol. [6]

[J12/P2/Q2]

Suggested Solution:

(a) equation: $CH_3OH(l) + \frac{3}{2}O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$

definition: Standard enthalpy change of combustion of a substance is the heat evolved when 1 mole of a substance, at its standard state, is completely burned in oxygen under standard conditions of 1 atm and 298 K.

(b)

$$\begin{tikzcd} CO(g) + 2H_2(g) \arrow[rr, "\Delta H^{\ominus}_{\text{reaction}}"] \arrow[dr] & & CH_3OH(g) \arrow[dl] \\ & CO_2(g) + H_2O(l) & \end{tikzcd}$$

$$\Delta H^{\ominus}_{\text{reaction}} + (-726) = (-283) + (-286 \times 2)$$

$$\Delta H^{\ominus}_{\text{reaction}} = -283 - 572 + 726$$

$$= -129 \text{ kJ/mol}$$

(c) pressure: Increasing the pressure will increase the rate of formation of methanol, due to the fact that the frequency of collisions will increase, thereby leading to more fruitful collisions.

temperature: Increasing the temperature will increase methanol's rate of formation as more molecules would have energy greater than the activation energy, and thus more successful collisions will take place.

catalyst: The use of a catalyst would increase the rate of formation of methanol by providing an alternative route of lower activation energy to the reaction pathway.

Learning CORNER

(c) Pressure: Increasing the pressure will increase the concentration of the reactants i.e. there would be more moles of reactant particles per unit volume.

Temperature: Increasing the temperature would cause the reactant molecules to have more energy. The molecules having Energy greater than E_a could then take part in the reaction.

Catalyst: A catalyst always lowers the E_a by providing an alternate chemical pathway.

Question 4

Propane, C_3H_8, and butane, C_4H_{10}, are components of Liquefied Petroleum Gas (LPG) which is widely used as a fuel for domestic cooking and heating.

(a) (i) To which class of compounds do these two hydrocarbons belong?

(ii) Write a balanced equation for the complete combustion of butane. [2]

(b) When propane or butane is used in cooking, the saucepan may become covered by a solid black deposit.

(i) What is the chemical name for this black solid?

(ii) Write a balanced equation for its formation from butane. [2]

(c) Propane and butane have different values of standard enthalpy change of combustion.

Define the term *standard enthalpy change of combustion*. [2]

(d) A 125 cm^3 sample of propane gas, measured at 20 °C and 101 kPa, was completely burnt in air.
The heat produced raised the temperature of 200 g of water by 13.8 °C. Assume no heat losses occurred during this experiment.

(i) Use the equation $pV = nRT$ to calculate the mass of propane used.

(ii) Use relevant data from the *Data Booklet* to calculate the amount of heat released in this experiment.

(iii) Use the data above and your answers to **(i)** and **(ii)** to calculate the energy produced by the burning of 1 mol of propane. [5]

(e) The boiling points of methane, ethane, propane, and butane are given below.

compound	CH_4	CH_3CH_3	$CH_3CH_2CH_3$	$CH_3(CH_2)_2CH_3$
boiling point / K	112	185	231	273

(i) Suggest an explanation for the increase in boiling points from methane to butane.

(ii) The isomer of butane, 2-methylpropane, $(CH_3)_3CH$, has a boiling point of 261 K.
Suggest an explanation for the difference between this value and that for butane in the table above. [4]

[N13/P2/Q5]

Suggested Solution:

(a) (i) Alkanes.

(ii) $2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O$

(b) (i) Carbon

(ii) $2C_4H_{10} + 5O_2 \rightarrow 8C + 10H_2O$

(c) The enthalpy change when 1 mole of a substance is completely burnt in an excess of oxygen under standard conditions of 298 K and 1 atm.

(d) (i) $pV = nRT \quad \Rightarrow \quad n = \dfrac{pV}{RT}$

as $\quad n = \dfrac{\text{mass in grams}}{M_r}$

$\Rightarrow \quad \dfrac{\text{mass of propane used}}{M_r \text{ of propane}} = \dfrac{pV}{RT}$

$\Rightarrow \quad \text{mass of propane used} = \dfrac{pVM_r}{RT}$

$$= \frac{(101\times10^3)\times(125\times10^{-6})\times\left((12\times3)+(1\times8)\right)}{8.31\times(273.15+20)}$$

$$= \frac{555.5}{2436.08}$$

$$= 0.228 \approx 0.23 \text{ g}$$

(ii) Heat Released $= mc\Delta T$

$= 200\times4.18\times13.8$

$= 11\,536.8 \text{ J} \approx 11.5 \text{ kJ}$

(iii) Moles of propane $= \dfrac{0.228}{44} = 0.005182$

energy produced by 0.00518 moles of propane = 11536.8 J

$\therefore$ energy produced by 1 mole of propane $= \dfrac{11536.8}{0.005182}$

$= 2226321.9 \text{ J mol}^{-1}$

$= 2226.3219 \text{ kJ mol}^{-}$

$\approx 2200 \text{ kJ mol}^{-1}$

(e) (i) From methane to butane, there are more number of electrons in the molecule. This causes more stronger Vander Waal's forces of attraction to develop between the molecules leading to an increase in boiling points.

