New Specification

Rewarding Learning

Centre Number

71	

Candidate Number

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
January 2009

Chemistry

Assessment Unit AS 1

assessing

Basic Concepts in Physical and Inorganic Chemistry

[AC111]

AC111

FRIDAY 16 JANUARY, MORNING

TIME

1 hour 30 minutes.

INSTRUCTIONS TO CANDIDATES

Write your Centre Number and Candidate Number in the spaces provided at the top of this page.
Answer **all sixteen** questions.
Answer **all ten** questions in **Section A**. Record your answers by marking the appropriate letter on the answer sheet provided. Use only the spaces numbered 1 to 10. Keep in sequence when answering.
Answer **all six** questions in **Section B**. Write your answers in the spaces provided in this question paper.

INFORMATION FOR CANDIDATES

The total mark for this paper is 100.
Quality of written communication will be assessed in question **16(a)(iii)**.
In Section A all questions carry equal marks, i.e. **two** marks for each question.
In Section B the figures in brackets printed down the right-hand side of pages indicate the marks awarded to each question or part question.
A Periodic Table of Elements (including some data) is provided.

For Examiner's use only	
Question Number	Marks
Section A	
1–10	
Section B	
11	
12	
13	
14	
15	
16	
Total Marks	

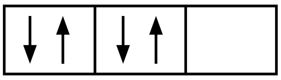
Section A

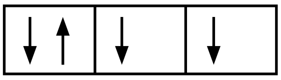
For each of the following questions only **one** of the lettered responses (A – D) is correct.

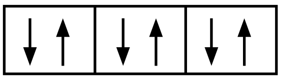
Select the correct response in each case and mark its code letter by connecting the dots as illustrated on the answer sheet

1 Which one of the following represents the p electrons in an oxide, O^{2-} ion?

A 

B 

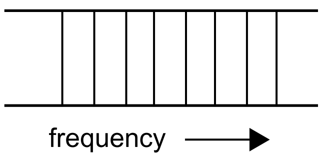
C 

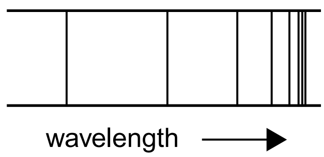
D 

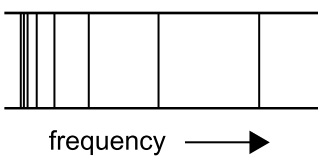
2 Which one of the following represents the electronic configuration for a chromium atom in its ground state?

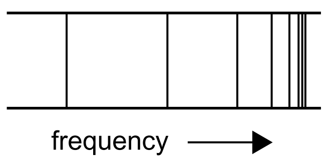
A $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^3$
B $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6$
C $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^1$
D $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^4\ 4s^2$

3 Which one of the following represents the line emission spectrum of atomic hydrogen?

A


B


C


D


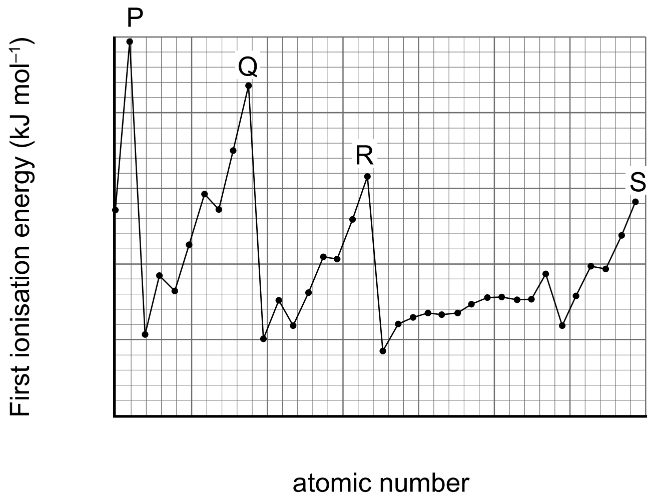
4 The intermolecular forces of attraction in solid iodine are

A covalent bonds.
B hydrogen bonds.
C permanent dipole attractions.
D van der Waals forces.

5 Which one of the following molecules contains the most polar bond?

A CH_4
B NH_3
C H_2O
D HF

6 The graph below represents the variation in the first ionisation energy with atomic number.



The elements indicated by the letters P, Q, R and S are

A alkali metals.
B halogens.
C noble gases.
D transition metals.

7 The melting point of the elements going across the Periodic Table from sodium to argon

A increases steadily.
B decreases steadily.
C increases to silicon and then decreases.
D decreases to silicon and then increases.

8 How many moles of hydrogen ions are present in $40\,cm^3$ of 0.2 M sulphuric acid?

A 8×10^{-3}
B 1.6×10^{-2}
C 0.2
D 0.4

9 In which one of the following do both molecules obey the octet rule?

A $BeCl_2$ and NH_3
B BF_3 and CH_4
C CH_4 and NH_3
D BF_3 and $BeCl_2$

10 Phosphoric acid is manufactured by the reaction of sulphuric acid with calcium phosphate according to the equation:

$$3H_2SO_4 \quad + \quad Ca_3(PO_4)_2 \quad \rightarrow \quad 2H_3PO_4 \quad + \quad 3CaSO_4$$

What mass of phosphoric acid would be obtained from reacting 60 kg of sulphuric acid with 60 kg of calcium phosphate?

A 19 kg
B 38 kg
C 40 kg
D 60 kg

Section B

Answer **all six** questions in this section

Examiner Only

Marks	Remark

11 Complete the table stating the shape of each of the molecules.

Molecule	**Shape**
Ammonia	
Carbon dioxide	
Methane	

[3]

12 **(a)** Metal ions can be identified by the characteristic flame colour observed when their solutions are sprayed into a blue Bunsen flame.

Complete the table by stating the flame colour for each of the ions listed.

Metal ion	Flame colour
Ba^{2+}	
Ca^{2+}	
Cu^{2+}	

[3]

(b) Explain, with the help of an energy level diagram, how flame colours arise.

____

____

____ [3]

(c) Flame colours are a consequence of the emission spectrum of an element.

Explain how the emission spectrum can be used to calculate the first ionisation energy of an element.

____

____

____

____ [2]

Examiner Only	
Marks	Remark

13 Washing soda is hydrated sodium carbonate and can be represented by the formula $Na_2CO_3.xH_2O$. The value of x can be found by titrating a solution of washing soda against standard hydrochloric acid solution.

(a) (i) What is meant by the term **standard** solution?

____ [1]

(ii) Write the equation for the reaction between sodium carbonate, Na_2CO_3, and excess hydrochloric acid.

____ [2]

(b) In one experiment a 2.80 g sample of washing soda was made up to 250 cm^3 of solution in a volumetric flask. 25 cm^3 of this solution required 22.4 cm^3 of 0.1 M hydrochloric acid for neutralisation. Find the value of x using the following headings.

Moles of hydrochloric acid used

____

Moles of sodium carbonate in 25 cm^3

____

Moles of sodium carbonate in the sample

____

Mass of sodium carbonate in the sample

____

Mass of water in the sample

____

Moles of water in the sample

____

Value of x

____ [5]

(c) Suggest a suitable indicator for the titration, stating the colour change expected.

Indicator: ______________________

Colour change: from ________________ to ________________ [3]

Examiner Only	
Marks	Remark

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Examiner Only	
Marks	Remark

14 (a) Magnesium is extracted from dolomite, $MgCO_3.CaCO_3$. Dolomite is heated to form the metal oxides and carbon dioxide. After purification the magnesium oxide is heated with coke (carbon) in a stream of chlorine to form magnesium chloride and carbon monoxide. Magnesium is formed by the electrolysis of molten magnesium chloride.

(i) Write the equation for the effect of heat on dolomite.

____ [2]

(ii) Write the equation for the formation of magnesium chloride from magnesium oxide.

____ [2]

(iii) Magnesium chloride is ionic. Explain why it must be molten for the electrolysis to take place.

____

____ [1]

(iv) Draw dot and cross diagrams to show the formation of magnesium and chlorine atoms from their ions.

[4]

(b) Magnesium is a typical metal.

(i) Draw a labelled diagram to show the bonding in magnesium metal.

[2]

(ii) Explain why magnesium conducts electricity.

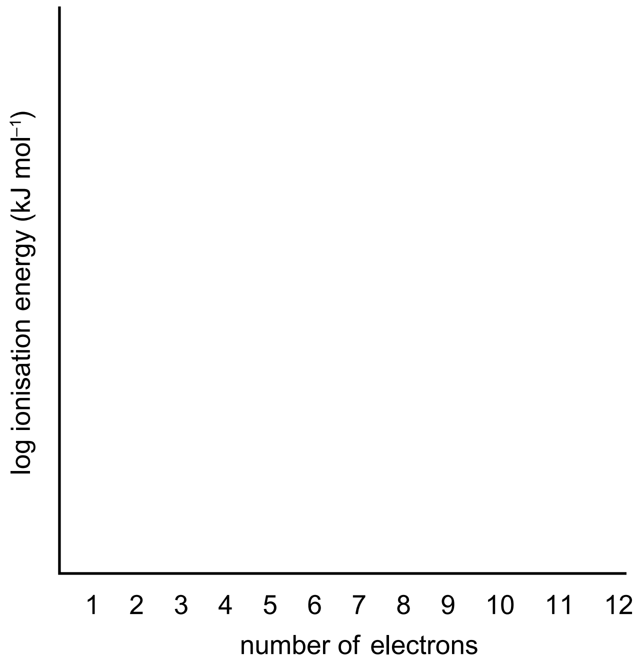
____

____ [2]

(c) **(i)** Write the equation, including state symbols, for the first ionisation energy of magnesium.

____ [2]

(ii) Using the axes below, sketch a graph to show the twelve successive ionisation energies of magnesium.



[2]

Examiner Only	
Marks	Remark

(iii) State **two** reasons why the first ionisation energy of calcium is less than that of magnesium.

____

____

____ [2]

Examiner Only	
Marks	Remark

15 Xenon was first isolated by Ramsey and Travers in 1898.

Examiner Only	
Marks	Remark

(a) Xenon makes up 1 part in 20 000 000 by volume of air.

Calculate the number of atoms of xenon in 1 dm^3 of air at room temperature and pressure using the following headings.

Molar gas volume = 24 dm^3 at room temperature and pressure.

Number of moles in 1 dm^3 of air.

____

Number of particles (atoms and molecules) in 1 dm^3 of air.

____

Number of atoms of xenon in 1 dm^3 of air.

____ [3]

(b) Xenon has a number of naturally occurring isotopes. The table lists the principal isotopes of xenon.

Relative isotopic mass	% abundance
129	27
131	23
132	28
134	12
136	10

(i) Explain what is meant by the term **isotope**.

____

____ [2]

(ii) Use the information in the table to calculate the relative atomic mass of xenon.

____

____

____ [2]

(c) The first compounds of xenon were isolated by Bartlett in 1962.

The reaction between the hydrogenxenate and hydroxide ions can be represented as follows:

$$2OH^- + 2HXeO_4^- \rightarrow Xe + 2H_2O + XeO_6^{4-} + O_2$$

(i) Deduce the oxidation number of xenon in each of the following.

$HXeO_4^-$ __________

Xe __________

XeO_6^{4-} __________ [3]

(ii) Explain why this is considered to be a disproportionation reaction.

____

____

____ [2]

Examiner Only	
Marks	Remark

16 The halogens are found in Group VII of the Periodic Table.

(a) (i) Complete the table.

	Fluorine	**Chlorine**	**Bromine**	**Iodine**
Atomic number	9	17	35	53
Appearance at 20 °C	Yellow gas	Green-yellow gas		
Boiling point (°C)	–188	–35	59	183
Electro-negativity	4.1	2.9	2.8	2.2

[2]

(ii) Explain the change in boiling point of the halogens.

____

____

____ [2]

(iii) State what is meant by the term **electronegativity** and explain the trend for the halogens.

____

____

____

____

____

____ [3]

Quality of written communication [2]

Examiner Only

Marks	Remark

Examiner Only	
Marks	Remark

(b) Chlorine reacts with cold dilute sodium hydroxide to form sodium chlorate(I) and with hot concentrated sodium hydroxide to form sodium chlorate(V).

(i) Write the equation for the reaction of chlorine with cold dilute sodium hydroxide.

____ [2]

(ii) What is the formula of sodium chlorate(V)?

____ [1]

(c) Iodine is more soluble in hexane than in water.

(i) Explain why iodine is more soluble in hexane.

____

____ [2]

(ii) What colour is a solution of iodine in hexane?

____ [1]

(d) The hydrogen halides can be formed by the reaction of concentrated sulphuric acid with the corresponding solid sodium halide.

(i) Write the equation for the reaction of concentrated sulphuric acid with solid sodium chloride at room temperature.

____ [2]

(ii) Name **two** products, other than hydrogen iodide, which are formed when sodium iodide reacts with concentrated sulphuric acid.

____ [2]

(iii) State and explain the trend in thermal stability of the hydrogen halides.

____

____

____ [2]

(e) The presence of halide ions can be detected using silver ions and aqueous ammonia.

Examiner Only

Marks	Remark

(i) Write the ionic equation for the reaction between silver ions and chloride ions.

____ [1]

(ii) Complete the table below.

Halide ion	**Colour of silver salt**	**Effect of adding aqueous ammonia**	
		dilute	**concentrated**
Chloride			
Bromide			
Iodide			

[3]

(f) Explain why the public water supply may be fluoridated and why some people are opposed to this.

____

____

____ [2]

THIS IS THE END OF THE QUESTION PAPER

Published Mark Scheme for GCE AS Chemistry

January 2009

Issued: April 2009

NORTHERN IRELAND GENERAL CERTIFICATE OF SECONDARY EDUCATION (GCSE) AND NORTHERN IRELAND GENERAL CERTIFICATE OF EDUCATION (GCE)

MARK SCHEMES (2009)

Foreword

Introduction

Mark Schemes are published to assist teachers and students in their preparation for examinations. Through the mark schemes teachers and students will be able to see what examiners are looking for in response to questions and exactly where the marks have been awarded. The publishing of the mark schemes may help to show that examiners are not concerned about finding out what a student does not know but rather with rewarding students for what they do know.

The Purpose of Mark Schemes

Examination papers are set and revised by teams of examiners and revisers appointed by the Council. The teams of examiners and revisers include experienced teachers who are familiar with the level and standards expected of 16- and 18-year-old students in schools and colleges. The job of the examiners is to set the questions and the mark schemes; and the job of the revisers is to review the questions and mark schemes commenting on a large range of issues about which they must be satisfied before the question papers and mark schemes are finalised.

The questions and the mark schemes are developed in association with each other so that the issues of differentiation and positive achievement can be addressed right from the start. Mark schemes therefore are regarded as a part of an integral process which begins with the setting of questions and ends with the marking of the examination.

The main purpose of the mark scheme is to provide a uniform basis for the marking process so that all the markers are following exactly the same instructions and making the same judgements in so far as this is possible. Before marking begins a standardising meeting is held where all the markers are briefed using the mark scheme and samples of the students' work in the form of scripts. Consideration is also given at this stage to any comments on the operational papers received from teachers and their organisations. During this meeting, and up to and including the end of the marking, there is provision for amendments to be made to the mark scheme. What is published represents this final form of the mark scheme.

It is important to recognise that in some cases there may well be other correct responses which are equally acceptable to those published: the mark scheme can only cover those responses which emerged in the examination. There may also be instances where certain judgements may have to be left to the experience of the examiner, for example, where there is no absolute correct response – all teachers will be familiar with making such judgements.

The Council hopes that the mark schemes will be viewed and used in a constructive way as a further support to the teaching and learning processes.