(ii) Straight Chain molecules such as $CH_3CH_2CH_2CH_3$, can pack more closely and therefore develop stronger Vander Waal's forces of attraction between molecules. This causes straight chain molecules to have a higher boiling point than branched molecules.

Question 5

Chemical reactions are accompanied by enthalpy changes.

(a) Explain the meaning of the term *standard enthalpy change* of reaction. [2]

(b) The enthalpy change of hydration of anhydrous magnesium sulfate, ΔH_{hyd} $MgSO_4$, can be calculated by carrying out two separate experiments.

In the first experiment 45.00 g of water was weighed into a polystyrene cup and 3.01 g of $MgSO_4$ was added and stirred until it was completely dissolved. The temperature of the water rose from 23.4 °C to 34.7 °C.

(i) Calculate the amount of heat energy transferred to the water during this dissolving process.
You can assume that the specific heat capacity of the solution is the same as that of water, 4.18 J g^{-1} K^{-1}. [1]

(ii) Calculate the amount, in moles, of $MgSO_4$ dissolved. [1]

(iii) Calculate the enthalpy change of solution, ΔH_{soln}, of $MgSO_4(s)$.
You must include a sign with your answer. [1]

In the second experiment, the enthalpy change of solution for the hydrated salt, $MgSO_4.7H_2O(s)$, was calculated and found to be +9.60 kJ mol^{-1}.

(iv) Use the equation below for the hydration of anhydrous magnesium sulfate to construct a suitable, fully labelled energy cycle that will allow you to calculate the enthalpy change for this reaction, ΔH_{hyd} $MgSO_4$.

$$MgSO_4(s) + 7H_2O(l) \rightarrow MgSO_4{\cdot}7H_2O(s)$$

[1]

(v) Calculate the enthalpy change for this reaction, ΔH_{hyd} $MgSO_4$. Include a sign in your answer. [1]

[N15/P2/Q2]

Suggested Solution:

(a) It is the enthalpy change when moles of reactants in an equation react together to give products under standard conditions.

(b) (i) Heat energy, $Q = mc\Delta T$

$$= 45 \times 4.18 \times (34.7 - 23.4)$$
$$= 2125 \text{ J}$$

(ii) $n = \dfrac{3.01}{(24+32.1+64)}$

$= 0.0251$ mol.

(iii) 0.025 mol —— 2125 J

1 mol —— $\dfrac{2125}{0.025}$

$= 85000\ \text{J mol}^{-1} = 85\ \text{kJ mol}^{-1}$

$\therefore\ \ \Delta H_{soln}$ of $MgSO_4 = -85\ \text{kJ mol}^{-1}$

(iv) $MgSO_4(s) + 7H_2O(l) \rightarrow MgSO_4{\cdot}7H_2O(s)$

$-85\ \text{kJ mol}^{-1}$ $\searrow$ $\swarrow$ $+9.60\ \text{kJ mol}^{-}$

$MgSO_4\,(aq)$

(v) $\Delta H_{hyd}\ MgSO_4 = -85 - 9.6$

$= -94.6\ \text{kJ mol}^{-1}$

Question 6

Trihalomethanes are organic molecules in which three of the hydrogen atoms of methane are replaced by halogen atoms, for example CHF_3.

(a) The equation shows a reaction to produce CHF_3.

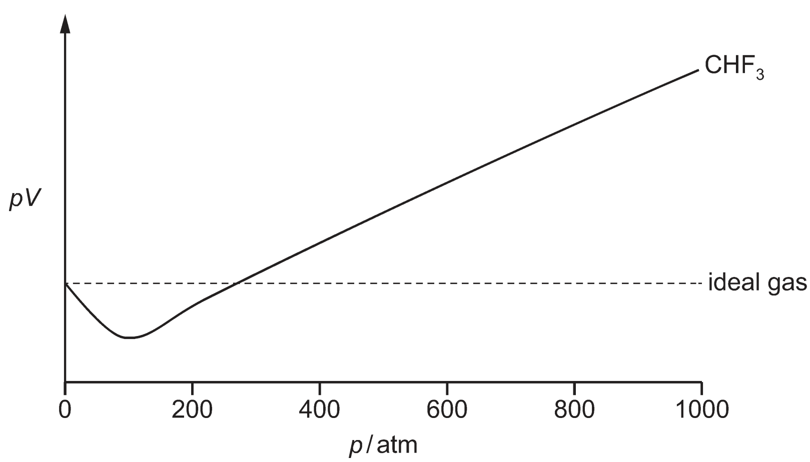
$$CHI_3(s) + 3AgF(s) \rightarrow CHF_3(g) + 3AgI(s)$$