CONTENTS

New Specification

Rewarding Learning

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
January 2009

Chemistry

Assessment Unit AS 1

assessing

Basic Concepts in Physical and Inorganic Chemistry

[AC111]

FRIDAY 16 JANUARY, MORNING

MARK SCHEME

		AVAILABLE MARKS
Section A		
1 D		
2 C		
3 D		
4 D		
5 D		
6 C		
7 C		
8 B		
9 C		
10 B		
[2] for each correct answer	[20]	20
	Section A	**20**

AVAILABLE MARKS

Section B

11

Molecule	Shape
Ammonia	**Pyramidal**
Carbon dioxide	**Linear/Straight**
Methane	**Tetrahedral**

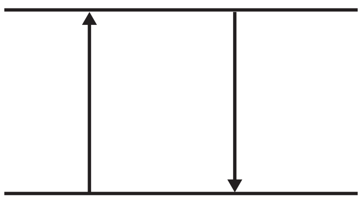
(1 mark each) [3] 3

12 **(a)**

Metal ion	Flame colour
Barium, Ba^{2+}	**Green**
Calcium, Ca^{2+}	**Red/Brick red**
Copper, Cu^{2+}	**Blue-green**

(1 mark each) [3]

(b)



[1]

Electron(s) excited to higher level, returns to lower level [1] releasing energy as light. [1] [3]

(c) Frequency at convergence limit [1],
Multiply by Planck's constant/use E = hf [1] [2] 8

	AVAILABLE MARKS

13 **(a)** **(i)** A solution of known concentration [1]

(ii) $Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$
(unbalanced = 1, balanced = 2) [2]

(b) Moles of HCl = 2.24×10^{-3}
Moles of Na_2CO_3 in 25 cm^3 = 1.12×10^{-3}
Moles of Na_2CO_3 in sample = 1.12×10^{-2}
Mass of Na_2CO_3 in sample = 1.19 g
Mass of H_2O in sample = 1.61 g
Moles of H_2O in sample = 8.94×10^{-2}
Value of *x* = 8
(–1 for each mistake) [5]

(c) Methyl orange [1]
Yellow [1] to orange/pink [1] [3] 11

14 **(a)** **(i)** $MgCO_3.CaCO_3 \rightarrow MgO + CaO + 2CO_2$ [2]

(ii) $MgO + C + Cl_2 \rightarrow MgCl_2 + CO$ [2]

(iii) Ions are free to move [1]

(iv) $Mg^{2+} \rightarrow \overset{\bullet\bullet}{Mg}$

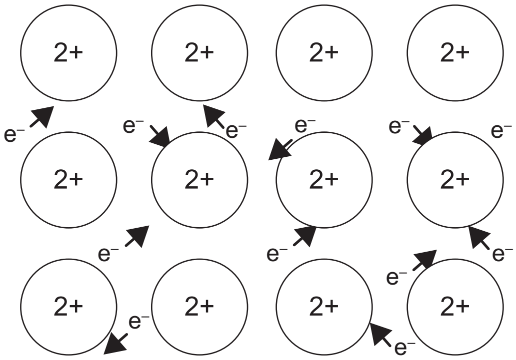
$\overset{xx}{\underset{xx}{{}^{x}_{x}Cl}}{}^{x}_{\bullet} \rightarrow \overset{xx}{\underset{xx}{{}^{x}_{x}Cl}}{}^{x} + e$

or

$2\overset{xx}{\underset{xx}{{}^{x}_{x}Cl}}{}^{x\ominus}_{\bullet} + Mg^{2+} \rightarrow 2\overset{xx}{\underset{xx}{{}^{x}_{x}Cl}}{}^{x} + \overset{\bullet\bullet}{Mg}$ [4]

AVAILABLE MARKS

(b) (i)

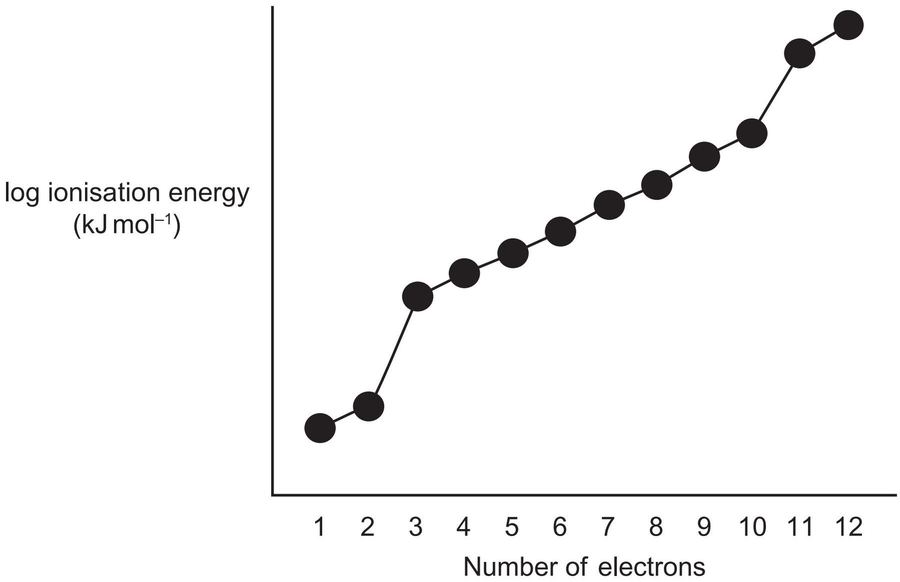


Regular array of positive ions [1] with free electrons around [1]
(–1 if no labels) [2]

(ii) Free electrons [1] can move [1] [2]

(c) (i) $Mg(g) \rightarrow Mg^{+}(g) + e^{-}$
(no state symbols –1) [2]

(ii)



(–1 for each mistake) [2]

(iii) Outer electron further from the nucleus [1]
Increased shielding (from inner electrons) [1] [2] 19

	AVAILABLE MARKS

15 (a) $1/24 = 0.0417$
$0.0417 \times (6.02 \times 10^{23}) = 2.508 \times 10^{22}$
$2.508 \times 10^{22}/2 \times 10^{7} = 1.254 \times 10^{15}$ [3]

(b) (i) Same atomic number/number of protons [1]
different mass number/different number of neutrons [1] [2]

(ii) $$\frac{((129\times 27)+(131\times 23)+(132\times 28)+(134\times 12)+(136\times 10))}{100} = 131.6$$

(–1 for each mistake) [2]

(c) (i) $HXeO_4^-$: +6
Xe: 0
XeO_6^{4-}: +8
(1 mark each) [3]

(ii) Same species (Xe) both oxidised, $+6 \rightarrow +8$, and reduced, $+6 \rightarrow 0$, in the same reaction [1], this is disproportionation (simultaneous oxidation and reduction) [1] [2] 12

AVAILABLE MARKS

16 (a) (i) Bromine: Red-brown liquid [1]
Iodine: Grey-black solid [1] [2]

(ii) Greater mass [1] – greater van der Waals forces [1]
or More electrons [1] – greater van der Waals forces [1] [2]

(iii) Electronegativity: the ability of an element to attract the electrons in a covalent bond towards it. [1] – essential
As the group is descended increased shielding from inner electrons [1]
and distance of bonding electrons from the nucleus [1]
means the nucleus has less pull on the bonding electrons. [1]
Any 2 from 3 [3]
Quality of written communication [2]

(b) (i) $2NaOH + Cl_2 \rightarrow NaOCl + NaCl + H_2O$ [2]

(ii) $NaClO_3$ [1]

(c) (i) Water is polar, hexane non-polar [1]
Hexane and Iodine non-polar [1] [2]

(ii) Purple [1]

(d) (i) $NaCl + H_2SO_4 \rightarrow NaHSO_4 + HCl$ [2]

(ii) Sodium hydrogen sulphate, sulphur dioxide, hydrogen sulphide, iodine, sulphur, water (any two) [2]

(iii) Thermal stability decreases down the group [1]
as bond enthalpy decreases [1] [2]

AVAILABLE MARKS

(e) **(i)** $Ag^+ + Cl^- \rightarrow AgCl$ [1]

(ii)

Halide ion	Colour of silver salt	Effect of adding aqueous ammonia	
		dilute	concentrated
Chloride	White	Soluble	Soluble
Bromide	Cream	Insoluble	Soluble
Iodide	Yellow	Insoluble	Insoluble

[–1] per error [3]

(f) Fluoride is added to water to reduce tooth decay [1]
Lack of choice for individuals/mass medication [1] [2] 27

Section B **80**

Total **100**

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Candidate Number

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
2009

Chemistry

Assessment Unit AS 1

assessing

Module 1: Basic Concepts in Physical and Inorganic Chemistry

[AC111]

WEDNESDAY 3 JUNE, MORNING

AC111

TIME

1 hour 30 minutes.

INSTRUCTIONS TO CANDIDATES

Write your Centre Number and Candidate Number in the spaces provided at the top of this page.
Answer **all seventeen** questions.
Answer **all ten** questions in **Section A**. Record your answers by marking the appropriate letter on the answer sheet provided. Use only the spaces numbered 1 to 10. Keep in sequence when answering.
Answer **all seven** questions in **Section B**. Write your answers in the spaces provided in this question paper.

INFORMATION FOR CANDIDATES

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A Periodic Table of Elements (including some data) is provided.

For Examiner's use only	
Question Number	**Marks**
Section A	
1–10	
Section B	
11	
12	
13	
14	
15	
16	
17	
Total Marks	

Section A

For each of the following questions only **one** of the lettered responses (A – D) is correct.

Select the correct response in each case and mark its code letter by connecting the dots as illustrated on the answer sheet

1 How many electrons are present in a potassium ion, K^+?

A 18
B 19
C 20
D 39

2 Which one of the following represents the first five ionisation energies in kJ mol^{-1} of an s-block element?

	1^{st}	2^{nd}	3^{rd}	4^{th}	5^{th}
A	580	1800	2700	11 600	14 800
B	740	1500	7700	10 500	13 600
C	1000	2300	3400	4600	7000
D	14 800	11 600	2700	1800	580

3 A sample of 4.64 g of hydrated sodium carbonate, $Na_2CO_3.xH_2O$, was dissolved in 1 dm^3 of water. 25.0 cm^3 of this solution required 20.0 cm^3 of 0.05 mol dm^{-3} hydrochloric acid for neutralisation. Which one of the following is the value of *x*?

A 0.5
B 5
C 7
D 13

4 Which one of the following contains a coordinate bond?

A Ammonium, NH_4^+
B Boron trifluoride, BF_3
C Sulphur hexafluoride, SF_6
D Water, H_2O

5 Which one of the following lists the colour of solid iodine and of iodine dissolved in the solvent stated?

	Solid	**Water**	**Hexane**
A	grey/black	purple	yellow/brown
B	dark purple	yellow/brown	purple
C	yellow/brown	grey/black	yellow/brown
D	grey/black	yellow/brown	purple

6 Which one of the following does **not** show the number of each bond present in the named molecules?

	Molecule	**Single bond**	**Double bond**	**Triple bond**
A	Ethene, C_2H_4	2	1	0
B	Nitrogen, N_2	0	0	1
C	Carbon dioxide, CO_2	0	2	0
D	Beryllium chloride, $BeCl_2$	2	0	0

7 In which one of the following molecules does the named element have two lone pairs of electrons?

A Beryllium in $BeCl_2$
B Carbon in CH_4
C Nitrogen in NH_3
D Oxygen in H_2O

8 Using the half-equations below, which one of the following is the balanced ionic equation for the reaction between acidified manganate(VII) ions and ethanedioate ions?

Acidified manganate(VII) ions:

$$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$$

Ethanedioate ions:

$$C_2O_4^{2-} \rightarrow 2CO_2 + 2e^-$$

A $2MnO_4^- + 16H^+ + C_2O_4^{2-} \rightarrow 2Mn^{2+} + 8H_2O + 2CO_2$
B $MnO_4^- + 8H^+ + 5C_2O_4^{2-} \rightarrow Mn^{2+} + 4H_2O + 10CO_2$
C $2MnO_4^- + 16H^+ + 5C_2O_4^{2-} \rightarrow 2Mn^{2+} + 8H_2O + 10CO_2$
D $5MnO_4^- + 40H^+ + 2C_2O_4^{2-} \rightarrow 5Mn^{2+} + 20H_2O + 4CO_2$

9 Which one of the following molecules is non-polar?

A Ammonia, NH_3
B Carbon dioxide, CO_2
C Hydrogen fluoride, HF
D Water, H_2O

10 The extraction and purification of uranium from its ore involves the following reaction between uranium(IV) fluoride and magnesium.

$$2Mg + UF_4 \rightarrow U + 2MgF_2$$

What mass of uranium can be extracted from 500 tonnes of uranium(IV) fluoride and 50 tonnes of magnesium?

A 192 tonnes
B 246 tonnes
C 379 tonnes
D 495 tonnes

Examiner Only	
Marks	Remark

Section B

Answer **all seven** questions in this section

11 **(a)** Complete the table naming the strongest intermolecular force between molecules in each of the following liquids.

Liquid	Intermolecular force
Ammonia, $NH_3(l)$	
Hydrogen chloride, $HCl(l)$	
Methane, $CH_4(l)$	

[3]

(b) Explain why ice has a lower density than water.

____

____

____ [2]

(c) Draw and explain the shape of an ammonia molecule.

____

____

____ [3]

12 Neon has several isotopes.

(a) Complete the table below.

	Number of protons	**Number of electrons**	**Number of neutrons**
Neon-20			
Neon-21			
Neon-22			

[2]

(b) The table below gives the abundance of each isotope of neon.

Calculate the relative atomic mass of neon to two decimal places.

Isotope	**% abundance**
Neon-20	90.92
Neon-21	0.26
Neon-22	8.82

____

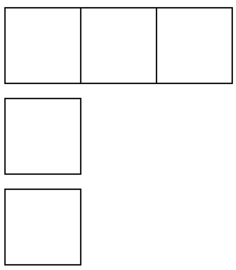
____

____ [2]

(c) Name the isotope used as the standard to compare the relative atomic mass of atoms.

____ [1]

(d) Label the sub-shells below and draw the electronic structure of neon in the ground state.



[2]

Examiner Only	
Marks	Remark

(e) Draw the shape of an s and of a p orbital.

s orbital p orbital [2]

Examiner Only	
Marks	Remark

Examiner Only	
Marks	Remark

13 The percentage of calcium carbonate present in egg shells can be found by back titration using excess hydrochloric acid and standard sodium hydroxide solution.

(a) Write an equation for the reaction between calcium carbonate and hydrochloric acid.

____ [2]

(b) Explain what is meant by a **standard** solution.

____

____ [1]

(c) 1.12 g of an egg shell was reacted with 20.0 cm^3 of 2M hydrochloric acid and the solution formed made up to 250 cm^3 in a volumetric flask. 25.0 cm^3 of this solution completely reacted with 18.6 cm^3 of 0.1 M sodium hydroxide.
Calculate the percentage of calcium carbonate in the egg shell using the headings below.

Moles of hydrochloric acid added to the egg shell

____

Moles of sodium hydroxide used

____

Moles of hydrochloric acid in 250 cm^3

____

Moles of hydrochloric acid which reacted with the egg shell

____

Mass of calcium carbonate in the egg shell

____

Percentage of calcium carbonate in the egg shell

____ [6]

Examiner Only	
Marks	Remark

(d) Name a suitable indicator for the titration of hydrochloric acid with sodium hydroxide solution. Give the colour change observed at the end point.

Indicator: ________________________

Colour change:

from ________________________ to ________________________ [3]

14 The Periodic Table identifies various relationships between elements.

Examiner Only	
Marks	Remark

(a) (i) What property is used to order the elements in the Periodic Table?

____ [1]

(ii) Explain why transition metals are classified as d-block elements.

____ [1]

(b) A number of distinct trends can be seen in the 3rd period from sodium to argon.

(i) Describe the change in melting point across this period.

____

____

____ [2]

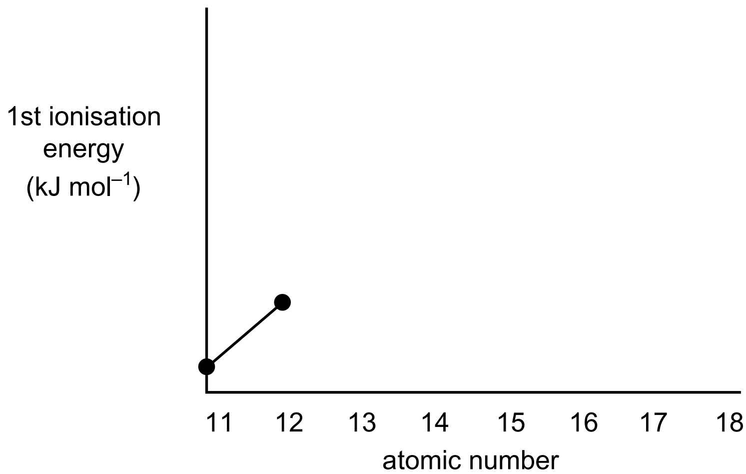
(ii) Describe and explain the change in atomic radius across this period.

____

____

____ [2]

(iii) On the axes below sketch the change in the 1st ionisation energy across the 3rd period.



[3]

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(Questions continue overleaf)

15 (a) Diamond and graphite have giant covalent structures.

(i) Explain what is meant by the term **covalent**.

____ [1]

(ii) Describe the structures of diamond and graphite.

Diamond: ____

____

____ [2]

Graphite: ____

____

____ [2]

(iii) Explain why graphite conducts electricity.

____

____

____ [2]

(iv) Explain why diamond is exceptionally hard.

____

____ [1]

Examiner Only	
Marks	Remark

(b) Carbon dioxide, CO_2, and beryllium chloride, $BeCl_2$, are both covalent compounds.

(i) Draw dot and cross diagrams for carbon dioxide and for beryllium chloride.

carbon dioxide beryllium chloride

[2]

(ii) State the **octet rule** and explain why beryllium chloride does not obey it.

____

____

____ [2]

Examiner Only	
Marks	Remark

16 Rock salt, impure sodium chloride, is found in large underground deposits at Kilroot.

Examiner Only

Marks	Remark

(a) (i) Describe how you would carry out chemical tests used to show that solid rock salt contains sodium chloride.

____

____

____

____

____

____ [5]

Quality of written communication [2]

(ii) Draw dot and cross diagrams to show how sodium chloride is formed from sodium and chlorine atoms.

[4]

(b) Chlorine is manufactured by the electrolysis of concentrated sodium chloride solution.

(i) Explain why sodium chloride solution conducts electricity but solid sodium chloride does not.

____

____ [1]

(ii) Household bleach is manufactured by reacting chlorine with sodium hydroxide solution.

$$Cl_2 + 2NaOH \rightarrow NaCl + NaOCl + H_2O$$

Using oxidation numbers, explain why this reaction is described as disproportionation.

____

____

____ [3]

(iii) Describe what you would observe when chlorine is bubbled through a solution of potassium bromide.

____

____

____ [2]

(iv) Write an ionic equation for the reaction of chlorine with potassium bromide.

____ [1]

(c) Concentrated sulphuric acid reacts with sodium halides to form the corresponding hydrogen halide.

(i) Write an equation for the reaction of concentrated sulphuric acid with sodium chloride.

____ [2]

(ii) Give **two** observations when concentrated sulphuric acid is added to sodium iodide.

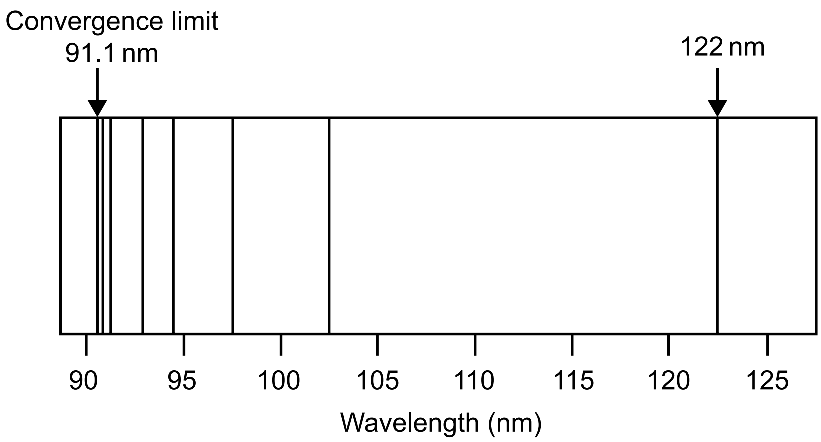
____

____ [2]

Examiner Only	
Marks	Remark

17 The electronic structure of atoms has been interpreted from analysis of emission spectra.

The diagram below shows the emission spectrum of hydrogen in the ultraviolet region.



(a) Draw the electron transition responsible for the line at 122 nm.

n = 4 ____________________________

n = 3 ____________________________

n = 2 ____________________________

n = 1 ____________________________ [2]

(b) Explain what is meant by the convergence limit.

____ [1]

Examiner Only

Marks	Remark

(c) The convergence limit can be used to calculate the ionisation energy for hydrogen.

(i) Write an equation, including state symbols, for the ionisation of atomic hydrogen.

____ [2]

(ii) Use the information below to calculate the frequency of the line at the convergence limit.
(speed of light = 3×10^{8} m s^{-1}, 1 nm = 1×10^{-9} m)

speed of light = frequency $\times$ wavelength

____

____ [1]

(iii) Use this frequency value to calculate the energy required to ionise one mole of hydrogen atoms.

Energy required to ionise one hydrogen atom

____

Energy required to ionise one mole of hydrogen atoms in kJ mol^{-1}

____ kJ mol^{-1} [2]

THIS IS THE END OF THE QUESTION PAPER

Examiner Only	
Marks	Remark

Published Mark Schemes for GCE AS Chemistry

Summer 2009

Issued: October 2009

NORTHERN IRELAND GENERAL CERTIFICATE OF SECONDARY EDUCATION (GCSE) AND NORTHERN IRELAND GENERAL CERTIFICATE OF EDUCATION (GCE)

MARK SCHEMES (2009)

Foreword

Introduction

Mark Schemes are published to assist teachers and students in their preparation for examinations. Through the mark schemes teachers and students will be able to see what examiners are looking for in response to questions and exactly where the marks have been awarded. The publishing of the mark schemes may help to show that examiners are not concerned about finding out what a student does not know but rather with rewarding students for what they do know.

The Purpose of Mark Schemes

Examination papers are set and revised by teams of examiners and revisers appointed by the Council. The teams of examiners and revisers include experienced teachers who are familiar with the level and standards expected of 16 and 18-year-old students in schools and colleges. The job of the examiners is to set the questions and the mark schemes; and the job of the revisers is to review the questions and mark schemes commenting on a large range of issues about which they must be satisfied before the question papers and mark schemes are finalised.

The questions and the mark schemes are developed in association with each other so that the issues of differentiation and positive achievement can be addressed right from the start. Mark schemes therefore are regarded as a part of an integral process which begins with the setting of questions and ends with the marking of the examination.

The main purpose of the mark scheme is to provide a uniform basis for the marking process so that all the markers are following exactly the same instructions and making the same judgements in so far as this is possible. Before marking begins a standardising meeting is held where all the markers are briefed using the mark scheme and samples of the students' work in the form of scripts. Consideration is also given at this stage to any comments on the operational papers received from teachers and their organisations. During this meeting, and up to and including the end of the marking, there is provision for amendments to be made to the mark scheme. What is published represents this final form of the mark scheme.

It is important to recognise that in some cases there may well be other correct responses which are equally acceptable to those published: the mark scheme can only cover those responses which emerged in the examination. There may also be instances where certain judgements may have to be left to the experience of the examiner, for example, where there is no absolute correct response – all teachers will be familiar with making such judgements.

The Council hopes that the mark schemes will be viewed and used in a constructive way as a further support to the teaching and learning processes.

CONTENTS

New Specification

Rewarding Learning

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
2009

Chemistry

Assessment Unit AS 1

assessing

Module 1: Basic Concepts in Physical and Inorganic Chemistry

[AC111]

WEDNESDAY 3 JUNE, MORNING

MARK SCHEME

		AVAILABLE MARKS

Section A

1	A		
2	B		
3	C		
4	A		
5	D		
6	A		
7	D		
8	C		
9	B		
10	B		
	[2] for each correct answer	[20]	20
		Section A	**20**

AVAILABLE MARKS

Section B

11 (a)

Molecule	Attractive force
Ammonia	Hydrogen bonds
Hydrogen chloride	Dipole-Dipole
Methane	Van der Waals

[1] each [3]

(b) More/longer/fixed Hydrogen bonds (between the water molecules) [1]
give ice a more open structure (and so a lower density) [1] [2]

(c)

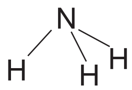


Diagram [1]
Repulsion between electron pairs [1]
Mention of **four** pairs or comment on lone pair [1] [3] 8

12 (a)

	Number of protons	Number of electrons	Number of neutrons
Neon-20	10	10	10
Neon-21	10	10	11
Neon-22	10	10	12

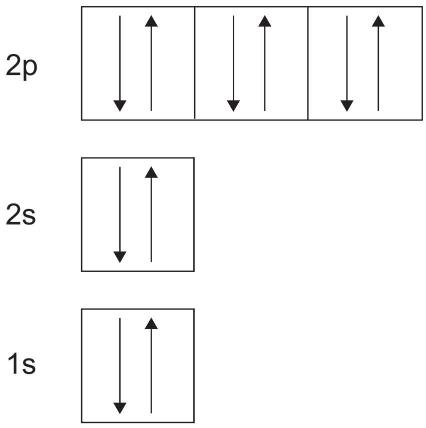
[–1] for each mistake [2]

(b) $$\frac{(20 \times 90.92) + (21 \times 0.26) + (22 \times 8.82)}{100}$$

= 20.18
[–1] for each mistake [2]

(c) Carbon-12 isotope [1]

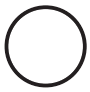
(d)



Subshell labels [1] electronic arrangement [1]
[–1] for each mistake [2]

AVAILABLE MARKS

(e)



s orbital

p orbital

[1] each [2] 9

13 **(a)** $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$

[2], [–1] for each mistake [2]

(b) A solution of known concentration [1]

(c) (2 x 20)/1000 = 0.04 mole

(0.1 x 18.6)/1000 = 0.00186 mole

0.0186

0.04 – 0.0186 = 0.0214

(0.0214/2) x 100 = 1.07(g)

(1.07/1.12) x 100 = 95.5%

[–1] for each mistake [6]

(d) phenolphthalein/methyl orange [1]

from colourless [1] to pink/from red [1] to yellow [1] [3] 12

14 **(a)** **(i)** Atomic number [1]

(ii) Their (outer) electrons are in the d-subshell [1]

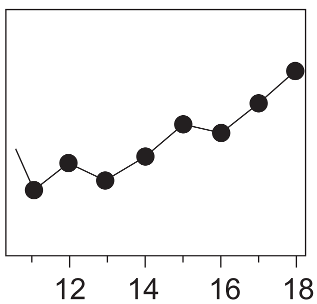
(b) **(i)** Melting point increases to silicon [1]

then decreases (to argon) [1] [2]

(ii) Atomic radius decreases across the period [1]

Shielding remains the same but nuclear charge increases [1] [2]

(iii)



General rise across the period [1]

Fall between Groups 2 and 3 [1]

Fall between Groups 5 and 6 [1] [3] 9

AVAILABLE MARKS

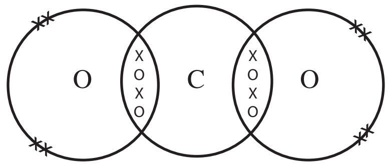
15 **(a)** **(i)** Pair(s) of electrons shared between (two) atoms [1]

(ii) Diamond: Carbon atoms joined to 4 others [1] tetrahedrally [1]
Graphite: hexagonal rings of carbon atoms [1] in layers [1] [4]

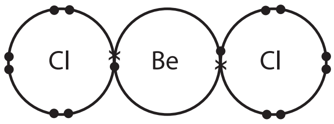
(iii) Free electrons [1] are able to move [1] around the layers [2]

(iv) Strong (covalent) bonds [1] throughout the giant (tetrahedral) structure [1]

(b) **(i)**



[–1] for each mistake [1]



[–1] for each mistake [1]

(ii) Octet rule: eight electrons in the outer shell (when bonded) [1]
Be (has less than 8) in beryllium chloride/has only 4 electrons in its outer shell [1] [2] 12

16 **(a)** **(i)** Sodium: nichrome wire [1]/(conc HCl) blue flame [1]/yellow [1]
Chloride: (make a solution) silver nitrate [1] white precipitate [1]
or dissolve in HNO_3 [1] (solution) [5]

Quality of written communication [2]

(ii) $Na^+ \cdot\ddot{\underset{\cdot\cdot}{Cl}}: \rightarrow Na^+ \,\overset{\times}{\cdot}\ddot{\underset{\cdot\cdot}{Cl}}:^-$ [4]

(b) **(i)** In the solution the ions are free to move, (they cannot move in the solid) [1]

(ii) Chlorine atoms are both oxidised (0 to +1) [1] and reduced (0 to –1) [1] this is disproportionation [1] [3]

(iii) Colourless solution [1]
turns yellow [1] [2]

(iv) $Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$ [1]

				AVAILABLE MARKS
(c)	**(i)**	$H_2SO_4 + NaCl \rightarrow NaHSO_4 + HCl$ (–1 if Na_2SO_4)	[2]	
	(ii)	Steamy fumes/purple vapour/yellow solid/fizzing/heat evolved/ grey-black solid/ rotten egg smell/choking gas (SO_2) any **two**, [1] each	[2]	22
17 (a)		Level 2 to level 1 [1] indication of downwards [1]	[2]	
(b)		An electron leaves the atom/energy levels come together	[1]	
(c)	**(i)**	$H(g) \rightarrow H^+(g) + e^-$ (–1 for each mistake)	[2]	
	(ii)	$3 \times 10^8 = 91.1 \times 10^{-9} \times$ frequency 3.29×10^{15}	[1]	
	(iii)	$E = hf = (6.63 \times 10^{-34}) \times (3.29 \times 10^{15}) = 2.18 \times 10^{-18}$ [1] $(2.18 \times 10^{-18}) \times (6.02 \times 10^{23}) = 1314\,kJ\,mol^{-1}$ [1]	[2]	8
			Section B	**80**
			Total	**100**

New Specification

Rewarding Learning

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
2009

Chemistry

Assessment Unit AS 2

assessing

Module 2: Further Physical and Inorganic Chemistry and Introduction to Organic Chemistry

[AC121]

THURSDAY 11 JUNE, AFTERNOON

MARK SCHEME

Quality of written communication:

2 marks	The candidate expresses ideas clearly and fluently through well-linked sentences and paragraphs. Arguments are generally relevant and well-structured. There are few errors of grammar, punctuation and spelling.
1 mark	The candidate expresses ideas clearly, if not always fluently. Arguments may sometimes stray from the point. There may be some errors of grammar, punctuation and spelling, but not such as to suggest a weakness in these areas.
0 marks	The candidate expresses ideas satisfactorily, but without precision. Arguments may be of doubtful relevance or obscurely presented. Errors in grammar, punctuation and spelling are sufficiently intrusive to disrupt the understanding of the passage.