Use the data to calculate the enthalpy change of reaction, ΔH_r, for this formation of CHF_3.

compound	enthalpy change of formation, ΔH_f / kJ mol^{-1}
$CHI_3(s)$	–182.1
$CHF_3(g)$	–692.9
$AgF(s)$	–204.6
AgI (s)	–61.8

[3]

(b) The graph shows the relationship between pV and p at a given temperature for CHF_3 and an ideal gas.

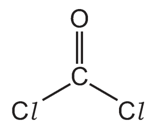


(i) CHF_3 is not an ideal gas.

State **three** basic assumptions that scientists make about the properties of ideal gases. [3]

(ii) Explain why CHF_3 deviates from the properties of an ideal gas at pressures greater than 300 atm. [2]

(c) A different trihalomethane, $CHCl_3$, reacts with O_2 to produce carbonyl dichloride. $HCl(g)$ is also released as a product of this reaction.

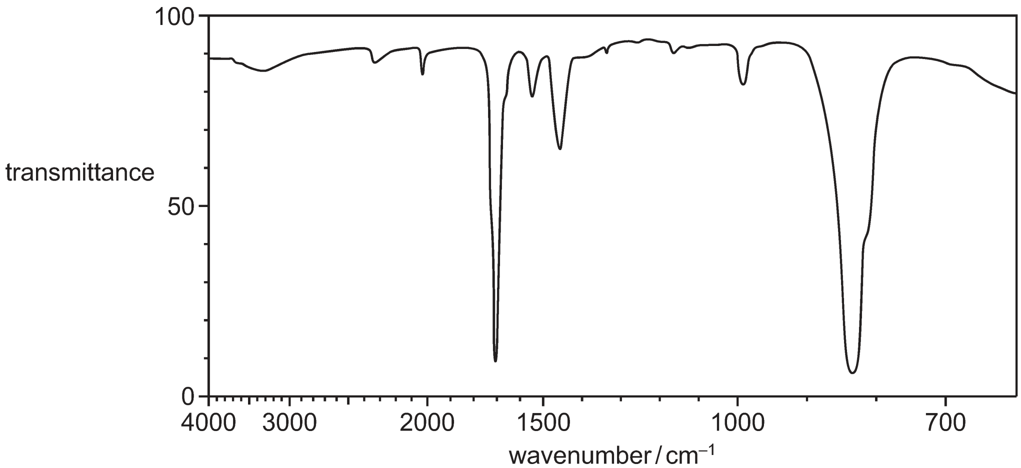


carbonyl dichloride

(i) Write an equation for this reaction of $CHCl_3$ with O_2. [1]

(ii) The conversion of $CHCl_3$ to carbonyl dichloride can be monitored by infra-red spectroscopy.

The infra-red spectrum of carbonyl dichloride is shown.



On the infra-red spectrum of carbonyl dichloride identify with an **X** the absorption that would **not** be present in an infra-red spectrum of $CHCl_3$.

Explain your answer. [2]

(iii) Suggest another difference between the infra-red spectra of $CHCl_3$ and carbonyl dichloride. [1]

[N18/P2/Q3]

Suggested Solution:

(a) ΔH_r = Enthalpy of products − Enthalpy of reactants

$= (-692.9) + 3(-61.8) - [(-182.1) + 3(-204.6)]$

$= -692.9 - 185.4 + 182.1 + 613.8$

$= -82.4\ \text{kJmol}^{-1}$

(b) (i) 1. Gas molecules experience forces only during collisions, which are completely elastic.