Section A

		AVAILABLE MARKS
1 A		
2 C		
3 D		
4 D		
5 C		
6 B		
7 B		
8 C		
9 B		
10 B		
[2] for each correct answer	[20]	20
	Section A	**20**

Section B

		AVAILABLE MARKS
11 copper sulphate [1] sulphuric acid [1] sodium chromate [1] potassium sulphate [1]	[4]	4

12 **(a)** reaction is endothermic + mention of high temperature [1]
reaction moves to RHS to absorb heat [2]
increased temperature speeds up reaction [1]
To a maximum of [3] [3]

(b) Comment/statement on formula for % atom economy [2]

$CH_4(g) + 2H_2O(g) \rightarrow CO_2(g) + 4H_2(g)$
16 36 44 8

$$\frac{8}{52} \times 100 = 15.4\% \quad [2]$$

atom economy very low [1]
(no problem with disposal of carbon dioxide [1])
except that it causes global warming [1]
need to find a use for the carbon dioxide [1]
large amount of waste produced [1]
To a maximum of [4] [4]

(c) energy needed to break bonds in reactants
4C—H = 4 × 413 = 1652
4O—H = 4 × 464 = 1856 total = 3508

energy needed to break bonds in products
2C═O = 2 × 805 = 1610
4H—H = 4 × 436 = 1744 total = 3354

+3508 – 3354 = (+)154 kJ [3]

(d) high pressure (is expensive to maintain/thicker pipes needed) [1]
high temperature/endothermic (needs fuel which is expensive) [1] [2]

(e) **(i)** carbon monoxide is poisonous [1]

(ii) use an excess of water/steam [1]

(iii) $1CH_4 \rightarrow 3H_2$
$16\,g \rightarrow 3 \times 24\,dm^3$
16 tonnes $\rightarrow 10^6 \times 3 \times 24\,dm^3$
4 tonnes $\rightarrow 10^6 \times 3 \times 6\,dm^3$
$= 1.8 \times 10^7\,dm^3$ [3] 17

AVAILABLE MARKS

13 **(a)** **(i)** conc. ammonia [1]
white smoke [1] [2]

(ii) $CH_3OH + HCl \rightarrow CH_3Cl + H_2O$ [1]

(iii) the gas would escape from the reaction flask/needs to be condensed/difficult to collect [1]

(b) **(i)** $CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$ [1]

(ii) $1s^2 2s^2 2p^6 3s^2 3p^5$ [1]

(iii) provides the energy [1]
to break the Cl—Cl bond [1] [2]

(iv) $^{\bullet}CH_3 + {}^{\bullet}CH_3 \rightarrow C_2H_6$ [1]

(c) $CH_3NH_2/CH_3\overset{+}{N}H_3\,Cl^-$ [1]
CH_3OH [1] [2] 11

14 **(a)** **(i)** number of molecules [1]

(ii) no molecules hence no energy [1]

(iii) there are always some molecules with (a higher/some) energy [1]

(iv) the number of particles [1]
with (energy greater than) the activation energy [1] [2]

(b) **(i)** peak moves to the RHS [1]
and is lower [1] [2]

(ii) more particles have (energy greater than) the activation energy [1]

(c) the activation energy is less [1]
the number of particles with the required activation energy is greater [1] [2] 10

AVAILABLE MARKS

15 **(a)** the outer shells contain s electrons [1]

(b) the solubility of the sulphates decreases down the group [1]

(c) **(i)** $CaSO_4.2H_2O = 40 + 32 + 64 + 2 \times 18 = 172$
$CaSO_4 = 40 + 32 + 64 = 136$
172 g of gypsum gives 136 g of anhydrous calcium sulphate
34.4 g of gypsum gives $136/172 \times 34.4$ g of anhydrous
calcium sulphate = 27.2 g
percentage yield = $26.0/27.2 \times 100 = 95.588 = 95.6\%$ [3]

(ii) $CaSO_4 \rightarrow CaO + SO_3$ [1]

(iii) the size of the cation increases down the group [1]
the polarising power of the cation is decreased [1]
charge density less [1]
anion is "deformed" [1]
Any three from four [3]

(d) **(i)** the enthalpy change in a reaction is independent of the route taken [2]

(ii)

$$\begin{tikzcd}
MgSO_4(s) + 7H_2O(l) \arrow[rr, "\Delta H_1"] \arrow[dr, "\Delta H_2"'] & & MgSO_4.7H_2O(s) \arrow[dl, "\Delta H_3"] \\
& MgSO_4(aq) &
\end{tikzcd}$$

$$\Delta H_1 + \Delta H_3 = \Delta H_2$$

$\Delta H_2 = 10 \text{ x } 100 \text{ x } 4.2 \text{ x } 9 = 37.8 \text{ x } 10^3 \text{ J} = -37.8 \text{ kJ}$ (exothermic)
$\Delta H_3 = 10 \text{ x } 100 \text{ x } 4.2 \text{ x } 3 = 12.6 \text{ x } 10^3 \text{ J} = +12.6. \text{ kJ}$ (endothermic)

$\Delta H_1 + 12.6 = -37.8$ $\quad \Delta H_1 = -50.4$ kJ [4]

(iii) Any **three** from **four**
- insulated container/plastic cup [1]
- add solid to water [1] with stirring [1]
- measure temperature before and after with thermometer [1]

major error is heat loss [1]
increase insulation/use lid/stop draughts [1]
(Maximum [5]) [5]

Quality of written communication [2] 22

16 **(a)** **(i)** $C_3H_6 + NH_3 + 1.5O_2 \rightarrow CH_2{=}CHCN + 3H_2O$ [2]

AVAILABLE MARKS

(ii) catalyst [1]

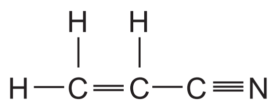
(iii) the IR spectra are unique for a compound they will differ [1]
because of the different position of the C=C absorption [1]/
fingerprint regions of each don't match
or
superimpose the spectra of each compound
they will be different [2]

(b) $CH_2{=}CHCl + NaCN \rightarrow CH_2{=}CHCN + NaCl$ [2]

(c) **(i)** C_3H_3N [1]

(ii) C_3H_3N [1]

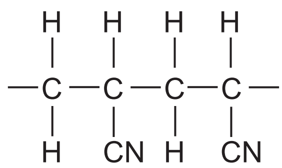
(iii)



[2]

(iv) no [1]
propenonitrile contains CH_2 group/explain by drawing structures [1] [2]

(d) **(i)**



[2]

(ii) addition polymerisation [1] 16

Section B	**80**
Total	**100**

New Specification

Rewarding Learning

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
2009

Chemistry

Assessment Unit AS 3

assessing

Module 3: Practical Examination 1

[AC131]

MONDAY 11 MAY, AFTERNOON

MARK SCHEME

AVAILABLE MARKS

Section A

1 Titration exercise

(a) **(i)** Rinse out a pipette with one of the solutions and (using a pipette filler) transfer/pipette a known volume of the solution into a conical flask [1]
(ii) Add 2 or 3 drops of phenolphthalein [1]
(iii) Rinse out the burette with the other solution and fill the burette [1]
(iv) Add the solution from the burette until the end point is reached [1]
(v) Repeat (for accuracy) [1]

To a maximum of [4]

(b) Table [1]
Significant figures [2]
Calculation of average titre [2]
Titration consistency [3]
Agreement with supervisor's titre [4] [12]

NOTES:

Table:
Table should include initial burette reading, final burette reading, and volume delivered for rough and accurate titrations. [1]

Significant figures:
All burette readings should be to at least one decimal place – each mistake is penalised by [1].
(However, initial burette readings of 0 are penalised once only)
If used, the second decimal position should be 0 or 5 only – other values are penalised by [1]. [2]

Average titre:
The average titre should be calculated and units should be included.
Accurate titrations only should be used.
The use of the rough value is [–1].
The average value can be two or three decimal places, e.g. 25.375/25.38
An incorrect calculation is 0. [2]

Titration consistency:
This is the difference between the first and second accurate readings.

Difference less than or equal to	Mark
0.1	3
0.2	2
0.3	1
0.4	0

[3]

AVAILABLE MARKS

Titration agreement with supervisor:
This is the difference between the candidate's calculated average titre and the supervisor's value.

Difference less than or equal to	Mark
±0.1	[4]
±0.2	[3]
±0.3	[2]
±0.4	[1]
±0.5	[0]

[4]

(c) Colourless to pink/red (or vice-versa depending on titration) [1]

(d) $CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l)$
[1] for equation, [1] for state symbols [2]

(e) **(i)** Correct calculation using volume in dm^3 [2]

$$\text{Number of moles} = \text{concentration } (\text{mol dm}^{-3}) \times \frac{\text{volume } (\text{cm}^3)}{1000}$$

(ii) Uses 1 : 1 ratio from equation
If equation incorrectly balanced in **(d)** the candidate's ratio should be used (i.e. carry error through (c.e.t.)) [1]

(iii) Correct calculation using volume in dm^3
Divide by volume (in dm^3) of ethanoic acid used in titration [2]

(iv) Correct calculation
Multiply by RMM of ethanoic acid [1] [6] 25

2 Observation/deduction

Safety goggles must be worn at all times and care should be exercised during this practical examination.

(a) You are provided with a mixture of two salts, labelled A, which have a common cation. Carry out the following experiments on the mixture. Record your observations and deductions in the spaces below and identify the two salts.

Experiment	Observations	Deductions	
1 Describe the appearance of A.	*White solid* [1]	*Does not contain a Transition metal ion. Group I or II/Ammonium compound/s-block ion present* [1]	[2]
2 Dip a wire loop in concentrated hydrochloric acid; touch sample A with the wire, then hold it in a blue Bunsen flame.	*Yellow/orange/golden* [1]	*Na^+/Sodium ion/compound (present)* [1]	[2]
In a fume cupboard: **3** Add about 1 cm^3 of concentrated sulphuric acid to a half spatula-measure of A in a test tube. Test the gas given off using a glass rod which has been dipped into concentrated ammonia solution.	*Bubbling/fizz/gas given off/ effervescence/frothing* [1] *Misty/Steamy fumes* [1] *White smoke/cloudy/ solid/fumes* [1]	*Hydrogen chloride/HCl (g) or chloride/Cl^-* [1]	[4]
4 Make up a solution of A by dissolving a half spatula-measure of A in a test tube half-full of dilute nitric acid. Put 1 cm^3 of the solution into each of two separate test tubes.	*No effervescence* [1] *Colourless solution* [1]	*Not carbonate or hydrogencarbonate or sulphite* [1]	
(a) (i) Add a few drops of silver nitrate solution into the first test tube.	*White solid/ precipitate* [1]		
(ii) Add about 1 cm^3 of concentrated ammonia into the first test tube.	*White precipitate dissolves/ disappears* [1]	*Chloride* [1]	
(b) Add a few drops of barium chloride solution into the second test tube.	*White solid/ precipitate* [1]	*Sulphate* [1]	[8]

Name the two salts present in A:

Sodium chloride [1]

Sodium sulphate [1]

AVAILABLE MARKS

(b) You are provided with an aqueous solution containing an organic substance X. Carry out the following experiments. Record your observations and deductions in the spaces below.

Experiment	Observations	Deductions	
1 Describe the solution and add a few drops on to Universal Indicator paper.	*Colourless (solution)* [1] *Shades of green* [1]	*Not a carboxylic acid/* *Neutral* [1]	[3]
In a fume cupboard: **2** Shake a small volume of the solution with bromine water.	*Yellow/orange colour* [1] *remains* [1]	*Saturated or no C═C* [1] *Not an alkene/not unsaturated*	[3]
3 Heat about 2 cm^3 of the solution with 2 cm^3 of acidified potassium dichromate solution.	*Orange* [1] *to green* [1] *Change in smell* [1]	*Primary or secondary alcohol/not a tertiary alcohol/aldehyde* [1] *can be oxidised/aldehyde or ketone formed* *is a reducing agent* [1]	[5]

AVAILABLE MARKS

Based on the above tests, suggest

A functional group(s) which may be present in X:

OH [1]

A functional group(s) which the tests used above show is absent in X:

C═C or —COOH [1]

To a maximum of [29] 29

Section A **54**

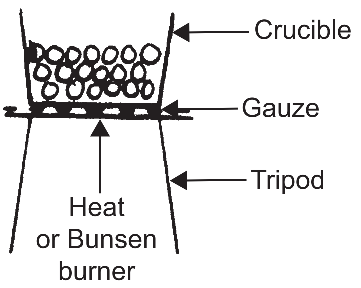
AVAILABLE MARKS

Section B

3 Planning

(a) (i) contains water of crystallisation [1]

(ii)



Heat/Bunsen burner – correctly labelled [1]
Tripod and gauze – correctly labelled [1]
Crucible – correctly labelled [1] [4]
(–1 for each omission)

(b) (i) Mass of container, e.g. crucible [1]
Mass of crucible + (hydrated) sodium carbonate [1]

(ii) Heat and weigh [1]
Repeat [1]
Until there is no further decrease in mass/to constant mass [1]

(iii) Allow apparatus (crucible) to cool/use gloves/use tongs [1] [6]

(c) (i) 11.44 – 4.24 = 7.20 g [1]

(ii) 7.2 ÷ 18 = 0.4 moles [1]

(iii) 4.24 ÷ 106 = 0.04 moles [1]

(iv)

Na_2CO_3	H_2O	
0.04	0.4	
1	10	[1]

$x = 10$ [1] [5]

(d) (i) 2.65 ÷ 106 = 0.025 mole [1]

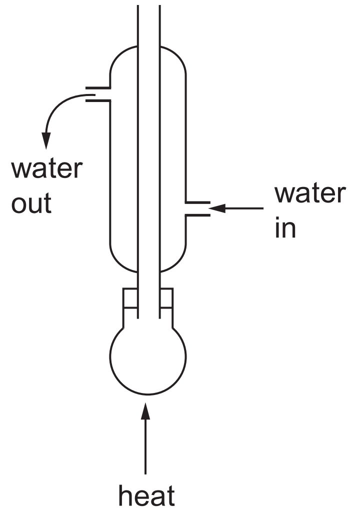
(ii) $q = m \times c \times \Delta T$
$50.0 \times 4.2 \times 4.8 = 1008\,J$ [2]
(units not required)

(iii) 1008 ÷ 1000 = 1.008 kJ [1]
1.008 ÷ 0.025 = –40.3 kJ mol^{-1}
(units not required) (minus sign required) [1] [5]

20

			AVAILABLE MARKS
4 **(a)**	mass = density × volume = 0.81 × 13.7 = 11.10 g	[1]	
(b)	11.10 ÷ 74 = 0.150 moles	[1]	
(c)	1:1 ratio 0.150 moles	[1]	
(d)	Actual yield = 10.28 ÷ 137 = 0.075 moles	[1]	
(e)	(Actual yield ÷ Theoretical yield) × 100 = % yield	[1]	
(f)	Percentage yield = (0.075 ÷ 0.150) × 100 = 50%	[1]	6

5 **(a)**



Condenser should have water in and water out as shown [3]
no heat source [–1]
no double jacket on condenser [–1]
condenser sealed at top [–1]
top of flask must not be open [–1]
no labels [–1]

(b) To ensure smooth boiling [1] 4

				AVAILABLE MARKS
6	**(a)**	Sodium hydroxide or NaOH	[1]	
	(b)	Potassium thiocyanate or KCNS	[1]	
	(c)	Blood red solution forms (If candidate gives the result of a positive test for part **(a)** this can be credited i.e. $Fe^{2+} \rightarrow$ green precipitate; $Fe^{3+} \rightarrow$ red/brown precipitate	[1]	3
7	**(a)**	Add water The layer which increases in volume is the aqueous layer Qualitative explanation using densities [2]	[1] [1]	
	(b)	Acts as a drying agent	[1]	3
			Section B	**36**
			Total	**90**

New Specification

Rewarding Learning

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
2009

Chemistry

Assessment Unit AS 3

assessing

Module 3: Practical Examination 2

[AC132]

FRIDAY 15 MAY, MORNING

MARK SCHEME

AVAILABLE MARKS

Section A

1 Titration exercise

(a) (i) Rinse out a pipette with one of the solutions and (using a pipette filler) transfer/pipette a known volume of the solution into a conical flask [1]
(ii) Add 2 or 3 drops of phenolphthalein [1]
(iii) Rinse out the burette with the other solution and fill the burette [1]
(iv) Add the solution from the burette until the end point is reached [1]
(v) Repeat (for accuracy) [1]

To a maximum of [4]

(b) Table [1]
Significant figures [2]
Calculation of average titre [2]
Titration consistency [3]
Agreement with supervisor's titre [4] [12]

NOTES:

Table:
Table should include initial burette reading, final burette reading, and volume delivered for rough and accurate titrations. [1]

Significant figures:
All burette readings should be to at least one decimal place – each mistake is penalised by [1]
(However, initial burette readings of 0 are penalised once only)
If used, the second decimal position should be 0 or 5 only – other values are penalised by [1] [2]

Average titre:
The average titre should be calculated and units should be included.
Accurate titrations only should be used.
The use of the rough value is [–1]
The average value can be two or three decimal places, e.g. 25.375/25.38
An incorrect calculation is 0. [2]

Titration consistency:
This is the difference between the first and second accurate readings.

Difference less than or equal to	Mark
0.1	3
0.2	2
0.3	1
0.4	0

[3]

Titration agreement with supervisor:
This is the difference between the candidate's calculated average titre and the supervisor's value.

Difference	Mark
±0.1	[4]
±0.2	[3]
±0.3	[2]
±0.4	[1]
±0.5	[0]

[4]

AVAILABLE MARKS

(c) Colourless to pink/red (or vice-versa depending on titration) [1]

(d) $CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l)$
[1] for equation, [1] for state symbols [2]

(e) **(i)** Correct calculation using volume in dm^3 [2]

$$\text{Number of moles} = \text{concentration } (mol\,dm^{-3}) \times \frac{\text{volume } (cm^3)}{1000}$$

(ii) Uses 1 : 1 ratio from equation
If equation incorrectly balanced in **(d)** the candidate's ratio should be used (i.e. carry error through (c.e.t.)) [1]

(iii) Correct calculation using volume in dm^3
Divide by volume (in dm^3) of ethanoic acid used in titration [2]

(iv) Correct calculation
Multiply by RMM of ethanoic acid [1] [6] 25

2 Observation/deduction

Safety goggles must be worn at all times and care should be exercised during this practical examination.

(a) You are provided with a mixture of two salts, labelled B, which have a common cation. Carry out the following experiments on the mixture. Record your observations and deductions in the spaces below and identify the two salts.

Experiment	Observations	Deductions	
1 Describe the appearance of B.	*White solid* [1]	*Does not contain a Transition metal ion. Group I or II/Ammonium compound/s-block ion present* [1]	[2]
2 Dip a wire loop in concentrated hydrochloric acid; touch sample B with the wire, then hold it in a blue Bunsen flame.	*Pink/Purple/Lilac* [1]	*K^+/Potassium ion/ compound (present)* [1]	[2]
In a fume cupboard: **3** Add about 1 cm^3 of concentrated sulphuric acid to a half spatula-measure of B in a test tube. Heat the test tube gently.	*Grey/black solid* [1] *Steamy fumes* [1] *Purple gas/clouds/fumes* [1]	*Iodine or iodide I^-* [1]	[4]
4 Make up a solution of B by dissolving a half spatula-measure of B in a test tube half-full of water. Put 1 cm^3 of the solution into each of two separate test tubes. **(a) (i)** Add a few drops of silver nitrate solution into the first test tube. **(ii)** Add about 2 cm^3 of concentrated ammonia into the first test tube. **(b)** Add a few drops of barium chloride solution to the second test-tube and then add 2 cm^3 of dilute nitric acid.	*Colourless solution* [1] *Yellow solid/ precipitate* [1] *Yellow precipitate remains* [1] *White solid precipitate* [1] *No effervescence/does not dissolve* [1]	 *Iodide* [1] *Sulphite/Sulphate* [1] *Not carbonate or hydrogen carbonate/sulphite* [1]	[8]

Name the two salts present in B:

Potassium iodide [1]

Potassium sulphate [1]

AVAILABLE MARKS

(b) You are provided with an aqueous solution containing an organic substance Y. Carry out the following experiments. Record your observations and deductions in the spaces below.

Experiment	Observations	Deductions	
1 Describe the solution and add a few drops on to Universal Indicator paper.	*Colourless (solution)* [1] *Shades of green* [1]	*Not a cartoxylic acid/Neutral* [1]	
In a fume cupboard: **2** Shake a small volume of the solution with bromine water.	*Yellow/orange colour* [1] *remains* [1]	*Saturated/no C═C* [1] *Not an alkene/not unsaturated*	[3]
3 Heat about 2 cm^3 of the solution with a few drops of acidified potassium dichromate solution.	*Orange* [1] *to green* [1] *Change in smell* [1]	*Primary or secondary alcohol/ not a tertiary alcohol/aldehyde* [1] *can be oxidised/aldehyde or ketone formed* *is a reducing agent* [1]	[3] [5]

AVAILABLE MARKS

Based on the above tests, suggest

A functional group(s) which may be present in Y:

OH [1]

A functional group(s) which is absent from Y:

C═C or —COOH [1]

To a maximum of [29] 29

Section A **54**

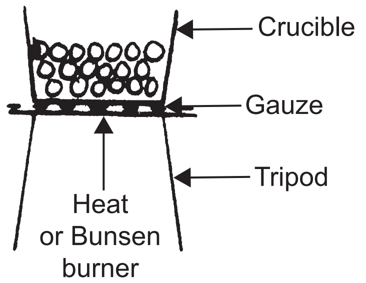
Section B

AVAILABLE MARKS

3 Planning

(a) **(i)** contains water of crystallisation [1]

(ii)



Heat/Bunsen burner – correctly labelled [1]
Tripod and gauze – correctly labelled [1]
Crucible – correctly labelled [1] [4]
([–1] for each omission)

(b) **(i)** mass of container, e.g. crucible [1]
Mass of crucible + hydrated zinc sulphate [1]

(ii) Heat and weigh [1]
Repeat [1]
Until there is no further decrease in mass/to constant mass [1]

(iii) Allow apparatus (crucible) to cool/use gloves/use tongs [1] [6]

(c) **(i)** 8.63 – 4.85 = 3.78 g [1]

(ii) 3.78 ÷ 18 = 0.21 moles [1]

(iii) 4.85 ÷ 161 = 0.03 moles [1]

(iv)

$ZnSO_4$	H_2O	
0.03	0.21	
1	7	[1]

$x = 7$ [1] [5]

(d) **(i)** $0.5 \times (50.0 \div 1000) = 0.025$ mole [1]

(ii) $q = m \times c \times \Delta T$ [1]
$50 \times 4.2 \times 25.3 = 5313$ J [1]
(units not required)

(iii) $5313 \div 1000 = 5.313$ kJ [1]
$5.313 \div 0.025 = -212.5$ kJ mol^{-1} [1] [5] 20
(units not required) (minus sign required)

AVAILABLE MARKS

4 **(a)** mass = density × volume = 1.05 × 11.4 = 11.97 g [1]

(b) 11.97 ÷ 60 = 0.200 moles [1]

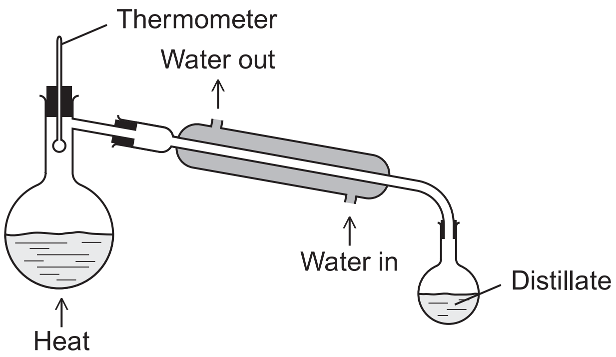
(c) 0.200 moles [1]

(d) Actual yield = 5.28 ÷ 88 = 0.06 moles [1]

(e) (Actual yield ÷ Theoretical yield) × 100 = % yield [1]

(f) Percentage yield = (0.06 ÷ 0.200) × 100 = 30% [1] 6

5 **(a)** [2]



Condenser should have water in and water out as shown [1]
no heat source [–1]
no double jacket on condenser [–1]
distillation flask open at top [–1]
thermometer missing [–1]
no labels [–1]

(b) To ensure smooth boiling [1] 4

6 **(a)** Ammonia solution/$NH_3(aq)$ [1]

(b) (pale) blue precipitate/solid forms [1]
precipitate dissolves in excess to form a (deep) blue solution [1] 3

7 **(a)** Add water [1]
the layer which does not increase in volume is the organic layer [1]
Qualitative explanation using densities [2]

(b) Acts as a drying agent [1] 3

Section B	**36**
Total	**90**

New Specification

Rewarding Learning

Centre Number

71	

Candidate Number

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
January 2010

Chemistry

Assessment Unit AS 1

assessing

Basic Concepts in Physical and Inorganic Chemistry

[AC111]

AC111

THURSDAY 14 JANUARY, MORNING

TIME

1 hour 30 minutes.

INSTRUCTIONS TO CANDIDATES

Write your Centre Number and Candidate Number in the spaces provided at the top of this page.
Answer **all fifteen** questions.
Answer **all ten** questions in **Section A**. Record your answers by marking the appropriate letter on the answer sheet provided. Use only the spaces numbered 1 to 10. Keep in sequence when answering.
Answer **all five** questions in **Section B**. Write your answers in the spaces provided in this question paper.

INFORMATION FOR CANDIDATES

The total mark for this paper is 100.
Quality of written communication will be assessed in question **13(c)**.
In Section A all questions carry equal marks, i.e. **two** marks for each question.
In Section B the figures in brackets printed down the right-hand side of pages indicate the marks awarded to each question or part question.
A Periodic Table of Elements (including some data) is provided.

For Examiner's use only	
Question Number	**Marks**
Section A	
1–10	
Section B	
11	
12	
13	
14	
15	
Total Marks	

Section A

For each of the following questions only **one** of the lettered responses (A–D) is correct.

Select the correct response in each case and mark its code letter by connecting the dots as illustrated on the answer sheet.

1 2.65 g of anhydrous sodium carbonate, Na_2CO_3, was dissolved in water and the solution made up to $250\,cm^3$ in a volumetric flask. The concentration of the solution was

A $0.025\,mol\,dm^{-3}$
B $0.050\,mol\,dm^{-3}$
C $0.100\,mol\,dm^{-3}$
D $0.200\,mol\,dm^{-3}$

2 In which one of the following molecules does the central atom obey the octet rule?

A $BeCl_2$
B BF_3
C CF_4
D SF_6

3 Which one of the following statements about iodine is **not** correct?

A It has a molecular covalent structure.
B It contains non-polar molecules.
C It exists as a grey-black shiny solid.
D It is more soluble in water than hexane.

4 Elements Q and R have ground state electron structures $1s^22s^22p^63s^2$ and $1s^22s^22p^5$ respectively. Q and R combine to produce a compound with the formula

A QR
B QR_2
C Q_2R
D Q_2R_5

5 Which one of the following molecules is polar?

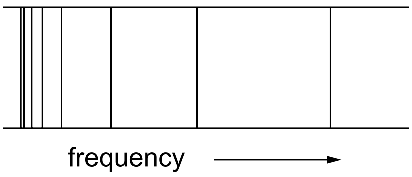
A BF_3
B CF_4
C OF_2
D F_2

6 4.88 g of hydrated barium chloride, $BaCl_2.xH_2O$, was heated to a constant mass of 4.16 g. What is the value of x?

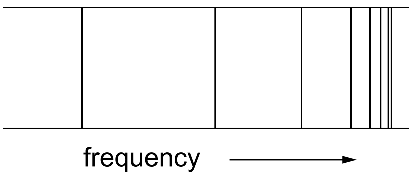
A 1
B 2
C 3
D 4

7 Which one of the following represents the emission spectrum of atomic hydrogen in the ultraviolet region?

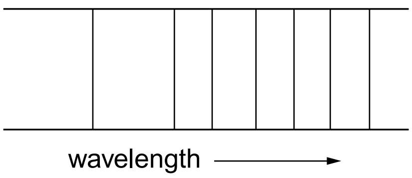
A



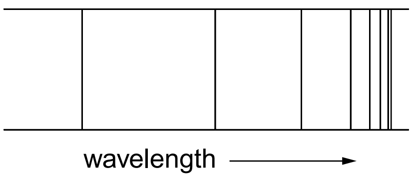
B



C



D



8 When burned in a plentiful supply of oxygen, propane (C_3H_8) produces carbon dioxide and water.

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

What is the number of molecules of carbon dioxide produced when 4.4 g of propane are burned?

A 6.02×10^{22}
B 1.81×10^{23}
C 6.02×10^{23}
D 1.81×10^{24}

9 A compound produces a lilac colour in a flame test. When chlorine is bubbled into an aqueous solution of the compound, the solution changes from colourless to yellow-orange. The compound is

A potassium bromide
B potassium iodide
C sodium bromide
D sodium iodide

10 Iron(III) oxide can be reduced by carbon to form iron.

$$2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$$

What is the maximum mass of iron which can be produced when 3.20 kg of iron(III) oxide is heated with 0.72 kg of carbon?

A 1.12 kg
B 2.24 kg
C 3.36 kg
D 4.48 kg

Section B

Answer **all five** questions in this section.

Examiner Only	
Marks	Remark

11 There are five isotopes of germanium.

(a) Atoms of the ^{74}Ge isotope contain 32 protons, 32 electrons and 42 neutrons. Complete the following table which shows the properties of each of these particles.

Particle	Relative mass	Relative charge
Proton		
Electron		
Neutron		

[3]

(b) State, in terms of protons and neutrons, the meanings of the following terms:

Mass number ____

____ [1]

Atomic number ____

____ [1]

Isotopes ____

____ [1]

(c) The terms atomic number and mass number can be used to deduce the numbers of protons, neutrons and electrons in an atom or ion.

Examiner Only

Marks | Remark

(i) An atom has 15 protons and 20 neutrons fewer than there are in the ^{70}Ge isotope. Deduce the symbol and mass number of this atom.

Symbol ______________ Mass Number ______________ [2]

(ii) Complete the table for the ions of the elements X, Y and Z. The letters X, Y and Z are not the symbols of the elements.

Ion	Atomic Number	Mass Number	Number of Neutrons	Electronic Structure
X^{2+}			20	$1s^2 2s^2 2p^6 3s^2 3p^6$
Y^-			18	$1s^2 2s^2 2p^6 3s^2 3p^6$
Z^{2-}		16		$1s^2 2s^2 2p^6$

[3]

(d) The mass spectrum of germanium is used to calculate its relative atomic mass.

(i) Define the term **relative atomic mass**.

____

____ [2]

(ii) The table below gives the percentage abundance of each isotope in the mass spectrum of germanium.

Relative Isotopic Mass	**70**	**72**	**73**	**74**	**76**
% Abundance	20.55	27.37	7.67	36.74	7.67

Use this information to calculate the relative atomic mass of germanium to one decimal place.

____

____

____ [2]

Examiner Only

Marks	Remark

12 Phosphorus and nitrogen are in Group V of the Periodic Table. Nitrogen forms a hydride called ammonia and the hydride of phosphorus is called phosphine, PH_3.

Examiner Only	
Marks	Remark

(a) (i) Draw a dot and cross diagram to show the bonding in phosphine.

[2]

(ii) Draw and name the shape of a phosphine molecule.

____ [2]

(iii) Explain why a phosphine molecule has the shape you have drawn.

____

____

____

____ [2]

(iv) Suggest a value for the bond angles in phosphine. Explain your answer.

____

____

____

____ [2]

(b) Phosphine reacts with hydrogen ions to form phosphonium ions, PH_4^+. Ammonia similarly forms ammonium ions, NH_4^+.

(i) Name the type of bond formed when phosphine reacts with a hydrogen ion.

____ [1]

(ii) Explain how this bond is formed.

____

____ [2]

(c) Suggest, in terms of intermolecular forces, why ammonia has a higher boiling point than phosphine.

____

____

____

____ [2]

Examiner Only	
Marks	Remark

13 Analysis of a vinegar solution was carried out using the following procedure:

Transfer 25.0 cm^3 of undiluted vinegar into a 250 cm^3 volumetric flask and make the solution up to the mark using de-ionised water. Transfer 25.0 cm^3 portions of the diluted vinegar into three separate conical flasks and add a few drops of indicator to each flask. Titrate each solution with 0.1 mol dm^{-3} sodium hydroxide until an end point is reached.