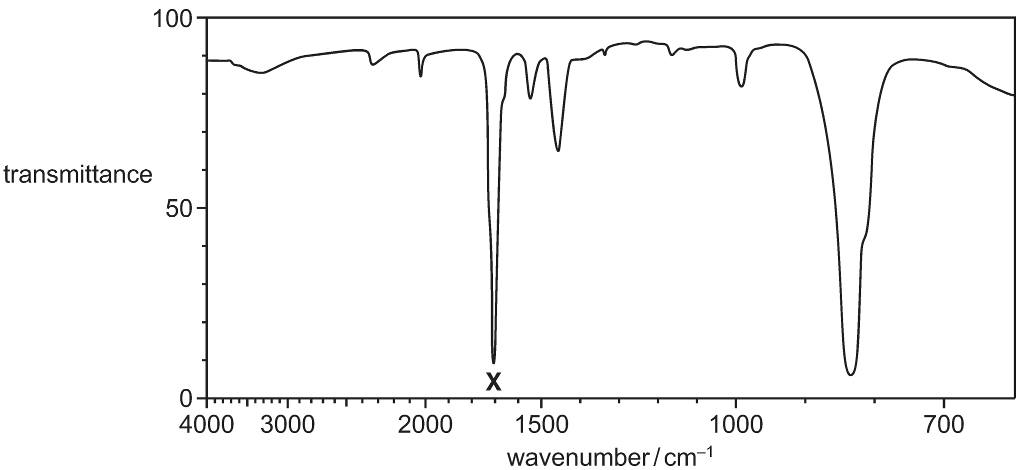
2. There exists negligible forces of attraction between molecules.

3. The volume occupied by the gas molecules themselves is negligible compared to the volume occupied by the gas.

(ii) At high pressure, the gas molecules come closer together to an extent that the repulsive forces between them become significant. This leads to deviation from the properties of an ideal gas.

(c) (i) $CHCl_3 + \frac{1}{2}O_2 \rightarrow COCl_2 + HCl$

(ii)

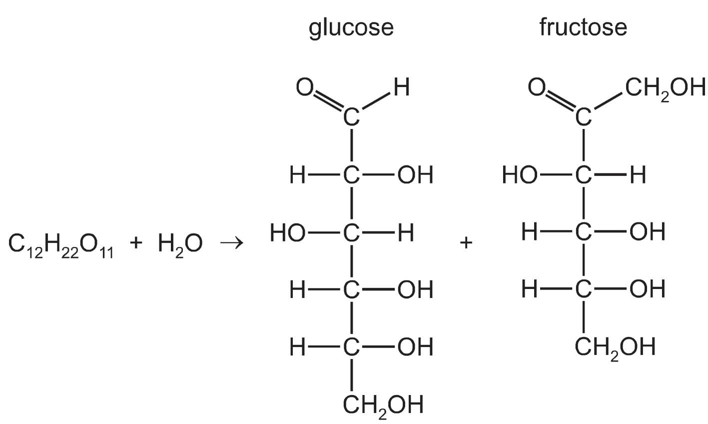


The marked absorption (in wave number) corresponds to a $C=O$ bond which is not present in $CHCl_3$.

(iii) Unlike carbonyl dichloride, $CHCl_3$ contains a C — H bond which shows a peak at 2850 – 2950 cm^{-1}.

Question 7

Sucrose, $C_{12}H_{22}O_{11}$, reacts with water to form glucose and fructose in reaction **A**.



reaction **A**

(a) Suggest a name for this type of reaction. [1]

(b) Explain in detail, why glucose and fructose are a pair of structural isomers. Your answer should refer specifically to these two molecules. [2]

(c) Reaction **A** occurs faster in the presence of an enzyme. This is reaction **B**.

(i) The activation energy for reaction **B** is +29 $kJ\,mol^{-1}$.
Predict a value for the activation energy of reaction **A**. [1]

(ii) The enthalpy change for reaction **A** is –14 $kJ\,mol^{-1}$.
Predict a value for the enthalpy change for reaction **B**. [1]

(iii) Sketch a labelled energy level diagram for reaction **B**. Use relevant values from **(c)(i)** and **(c)(ii)**.

[2]

(d) 1.00 g of sucrose, $C_{12}H_{22}O_{11}$, is completely combusted. The heat energy produced is used to increase the temperature of 250 g of water inside a calorimeter from 25.0 °C to 40.7 °C.

These data can be used to calculate the enthalpy change of combustion of sucrose.

(i) Explain what is meant by the term *enthalpy change of combustion of sucrose*. [2]

(ii) Use the *Data Booklet* to calculate the enthalpy change, in $kJ\,mol^{-1}$, for the combustion of sucrose.
Assume that all of the heat energy produced is transferred to the water.
Show your working. [3]

[J20/P2/Q3]

Suggested Solution:

(a) Hydrolysis.

(a) A disaccharide reacts with acid and is hydrolysed to monosaccharides.