A student obtained the following results:

	Initial burette reading (cm^3)	**Final burette reading (cm^3)**	**Titre (cm^3)**
Rough	0.0	21.7	21.7
1st accurate	21.7	43.1	
2nd accurate	0.0	21.3	

(a) (i) Name a suitable indicator for this titration.

____ [1]

(ii) State the colour change which would be obtained at the end point.

from ____________________ to ____________________ [2]

(b) (i) Write the equation for the reaction between vinegar (ethanoic acid) and sodium hydroxide.

____ [2]

(ii) Complete the results table and calculate the average titre.

____ [2]

(iii) Use the average titre to calculate the number of moles of sodium hydroxide used in the titration.

____ [1]

Examiner Only

Marks	Remark

(iv) Calculate the concentration of ethanoic acid in the diluted vinegar.

_______________________________ $mol\,dm^{-3}$ [1]

(v) Calculate the concentration of ethanoic acid in the undiluted vinegar.

_______________________________ $mol\,dm^{-3}$ [1]

(c) Describe, giving practical details, how you would prepare the solution of diluted vinegar and then transfer $25.0\,cm^3$ to a conical flask.

____

____

____

____

____

____

_______________________________ [4]

Quality of written communication [2]

Examiner Only	
Marks	Remark

Examiner Only
Marks	Remark

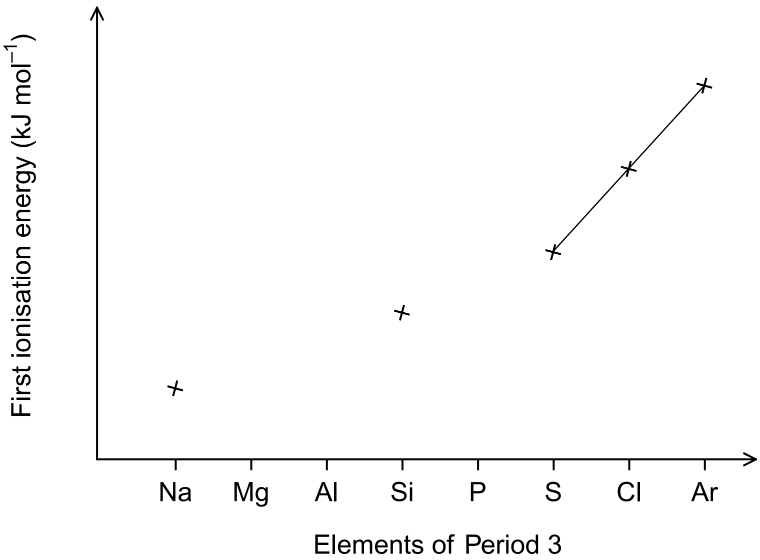
14 Ionisation energies provide evidence for the existence of shells and subshells in atoms.

(a) State the meaning of the term *first ionisation energy* of an element.

____

____ [2]

(b) There is a general increase in the first ionisation energies across Period 3. The graph below shows the variation of the first ionisation energies of some of the elements in Period 3.



(i) Use crosses to mark the relative positions of the first ionisation energies for the elements Mg, Al and P. Complete the graph by joining the crosses. [2]

(ii) Explain the general increase in first ionisation energy across the period.

____

____ [2]

(iii) Using s, p and d notation give the ground state electronic configuration of a magnesium atom.

____ [1]

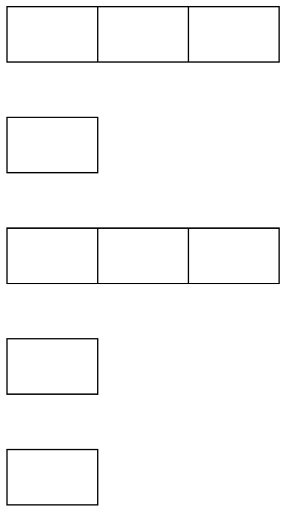
(iv) Explain the position of the first ionisation energy of magnesium relative to that of aluminium in your graph.

____

____ [2]

(c) The first four ionisation energies of aluminium are 578, 1817, 2745 and 11 578 kJ mol^{-1}.

(i) Label the subshells in the following diagram for an aluminium atom and use the electrons-in-boxes notation to show how the electrons are arranged in the Al^{2+} ion.



[2]

(ii) Write the equation, including state symbols, for the fourth ionisation energy of aluminium.

____ [2]

(iii) Explain why the third ionisation energy of aluminium is much smaller than the fourth ionisation energy.

____

____ [2]

Examiner Only

Marks	Remark

15 The elements in Group VII are all reactive non-metals.

Examiner Only	
Marks	Remark

(a) There is a trend in the electronegativity of the elements in Group VII.

(i) Define the term **electronegativity**.

____

____

____ [2]

(ii) State and explain the trend in the electronegativity of the elements down Group VII from fluorine to iodine.

____

____

____

____ [3]

(b) When concentrated sulphuric acid is added to solid sodium bromide, the acid reacts with bromide ions to form sulphur dioxide and bromine.

(i) State the change in the oxidation number of sulphur in this reaction.

____ [2]

(ii) Write the half-equation to show how bromine is formed from bromide ions.

____ [1]

(iii) Complete the half-equation to show how sulphur dioxide is formed from sulphuric acid.

$$H_2SO_4 \quad + \quad H^+ \qquad\qquad \rightarrow \qquad SO_2 \quad + \quad H_2O$$

[2]

(iv) Write the overall ionic equation for the reaction of bromide ions with sulphuric acid.

____ [1]

(v) State **one** observation in the above reaction.

____ [1]

(vi) State the role of bromide ions in this reaction.

____ [1]

(vii) Why are sulphur dioxide and chlorine not formed when concentrated sulphuric acid is added to solid sodium chloride?

____

____ [1]

(viii) Write the equation for the reaction of solid sodium chloride with concentrated sulphuric acid.

____ [2]

(c) When sodium bromide is dissolved in water, the presence of bromide ions can be established by using aqueous silver nitrate followed by concentrated ammonia solution.

(i) What is observed when aqueous silver nitrate is added to sodium bromide solution?

____ [2]

(ii) Write the ionic equation, including state symbols, for the reaction.

____ [2]

(iii) What is observed when an excess of concentrated ammonia solution is added?

____ [1]

Examiner Only	
Marks	Remark

Published Mark Schemes for GCE AS Chemistry

January 2010

Issued: April 2010

NORTHERN IRELAND GENERAL CERTIFICATE OF SECONDARY EDUCATION (GCSE) AND NORTHERN IRELAND GENERAL CERTIFICATE OF EDUCATION (GCE)

MARK SCHEMES (2010)

Foreword

Introduction

Mark Schemes are published to assist teachers and students in their preparation for examinations. Through the mark schemes teachers and students will be able to see what examiners are looking for in response to questions and exactly where the marks have been awarded. The publishing of the mark schemes may help to show that examiners are not concerned about finding out what a student does not know but rather with rewarding students for what they do know.

The Purpose of Mark Schemes

Examination papers are set and revised by teams of examiners and revisers appointed by the Council. The teams of examiners and revisers include experienced teachers who are familiar with the level and standards expected of 16- and 18-year-old students in schools and colleges. The job of the examiners is to set the questions and the mark schemes; and the job of the revisers is to review the questions and mark schemes commenting on a large range of issues about which they must be satisfied before the question papers and mark schemes are finalised.

The questions and the mark schemes are developed in association with each other so that the issues of differentiation and positive achievement can be addressed right from the start. Mark schemes therefore are regarded as a part of an integral process which begins with the setting of questions and ends with the marking of the examination.

The main purpose of the mark scheme is to provide a uniform basis for the marking process so that all the markers are following exactly the same instructions and making the same judgements in so far as this is possible. Before marking begins a standardising meeting is held where all the markers are briefed using the mark scheme and samples of the students' work in the form of scripts. Consideration is also given at this stage to any comments on the operational papers received from teachers and their organisations. During this meeting, and up to and including the end of the marking, there is provision for amendments to be made to the mark scheme. What is published represents this final form of the mark scheme.

It is important to recognise that in some cases there may well be other correct responses which are equally acceptable to those published: the mark scheme can only cover those responses which emerged in the examination. There may also be instances where certain judgements may have to be left to the experience of the examiner, for example, where there is no absolute correct response – all teachers will be familiar with making such judgements.

The Council hopes that the mark schemes will be viewed and used in a constructive way as a further support to the teaching and learning processes.

CONTENTS

New Specification

Rewarding Learning

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
January 2010

Chemistry

Assessment Unit AS 1

assessing

Basic Concepts in Physical and Inorganic Chemistry

[AC111]

THURSDAY 14 JANUARY, MORNING

MARK SCHEME

Section A

			AVAILABLE MARKS
1	C		
2	C		
3	D		
4	B		
5	C		
6	B		
7	B		
8	B		
9	A		
10	B		
[2] for each correct answer		[20]	20
		Section A	**20**

Section B

AVAILABLE MARKS

11 **(a)**

Particle	Relative mass	Relative charge
Proton	1	+ 1
Electron	$\frac{1}{1840}$	– 1
Neutron	1	0

1 mark per particle [3]

(b) Number of protons + Number of neutrons [1]

Number of protons [1]

Atoms which have the same number of protons but different numbers of neutrons [1]

(c) **(i)** Cl [1] 35 [1] [2]

(ii)

Ion	Atomic Number	Mass Number	Number of Neutrons	Electronic Structure
X^{2+}	**20**	**40**	20	$1s^22s^22p^63s^23p^6$
Y^-	**17**	**35**	18	$1s^22s^22p^63s^23p^6$
Z^{2-}	**8**	16	**8**	$1s^22s^22p^6$

1 mark per particle [3]

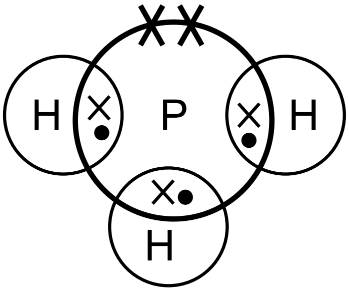
(d) **(i)** The average mass of an atom of an element [1] relative to (one twelfth of) the mass of an atom of carbon-12. [1] [2]

(ii) [(70 × 20.55) + (72 × 27.37) + (73 × 7.67) + (74 × 36.74) + (76 × 7.67)] ÷ 100

72.7 [2] 15

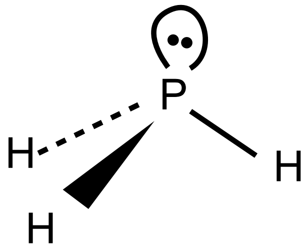
	AVAILABLE MARKS

12 (a) (i)



[2]

(ii)



[1]

Pyramidal [1]

(iii) In the outer shell of the central P atom there are (3) bonding pairs and (1) lone pair of electrons [1] which repel [1] [2]

(iv) approximately 107° [1]
Lone pair-bonding pair repulsion is greater than bonding pair-bonding pair repulsion [1] [2]

(b) (i) coordinate (dative covalent) bond [1]

(ii) the phosphorus atom shares its lone pair with the H^+ ion [2]

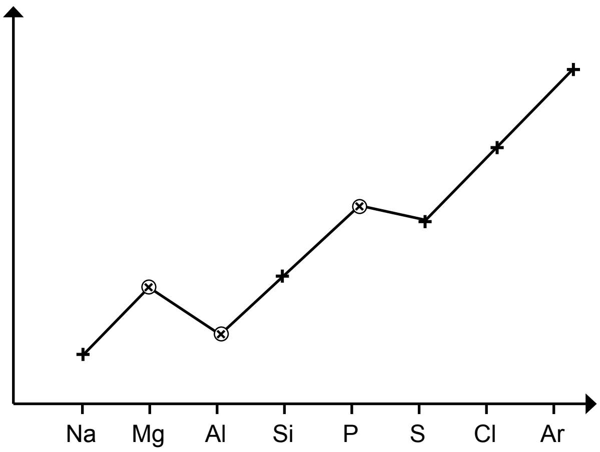
(c) Hydrogen bonding in ammonia [1]
is stronger than dipole-dipole (or van der Waals') in phosphine [1] [2] 13

					AVAILABLE MARKS
13	**(a)**	**(i)**	Phenolphthalein	[1]	
		(ii)	colourless [1] to pink [1]	[2]	
	(b)	**(i)**	$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$	[2]	
		(ii)	21.4 and 21.3 into table Average titre = 21.35 cm^3	[1] [1]	
		(iii)	$0.1 \times (21.35/1000) = 2.135 \times 10^{-3}$	[1]	
		(iv)	$(2.135 \times 10^{-3}) \div (0.025) = 0.0854$	[1]	
		(v)	0.854	[1]	
	(c)	Rinsing the pipette (2-3 times) with undiluted vinegar Use of pipette filler Meniscus on mark Mixing/inverting solution in volumetric flask Rinsing the pipette (2-3 times) with diluted vinegar Max = 4		[4]	
		Quality of written communication		[2]	16

14 **(a)** Energy required to convert one mole of gaseous atoms to gaseous ions with single positive charges. [2]

AVAILABLE MARKS

(b) **(i)**



Point for Mg above Na but not above that for Si
point for Al below that for Mg but not below Na [1]

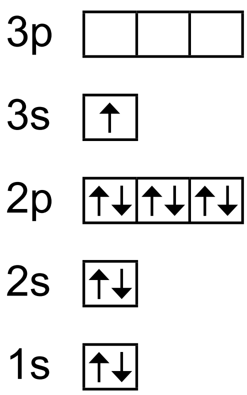
Point for P above that for S and below that for Cl [1]

(ii) increasing nuclear charge [1]
Shielding approximately constant/atomic radius decreases [1]

(iii) $1s^2 2s^2 2p^6 3s^2$ [1]

(iv) (stability of) filled subshell for Mg atom [1]
Stability not present in aluminium/$3p^1$/outer electron is in higher energy subshell [1] [2]

(c) **(i)**



Labels [1] Configuration [1] [2]

(ii) Al^{3+} (g) $\rightarrow$ Al^{4+} (g) + e^- [2]

(iii) Outer electron in Al^{2+} is further from nucleus/more shielded/ stability of sub shell full Al^{3+}(higher energy level) 2 from 3 [2] 15

					AVAILABLE MARKS
15	**(a)**	**(i)**	The ability/power of an atom to attract bonding electrons towards itself in a covalent bond	[2]	
		(ii)	decreases [1] Shared pair is further from nucleus [1] more shielded [1]	[3]	
	(b)	**(i)**	+6 [1] to +4 [1]	[2]	
		(ii)	$2Br^- \rightarrow Br_2 + 2e^-$	[2]	
		(iii)	$H_2SO_4 + \mathbf{2}H^+ + \mathbf{2e^-} \rightarrow SO_2 + \mathbf{2}H_2O$	[1]	
		(iv)	$H_2SO_4 + 2H^+ + 2Br^- \rightarrow SO_2 + 2H_2O + Br_2$	[1]	
		(v)	red-brown gas	[1]	
		(vi)	reducing agent	[1]	
		(vii)	chloride ions are not strong enough reducing agents to reduce concentrated sulphuric acid	[1]	
		(viii)	$NaCl + H_2SO_4 \rightarrow NaHSO_4 + HCl$	[2]	
	(c)	**(i)**	cream [1] precipitate [1]	[2]	
		(ii)	$Ag^+ (aq) + Br^- (aq) \rightarrow AgBr (s)$ Equation [1] State symbols [1]	[2]	
		(iii)	cream precipitate dissolves	[1]	21
				Section B	**80**
				Total	**100**

New Specification

Rewarding Learning

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
January 2010

Chemistry

Assessment Unit AS 2

assessing

Module 2: Further Organic, Physical and Inorganic Chemistry and Introduction to Organic Chemistry

[AC121]

THURSDAY 21 JANUARY, AFTERNOON

MARK SCHEME

Section A

			AVAILABLE MARKS
1	A		
2	C		
3	C		
4	C		
5	D		
6	D		
7	A		
8	C		
9	C		
10	D		
[2] marks for each correct answer		[20]	20
		Section A	**20**

AVAILABLE MARKS

Section B

11 **(a)** CH_3CH_2Cl

CH_3CH_2ONa

CH_3CH_2Cl

CH_3CH_2Br [4] 4

12 **(a)** **(i)** molecule broken down (into a smaller molecule) [1]
using heat (and a catalyst) [1]

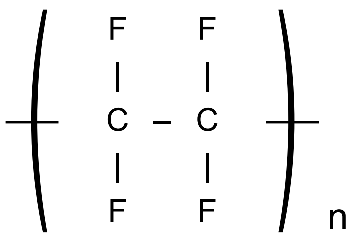
(ii) endothermic as ΔH is +ve [1]

(iii) low pressure [1]
more molecules on RHS [1]
but reaction rate is (slightly) slower [1]

(iv) high temperature
reaction is endothermic/ absorbs heat → RHS
and rate is increased [3]

(v) conc ammonia (solution) [1]
white smoke/fumes [1]

(b) **(i)**



[2]

(ii) contains a double bond/ C = C/ unsaturated [1]

(iii) addition [1] 15

			AVAILABLE MARKS
13 **(a)**	C_nH_{2n+2}	[1]	
(b)	same molecular formula different arrangement of atoms etc	[1] [1]	
(c)	fractional distillation	[2]	
(d)	2 – methylbutane or methylbutane 2, 2 – dimethylpropane or dimethylpropane	[1] [1]	
(e)	van der Waals forces greatest between chains more branched lower b.pt	[2]	
(f) **(i)**	$C_5H_{12} + 8O_2 \rightarrow 5CO_2 + 6H_2O$	[2]	
(ii)	$2C_5H_{12} + 11O_2 \rightarrow 10CO + 12H_2O$	[2]	

(g) **(i)**

$$\left.\begin{array}{llll} Cl & \frac{60.7}{35.5} = 1.71 & 1 & \\ C & \frac{34.2}{12} = 2.85 & 1.66 & 1\frac{2}{3} = \frac{5}{3} \\ H & \frac{5.1}{1} = 5.1 & 3 & \end{array}\right\} [1]$$

$$1 : \frac{5}{3} : 3$$

$$\therefore\ 3 : 5 : 9$$

$C_5H_9Cl_3$ or $Cl_3C_5H_9$ [3]

(ii) has a higher b.pt. because chloropentane has a greater mass than pentane/ polar /more van der Waals [1] 17

			AVAILABLE MARKS
14 **(a)**	more stable than ($BeCO_3$), $MgCO_3$	[1]	
	less stable than $SrCO_3$/ $BaCO_3$	[1]	
(b)	the size of cation increases down Gp = Group passim [1] charge density decreases down Gp polarisation power decreases down Gp max of 2	[2]	
(c)	produces CO_2 $CaCO_3 \rightarrow CaO + CO_2$	[1]	
	and CO_2 from combustion of fossil fuels	[1]	
(d)	**(i)** $CaO + 2HCl \rightarrow CaCl_2 + H_2O$	[2]	
	(ii) $CaO + H_2O \rightarrow Ca(OH)_2$	[1]	
(e)	add $Ca(OH)_2$ to water stir / until no more dissolves filter blow CO_2 through limewater/collect CO_2 in pipette/ add to container of CO_2 seal and shake goes milky/chalky/cloudy	[4]	
	max of 4		
	QWC	[2]	
(f)	$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$	[2]	
(g)	100g of $CaCO_3$ $\frac{0.05}{100} = 5 \times 10^{-4}$ mol of $CaCO_3$ 5×10^{-4} mol CO_2 $5 \times 10^{-4} \times 24\ dm^3 = 0.012\ dm^3$ $\% CO_2 = \frac{0.012}{5} \times 100 = 0.24\%$	[5]	22

AVAILABLE MARKS

15	**(a)**	**(i)**	$C_2H_4 + H_2O \rightarrow C_2H_5OH$	[1]
			steam/ high temp	[1]
			H_3PO_4	[1]
		(ii)	fermentation	[1]
			sugars	[1]

(b)

acidified dichromate		Iodine	[1]
warm	**or**	alkali	[1]
(orange) → green		yellow solid	[1]

(c)

bonds broken	5C—H	= 5 × 413	= 2065
	1C—C	= 1 × 347	= 347
	1C—O	= 1 × 360	= 360
	1O—H	= 1 × 464	= 464
	3O=O	= 3 × 498	= 1494
bonds formed	4C=O	= 4 × 805	= 3220
	6O—H	= 6 × 464	= 2784

bonds broken = +4730
bonds formed = –6004

$\Delta H = -1274\ kJ\ mol^{-1}$ [4]

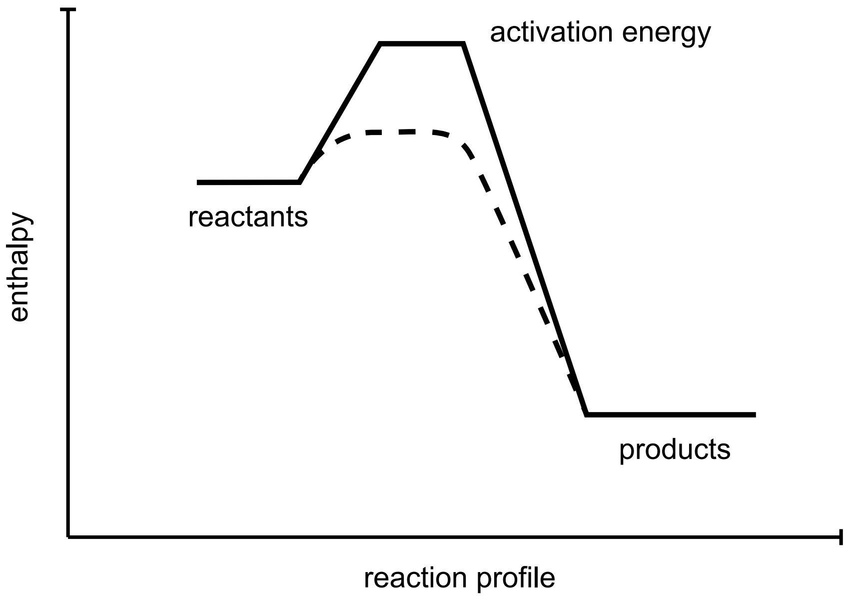
AVAILABLE MARKS

(d) **(i)** $CO \rightarrow CO_2$ [1]

$CxHy \rightarrow CO_2 + H_2O$ [1]

$NO_x \rightarrow N_2 (+CO_2)$ [1]

(ii) lead [1]
covers the surface [1]
molecules prevented from chemisorption/ catalyst effect reduced [1]



activation energy lowered by the catalyst [4] 22

Section B	**80**
Total	**100**

New Specification

Rewarding Learning

Centre Number

71	

Candidate Number

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
2010

Chemistry

Assessment Unit AS 1

assessing

Basic Concepts in Physical and Inorganic Chemistry

[AC111]

MONDAY 7 JUNE, MORNING

AC111

TIME

1 hour 30 minutes.

INSTRUCTIONS TO CANDIDATES

Write your Centre Number and Candidate Number in the spaces provided at the top of this page.
Answer **all fifteen** questions.
Answer **all ten** questions in **Section A**. Record your answers by marking the appropriate letter on the answer sheet provided. Use only the spaces numbered 1 to 10. Keep in sequence when answering.
Answer **all five** questions in **Section B**. Write your answers in the spaces provided in this question paper.

INFORMATION FOR CANDIDATES

The total mark for this paper is 100.
Quality of written communication will be assessed in question **13(d)**.
In Section A all questions carry equal marks, i.e. **two** marks for each question.
In Section B the figures in brackets printed down the right-hand side of pages indicate the marks awarded to each question or part question.
A Periodic Table of Elements (including some data) is provided.

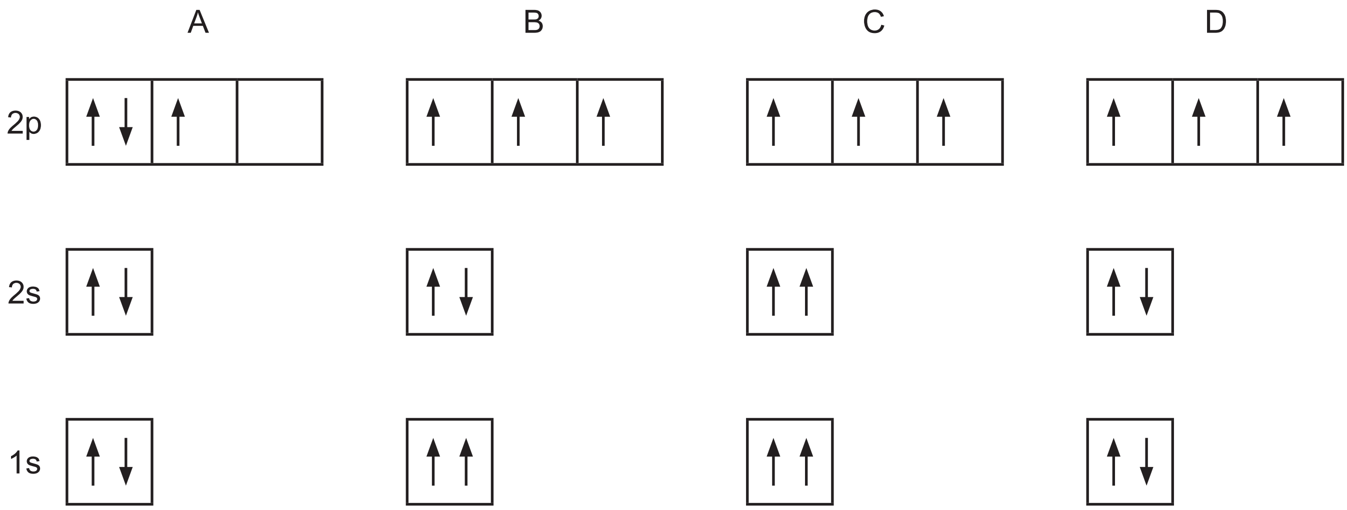
For Examiner's use only	
Question Number	**Marks**
Section A	
1–10	
Section B	
11	
12	
13	
14	
15	
Total Marks	

Section A

For each of the following questions only **one** of the lettered responses (A–D) is correct.

Select the correct response in each case and mark its code letter by connecting the dots as illustrated on the answer sheet.

1 Which one of the following represents the ground state electronic configuration of a nitrogen atom?



2 When burned in a plentiful supply of oxygen, methane produces carbon dioxide and water.

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

What is the number of molecules of oxygen required for the complete combustion of 1.6 g of methane?

A 6.0×10^{22}
B 1.2×10^{23}
C 6.0×10^{23}
D 1.2×10^{24}

3 In which one of the following molecules are the bond angles closest to 107°?

A BF_3
B CH_4
C H_2O
D NH_3

4 Which one of the following gives the correct flame colour for the named compound?

A	barium chloride	red
B	copper(II) chloride	blue-green
C	potassium chloride	yellow/orange
D	sodium chloride	lilac

5 An element was analysed using a mass spectrometer. The spectrum showed that there were four isotopes. The relative isotopic masses and relative abundances are given below.

Relative isotopic mass	**Relative abundance**
50	2
52	35
53	4
54	1

The relative atomic mass of this element is

A 52.00.
B 52.05.
C 52.25.
D 52.50.

6 Which one of the following is produced when concentrated sulphuric acid reacts with solid sodium chloride?

A chlorine
B hydrogen chloride
C hydrogen sulphide
D sulphur dioxide

7 Titanium is extracted in a two-stage process. The first stage involves the conversion of titanium(IV) oxide to titanium(IV) chloride. In the second stage, the titanium(IV) chloride is reduced using magnesium.

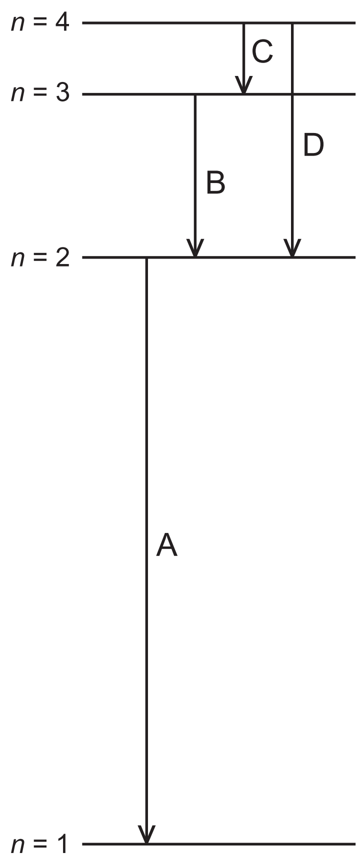
$$TiO_2 + C + 2Cl_2 \rightarrow TiCl_4 + CO_2$$

$$TiCl_4 + 2Mg \rightarrow Ti + 2MgCl_2$$

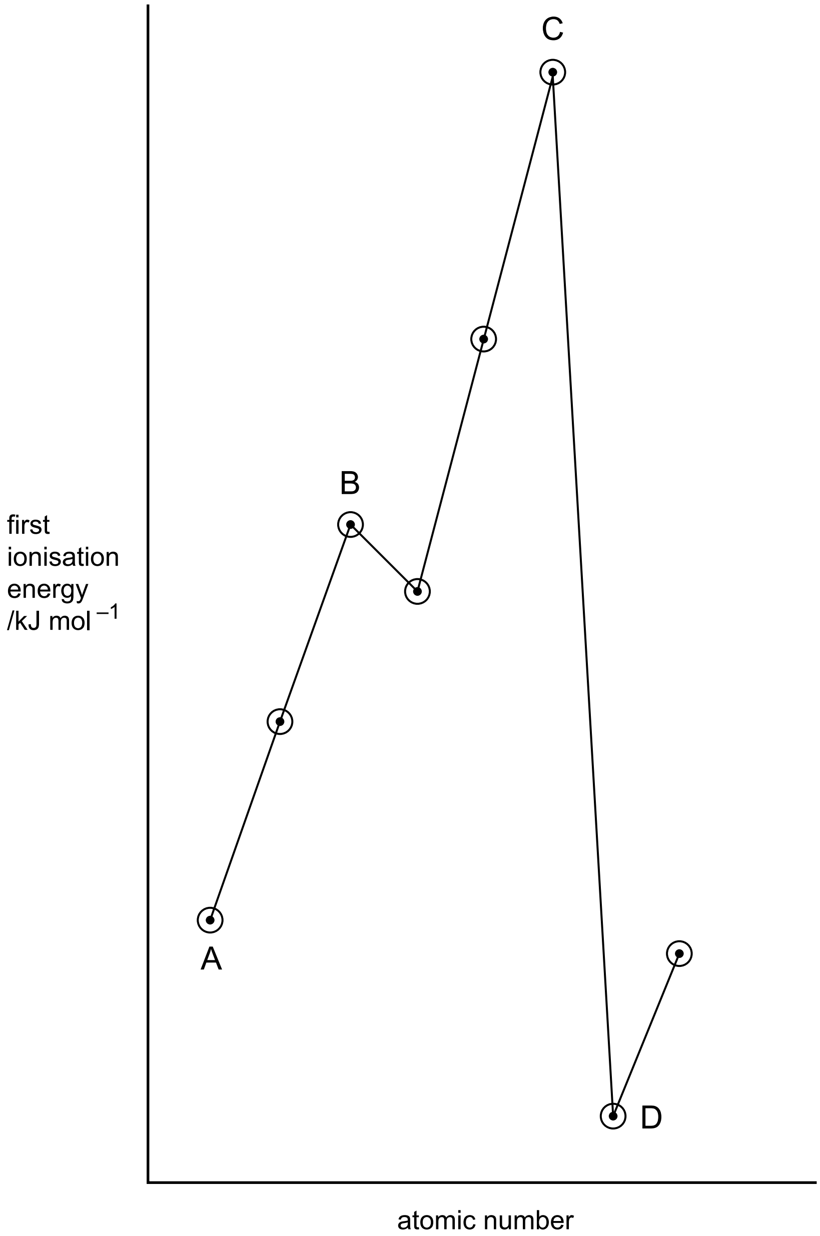
What is the maximum mass of titanium which can be obtained when 8.0 kg of titanium(IV) oxide is converted to titanium(IV) chloride and then reduced using 7.2 kg of magnesium?