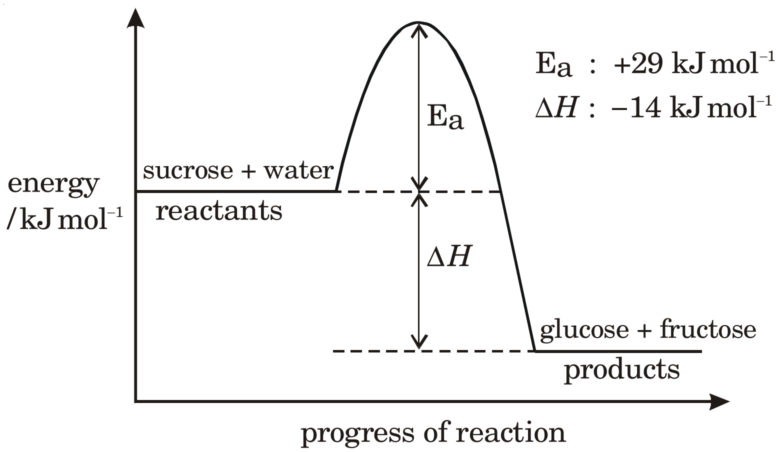
(b) Glucose and Fructose both have a molecular formula of $C_6H_{12}O_6$. The two compounds are isomers because there are same number and type of atoms present in their molecules but with different arrangements. The carbonyl group, for instance, is present at the end of the chain in glucose but in the middle in fructose.

Learning CORNER

(c) (i) Activation energy: +100 kJ mol^{-1}

(ii) Enthalpy change: –14 kJ mol^{-1}

(iii)



(c) (i) The presence of an enzyme lowers the activation energy of a reaction so any value greater than 29 kJ/mol would be acceptable.

(ii) Adding an enzyme have no effect on the enthalpy of the reaction.

(d) (i) Enthalpy change of combustion is the energy change when 1 mole of sucrose burns completely in excess of oxygen.

(ii) $\Delta H = -mc\Delta T$

$\Delta H = -[250 \times 4.18 \times (40.7 - 25)] = -16406.5$ J/g

Number of moles in 1 g of sucrose $= \dfrac{1}{342}$

Enthalpy change in kJ/mol $= \dfrac{-16406.6}{\frac{1}{342} \times 1000}$

$= -16406.6 \times \dfrac{342}{1000} = -5611$ kJ/mol

Question 8

PCl_5, PCl_3 and NCl_3 are halides of Group 15 elements.

(a) PCl_5 can be formed from the reaction of phosphorus with chlorine. PCl_5 has a melting point of 161 °C.

(i) Write an equation for the formation of PCl_5 from the reaction of phosphorus and chlorine. [1]

(ii) State the type of structure and bonding shown by liquid PCl_5. [1]

(b) A small amount of PCl_5 is added to excess water. The PCl_5 reacts vigorously to form a colourless solution.

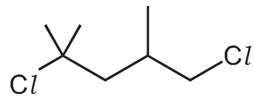
(i) Give **one** other observation you would make when PCl_5 reacts with excess water. [1]

(ii) Write the equation for the reaction of PCl_5 with excess water. [1]

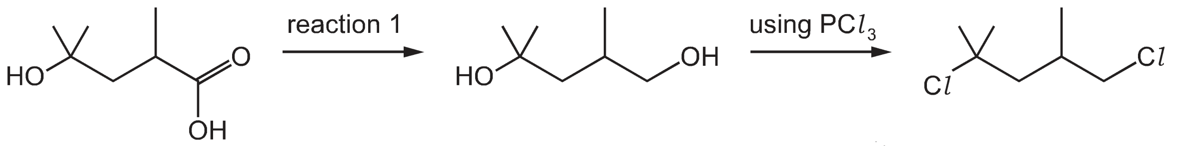
(iii) Estimate the pH of the resulting solution. [1]

(c) PCl_3 is used to convert alcohols to chloroalkanes, such as compound **T**.

T



A possible synthesis of **T** is shown.

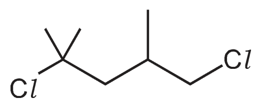


(i) Identify a reagent that could be used in reaction 1. [1]

(ii) **T** exhibits optical isomerism.

Explain what is meant by the term *optical isomer* and circle any atom(s) in **T** that give rise to optical isomerism.

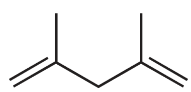
T



[2]

(iii) **T** is a **minor** product in the reaction of compound **S** with excess HCl.

S



Draw the structure of the **major** product of the reaction of **S** with excess HCl. [1]

(d) NCl_3 is a yellow liquid that can be used to bleach flour.

(i) Predict the shape of the NCl_3 molecule and the Cl–N–Cl bond angle. [2]

(ii) NCl_3 reacts with water to form $HOCl$, a weak Brønsted-Lowry acid.

Explain fully what is meant by the term weak *Brønsted-Lowry acid*. [2]

(iii) NCl_3(l) decomposes according to the equation shown.

$$2NCl_3(l) \rightarrow N_2(g) + 3Cl_2(g)$$

A sealed container of volume 250 cm^3 contains an unreactive gas at a pressure of 1.00×10^5 Pa.

0.241 g of NCl_3(l) was injected into the sealed container.

The sealed container was heated to make the NCl_3(l) decompose fully and then cooled to 20 °C.

Calculate the final **total** pressure inside the sealed container at 20 °C after the NCl_3(l) has fully decomposed. [4]

[N20/P2/Q3]

Suggested Solution:

(a) (i) $P_4 + 10Cl_2 \rightarrow 4PCl_5$

(ii) PCl_5 has simple covalent molecules.

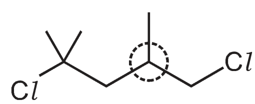
(b) (i) Formation of steamy fumes.

(ii) $PCl_5 + 4H_2O \rightarrow H_3PO_4 + 5HCl$

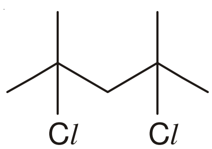
(iii) pH of resulting solution : 2

(c) (i) lithium tetrahydridoaluminate((III)) ($LiAlH_4$)

(ii) Optical isomerism is exhibited by molecules which have a non-super imposable mirror image.



(iii)



(d) (i) Shape: Trigonal pyramidal
Bond angle: 107°

(ii) Brønsted-Lowry acid is a proton donor that partially dissociates in the solution.

(iii) Moles of NCl_3 reacted $= \dfrac{0.241}{120.5} = 0.002$

1 mol of NCl_3 produces 2 mole of gas

$\therefore$ Moles of gas produced $= 0.002 \times 2 = 0.004$

$PV = nRT$

Where, $V = \dfrac{250}{1000000} = 0.00025 \text{ m}^3$

$n = 0.004, \quad R = 8.31 \text{ J mol}^{-1}\text{ K}^{-1}, \quad T = 293 \text{ Kelvin}$

Pressure from gas produced, $P = \dfrac{nRT}{V}$

$$= \frac{0.004 \times 8.31 \times 293}{0.00025} = 3.9 \times 10^4 \text{ Pa}$$

Total pressure $= (3.9 \times 10^4) + (1.00 \times 10^5) = 1.39 \times 10^5$ Pa

(b) (iii) As HCl is produced, the resulting solution will have pH between 0-4.

(c) (i) Carboxylic acids can be reduced to primary alcohols using $LiAlH_4$. For this reaction to occur, a strong reducing agent is required.

(ii) Molecules with chiral carbons (Carbon atoms bonded with 4 different atoms) generally exhibit optical isomerism.

(d) (i) Due to the lone pairs on Nitrogen, the shape of the molecule is trigonal pyramidal.

Question 9

Hydrogen iodide, HI, is a colourless gas at room temperature.

The standard enthalpy change of formation, $\Delta H_f^{\ominus}$, of HI(g) is +26.5 kJ mol^{-1}.
Define the term *standard enthalpy change of formation.* [2]

[N21/P2/Q1 b]

Suggested Solution:

Standard enthalpy change of formation represents energy change when one mole of a compound is formed from its elements in their standard states.

Question 10

When magnesium is heated in air, magnesium oxide, MgO, is the major product. Smaller amounts of magnesium nitride, Mg_3N_2, are also made.

(i) Calculate the oxidation number for magnesium and for the nitrogen species in Mg_3N_2 to complete Table 1.1.

Table 1.1

species	magnesium in Mg_3N_2	nitrogen in Mg_3N_2
oxidation number		

[1]

(ii) Identify the type of reaction which takes place between magnesium and nitrogen.

Explain your answer. [1]

(iii) Define enthalpy change of formation. [2]

(iv) When 3.645 g of Mg(s) burns in excess $N_2(g)$ to form $Mg_3N_2(s)$, 23.05 kJ of energy is released.

Calculate the enthalpy change of formation, ΔH_f, of Mg_3N_2. Show your working. [3]

[J22/P2/Q1 b]

Suggested Solution:

(i)

species	magnesium in Mg_3N_2	nitrogen in Mg_3N_2
oxidation number	**+2**	**–3**

(ii) It is a redox type reaction where Magnesium is oxidised and nitrogen is reduced.

(iii) Enthalpy change of formation is the energy change associated with the formation of one mole of a compound from its elements in their standard states.

(iv) $3Mg + N_2 \rightarrow Mg_3N_2$

Moles of Mg burned $= \dfrac{3.645}{24.3} = 0.15$ mol

Moles of Mg_3N_2 produced $= \dfrac{0.15}{3} = 0.05$ moles

Energy released $= \dfrac{23.05}{0.05} = 461$ kJ

ΔH_f, of $Mg_3N_2 = -461.0$ kJ/mol

(ii) Magnesium loses electrons during the reaction while its oxidation number increases. Nitrogen gains electrons while its oxidation number decreases.

Question 11

In industry, ethanol is made by reacting ethene with steam in the presence of H_3PO_4.

reaction 1 $C_2H_4(g) + H_2O(g) \rightleftharpoons C_2H_5OH(g)$

(a) Use the bond energy values in Table 4.1 to calculate the enthalpy change, ΔH_r, for reaction 1.

Table 4.1

bond	bond energy / kJ mol^{-1}
C–C	350
C=C	610
C≡C	840
C–H	410
C–O	360
C=O	740
O–H	460

[2]

(b) Reaction 1 reaches equilibrium at constant temperature and pressure. Deduce what effect increasing the pressure will have on the amount of ethanol in the new equilibrium mixture. Use Le Chatelier's principle to explain your answer. [2]

(c) Describe the covalent bonds present between the carbon atoms in an ethene molecule by completing Table 4.2.

Table 4.2

	sigma (σ)	pi (π)
type of orbitals involved in bond		
how the orbitals overlap		

[2]

[J23/P2/Q4 a,b,d]

Suggested Solution:

(a) Energy released in bond breaking $= 4(C-H) + (C{=}C) + 2(O-H)$
$= (410 \times 4) + 610 + 2(460)$
$= 3170$ kJ

Energy required for bond making $= (C-C) + 5(C-H) + (C-O) + (O-H)$
$= 350 + 5(410) + 360 + 460$
$= 3220$ kJ

Enthalpy change of reaction, $\Delta H_r = 3170 - 3220$
$= -50$ kJmol^{-1}

(b) Effect of increasing pressure: The equilibrium shifts towards the right and produces more ethanol.

Explanation: As the pressure increases, the equilibrium shifts to the side with less number of gas molecules which in this case is the right side.

(c)

	sigma (σ)	pi (π)
type of orbitals involved in bond	sp^2 - sp^2	(2)p - (2)p
how the orbitals overlap	direct	sideways

Question 12

Phosphoric(V) acid, H_3PO_4, is used in both inorganic and organic reactions.

(a) H_3PO_4 is made in a two-step process from phosphorus.

step 1 Phosphorus reacts with an excess of oxygen to form a white solid.

step 2 The white solid then reacts with water to form H_3PO_4.

(i) Write an equation for each step. [2]

(ii) H_3PO_4 is a weak Brønsted–Lowry acid.

Define weak Brønsted–Lowry acid. [2]

(b) H_3PO_4 is also formed in the process shown in reaction 1.

reaction 1 $4H_3PO_3 \rightarrow 3H_3PO_4 + PH_3$

Table 3.1 shows some relevant thermodynamic data.

Table 3.1

compound	enthalpy change of formation, ΔH_f / kJ mol^{-1}
H_3PO_3	–972
H_3PO_4	–1281
PH_3	+9

(i) Define enthalpy change of formation. [2]

(ii) Use the data in Table 3.1 to calculate the enthalpy change, ΔH_r, of reaction 1. [2]

(iii) Explain why reaction 1 is a disproportionation reaction.

Explain your reasoning with reference to relevant oxidation numbers. [2]

[N23/P2/Q3 a,b]

Suggested Solution:

(a) (i) Step 1: $P_4 + 5O_2 \rightarrow P_4O_{10}$

Step 2: $P_4O_{10} + 6H_2O \rightarrow 4H_3PO_4$

(ii) A weak Brønsted–Lowry acid donates a proton and dissociates only partially in an aqueous solution.

(b) (i) Enthalpy change of formation is the energy change when 1 mol of a compound is formed from its constituent elements in their standard states.

(ii) Enthalpy Change

= Enthalpy change of products − Enthalpy change of reactants.

$\Rightarrow \Delta H_r = \Delta H_f(3H_3PO_4 + PH_3) - \Delta H_f(4H_3PO_3)$

$\Rightarrow \Delta H_r = (3 \times -1281) + 9 - 4(-972)$

$= -3843 + 9 + 3888 = +54 \text{ kJ/mol}$

(iii) In this reaction, P is oxidised and reduced at the same time. Its oxidation state increases from +3 to +5, representing oxidation and it reduces from +3 to –3, representing reduction.

Question 13

Separate samples of Na_2CO_3 and $NaHCO_3$ react with $HCl(aq)$ to produce the same products, as shown in Table 2.1.

Table 2.1

reaction	equation	ΔH / kJ mol^{-1}
1	$Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$	ΔH_1
2	$NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2$	$\Delta H_2 = +27.2$

(a) Complete the reaction pathway diagram in Fig. 2.1 for reaction 2.

Label the diagram to show the enthalpy change, ΔH_2, and the activation energy, E_A. [2]

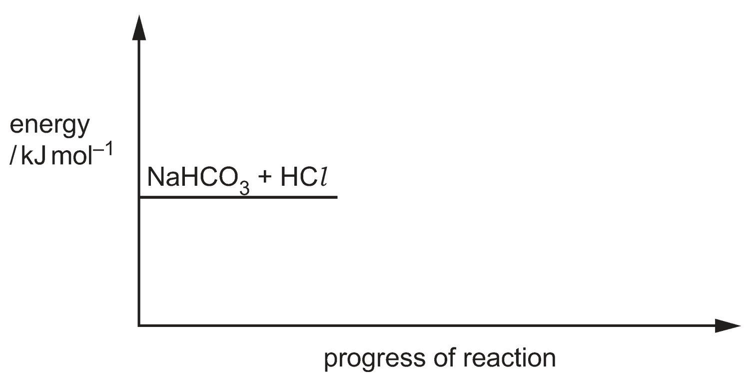


Fig. 2.1

(b) The value for ΔH_1, is determined by experiment using the following method.

- 50.0 cm^3 of 2.00 mol dm^{-3} $HCl(aq)$ is added to a polystyrene cup.
- The initial temperature of the acid is recorded as 19.6 °C.
- 0.0400 mol of Na_2CO_3 is added and the mixture is stirred.
- All the solid Na_2CO_3 disappears and a colourless solution is produced.

The maximum temperature recorded during the reaction is 26.2 °C.

(i) Describe **one** other observation that shows the reaction is complete. [1]

(ii) Calculate the value of ΔH_1 in kJ mol^{-1}.

Assume the specific heat capacity of the reaction mixture is the same as for water and no heat is lost to the surroundings.

Show your working. [3]

(iii) Thermal decomposition occurs when $NaHCO_3$ is heated.

reaction 3 $2NaHCO_3 \rightarrow Na_2CO_3 + H_2O + CO_2$

Calculate the enthalpy change for reaction 3, ΔH_r using the data in Table 2.1 and the value of ΔH_1 calculated in **(b)(ii)**.

(If you were unable to calculate a value for ΔH_1 in **(b)(ii)**, assume the enthalpy change is –38.4 kJ mol^{-1}. This is **not** the correct value.) [2]

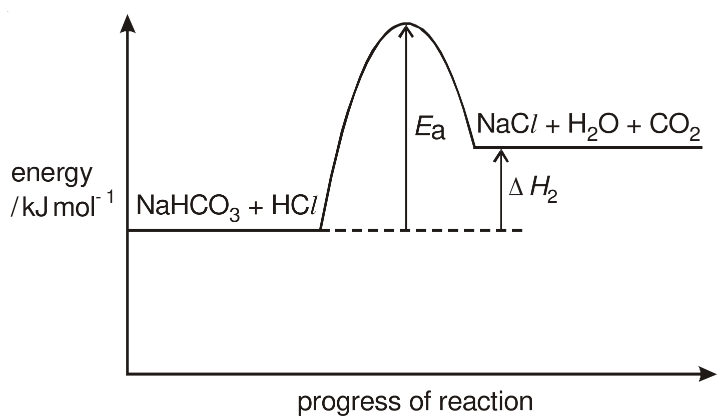
(c) **Z** is a salt that contains a Period 4 element from Group 2. When **Z** is heated brown gas forms.

Identify the formula of **Z** and use it to write an equation for the reaction. [2]

[J24/P2/Q2]

Suggested Solution:

(a)



(b) **(i)** Effervescence stops as no more gas is produced.

(ii) $Q = mc\Delta T$

$Q = 50 \times 4.18 \times (26.2 - 19.6) = 1379.4$ J

Moles of $Na_2CO_3 = 0.0400$

Energy change when 1 mol of Na_2CO_3 reacts $= \dfrac{1379.4}{0.0400} = 34,485$ J

Enthalpy change, $\Delta H_1 = -34.5 \text{ kJ mol}^{-1}$

(iii) Enthalpy change for reaction 3 $= 2(\Delta H_2) - \Delta H_1$

$= 2(27.2) - (-34.5) = 88.9$ kJ/mol

(c) Calcium nitrate : $Ca(NO_3)_2 \rightarrow CaO + 2NO_2 + \frac{1}{2}O_2$