A 2.4 kg
B 4.8 kg
C 9.6 kg
D 14.4 kg

8 Which one of the following electron transitions is responsible for the lowest frequency line in the visible region of the emission spectrum of atomic hydrogen?



9 Which one of the letters represents the first ionisation energy of an alkali metal?



10 The ionisation energy of hydrogen is 1312 kJ mol^{-1}. Use this value to calculate the frequency at convergence in the hydrogen emission spectrum.

A 2.179×10^{-21} Hz
B 1.312×10^{6} Hz
C 3.287×10^{15} Hz
D 1.979×10^{36} Hz

Section B

Answer **all five** questions in this section.

Examiner Only	
Marks	Remark

11 Aluminium is the most abundant metal in the Earth's crust. It is used in electrical cables and is present in high strength alloys.

(a) All atoms of aluminium have a mass number of 27.
How many neutrons are present in the nucleus of these atoms?

____ [1]

(b) (i) Write the equation, including state symbols, which represents the first ionisation energy of aluminium.

____ [2]

(ii) Explain why the first ionisation energy of boron has a larger value than the first ionisation energy of aluminium.

____

____

____ [2]

(iii) Explain why the first ionisation energy of magnesium has a larger value than the first ionisation energy of aluminium.

____

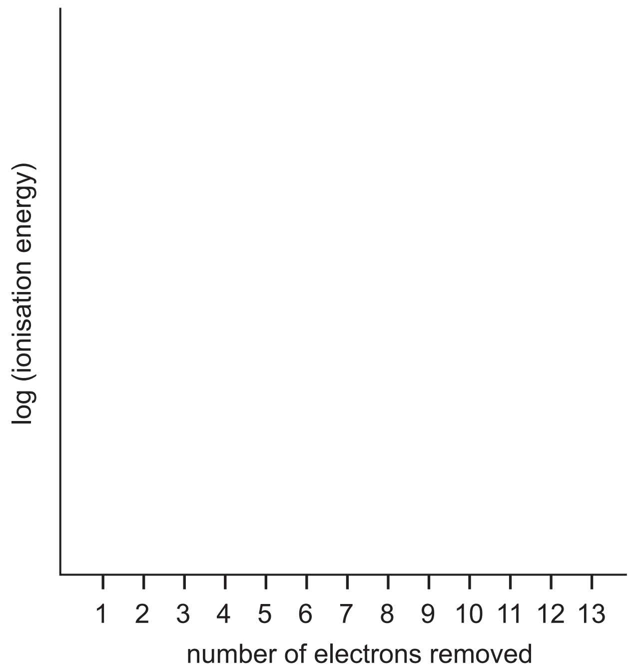
____

____ [2]

(iv) Give the ground state electronic configuration of the Al^{4+} ion.

____ [1]

(v) Sketch a graph to show the successive ionisation energies of aluminium.



[2]

Examiner Only

Marks	Remark

12 Fluorine is the most reactive non-metallic element. It combines with both metals and non-metals.

(a) **(i)** Using dot and cross diagrams, explain how strontium atoms combine with fluorine atoms to form strontium fluoride. Show the outer electrons only.

[4]

(ii) Compare the electrical conductivity of solid strontium metal with that of solid strontium fluoride. Explain your answer.

____

____

____

____ [3]

Examiner Only	
Marks	Remark

Examiner Only	
Marks	Remark

(b) Sulphur and fluorine combine to form a non-polar molecule sulphur hexafluoride, SF_6.

(i) Define the term **electronegativity**.

____

____ [2]

(ii) Label the diagram below to show the polarity of the S—F bond.

S—F

[1]

(iii) Draw a dot and cross diagram to show the bonding in SF_6 using outer shell electrons only.

[2]

(iv) Explain whether the SF_6 molecule obeys the octet rule.

____

____

____ [2]

(v) Draw, name and explain the shape of the SF_6 molecule.

____

____

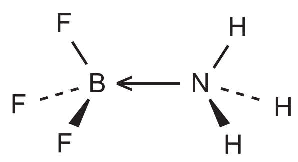
____ [4]

(vi) Suggest why SF_6 is a non-polar molecule, even though it contains polar bonds.

____

____ [2]

(c) Boron trifluoride can combine with ammonia to form the following molecule.



(i) Name the type of bond formed between the boron and nitrogen atoms.

____ [1]

(ii) Explain how this bond is formed.

____

____ [1]

13 Concentrated nitric acid (HNO_3) oxidises iodide ions to form iodine. In the reaction the nitric acid is reduced to form nitrogen monoxide (NO).

Examiner Only	
Marks	Remark

(a) Reduction and oxidation can be defined in different ways.

(i) Define oxidation in terms of electron transfer.

____ [1]

(ii) Define reduction in terms of changes in oxidation state.

____ [1]

(b) Deduce the oxidation number of nitrogen in

(i) HNO_3 _______________________________________ [1]

(ii) NO _______________________________________ [1]

(c) The half-equation for the reduction of concentrated nitric acid is shown below.

$$HNO_3 + 3H^+ + 3e^- \rightarrow NO + 2H_2O$$

(i) Write a half-equation for the oxidation of iodide ions to form an iodine molecule.

____ [1]

(ii) Combine the reduction and oxidation half-equations to give the overall ionic equation.

____ [2]

(d) At room temperature and pressure, iodine exists as a grey-black shiny solid. Describe the bonding in, and explain the structure of iodine crystals. Explain the relative solubilities of iodine in water and hexane.

____

____

____

____

____

____

____ [5]

Quality of written communication [2]

Examiner Only	
Marks	Remark

Examiner Only	
Marks	Remark

14 Chlorine is produced by the electrolysis of concentrated sodium chloride solution (brine). It is then used by other industries to produce a variety of useful products.

(a) The reaction between chlorine and cold dilute sodium hydroxide is used in the manufacture of bleach.

(i) Write the equation for this reaction.

____ [2]

(ii) This reaction is described as **disproportionation**. Explain the meaning of this term.

____

____

____ [2]

(b) Chlorine reacts with hydrogen to produce hydrogen chloride.

(i) Write the equation for this reaction.

____ [2]

(ii) Suggest why hydrogen chloride has a much lower boiling point than hydrogen fluoride.

____

____

____ [3]

(iii) Explain why hydrogen chloride is more thermally stable than hydrogen iodide.

____

____

____ [2]

Examiner Only	
Marks	Remark

(c) The presence of chloride ions in brine can be established by adding an aqueous solution of silver nitrate.

(i) What would be observed in this reaction?

____ [2]

(ii) Write an ionic equation, including state symbols, for the reaction.

____ [2]

(iii) State what is observed when an excess of dilute aqueous ammonia is then added.

____ [1]

15 The degree of hydration in samples of hydrated sodium carbonate ($Na_2CO_3.xH_2O$) can be determined by different methods.

Examiner Only	
Marks	Remark

(a) When 10.04 g of a sample was heated to constant mass, 3.97 g of anhydrous sodium carbonate was obtained.

(i) Explain the term **anhydrous**.

____

____ [1]

(ii) Calculate the number of moles of anhydrous sodium carbonate obtained.

____ [1]

(iii) Calculate the mass of water present in the sample.

____ [1]

(iv) Calculate the number of moles of water present.

____ [1]

(v) Calculate the value of *x* in the sample.

____ [1]

(b) The degree of hydration can also be determined by dissolving the sample in water and titrating with a standard solution of hydrochloric acid.

(i) What is a **standard** solution?

____ [1]

(ii) Give the equation for the reaction between sodium carbonate and hydrochloric acid.

____ [2]

(iii) Name a suitable indicator for this titration.

____ [1]

Examiner Only	
Marks	Remark

Examiner Only	
Marks	Remark

(c) 3.57 g of a second sample of sodium carbonate was dissolved in water and the resulting solution was made up to 250 cm^3 in a volumetric flask. A 25.0 cm^3 sample of this solution required 28.5 cm^3 of 0.1 mol dm^{-3} hydrochloric acid to reach the end point.

(i) Give the colour change which would be obtained at the end point, using the indicator given in **(b)(iii)**.

From ______________________ to ______________________ [2]

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

____ [1]

(iii) Calculate the number of moles of sodium carbonate present in 25.0 cm^3 of solution.

____ [1]

(iv) Calculate the number of moles of sodium carbonate present in 250 cm^3 of solution.

____ [1]

(v) Calculate the mass of sodium carbonate present in the second sample.

____ [1]

(vi) Calculate the mass of water present in the second sample.

____ [1]

(vii) Calculate the number of moles of water present in the second sample.

____ [1]

(viii) Calculate the value of *x* in the second sample.

____ [1]

THIS IS THE END OF THE QUESTION PAPER

Published Mark Schemes for GCE AS Chemistry

Summer 2010

Issued: October 2010

NORTHERN IRELAND GENERAL CERTIFICATE OF SECONDARY EDUCATION (GCSE) AND NORTHERN IRELAND GENERAL CERTIFICATE OF EDUCATION (GCE)

MARK SCHEMES (2010)

Foreword

Introduction

Mark Schemes are published to assist teachers and students in their preparation for examinations. Through the mark schemes teachers and students will be able to see what examiners are looking for in response to questions and exactly where the marks have been awarded. The publishing of the mark schemes may help to show that examiners are not concerned about finding out what a student does not know but rather with rewarding students for what they do know.

The Purpose of Mark Schemes

Examination papers are set and revised by teams of examiners and revisers appointed by the Council. The teams of examiners and revisers include experienced teachers who are familiar with the level and standards expected of 16 and 18-year-old students in schools and colleges. The job of the examiners is to set the questions and the mark schemes; and the job of the revisers is to review the questions and mark schemes commenting on a large range of issues about which they must be satisfied before the question papers and mark schemes are finalised.

The questions and the mark schemes are developed in association with each other so that the issues of differentiation and positive achievement can be addressed right from the start. Mark schemes therefore are regarded as a part of an integral process which begins with the setting of questions and ends with the marking of the examination.

The main purpose of the mark scheme is to provide a uniform basis for the marking process so that all the markers are following exactly the same instructions and making the same judgements in so far as this is possible. Before marking begins a standardising meeting is held where all the markers are briefed using the mark scheme and samples of the students' work in the form of scripts. Consideration is also given at this stage to any comments on the operational papers received from teachers and their organisations. During this meeting, and up to and including the end of the marking, there is provision for amendments to be made to the mark scheme. What is published represents this final form of the mark scheme.

It is important to recognise that in some cases there may well be other correct responses which are equally acceptable to those published: the mark scheme can only cover those responses which emerged in the examination. There may also be instances where certain judgements may have to be left to the experience of the examiner, for example, where there is no absolute correct response – all teachers will be familiar with making such judgements.

The Council hopes that the mark schemes will be viewed and used in a constructive way as a further support to the teaching and learning processes.

CONTENTS

New Specification

Rewarding Learning

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
2010

Chemistry

Assessment Unit AS 1

assessing

Basic Concepts in Physical and Inorganic Chemistry

[AC111]

MONDAY 7 JUNE, MORNING

MARK SCHEME

		AVAILABLE MARKS
1 D		
2 B		
3 D		
4 B		
5 B		
6 B		
7 B		
8 B		
9 D		
10 C		
[2] for each correct answer	[20]	20
	Section A	**20**

Section B

AVAILABLE MARKS

11 **(a)** 14 [1]

(b) **(i)** Al (g) → Al^+ (g) + e^-

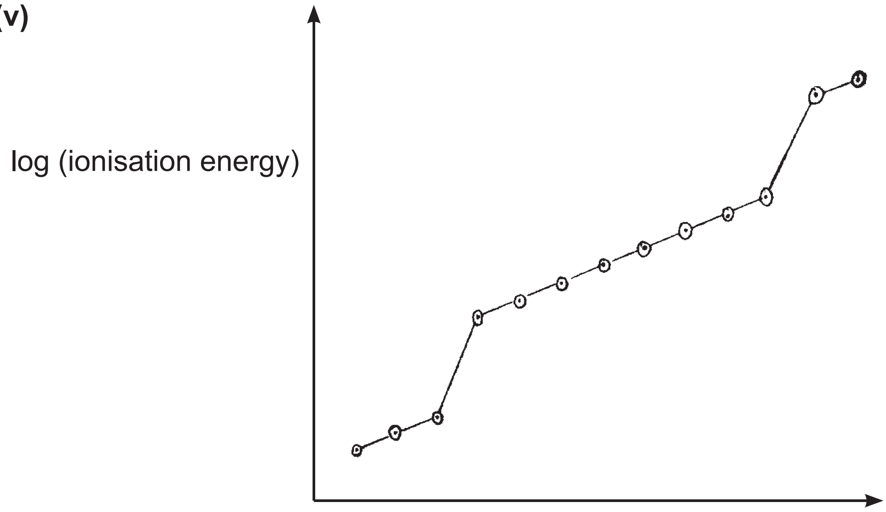
Equation [1] State symbols [1] [2]

(ii) Outer electron for boron is closer to nucleus [1]
and is less shielded [1] (than for aluminium) [2]

(iii) stability of filled 3s shell for magnesium [1]
aluminium has $3p^1$ configuration [1] [2]

(iv) $1s^22s^22p^5$ [1]

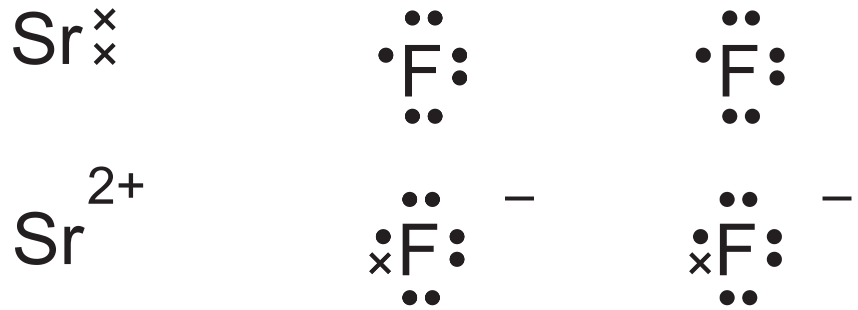
(v)



[2] 10

AVAILABLE MARKS

12 (a) (i)



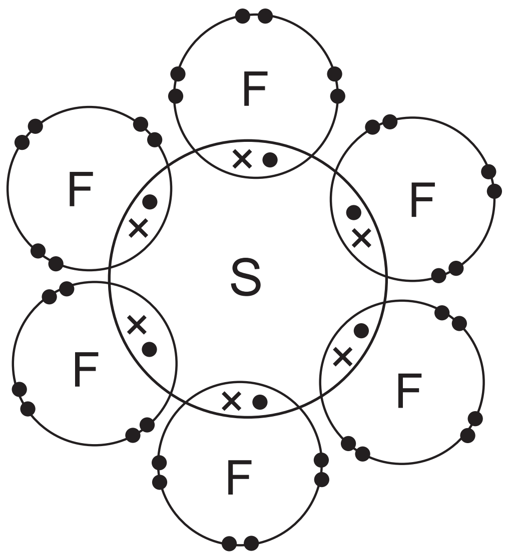
Sr atom [1]
F atom [1]
1 : 2 ratio [1]
correct electron transfer + charges [1]

(ii) solid strontium – good electrical conductor
solid strontium fluoride – poor electrical conductor } [1]
solid strontium – delocalised electrons [1]
solid strontium fluoride – ions are not free to move [1]

(b) (i) The ability/power of an atom to attract bonding electrons in a covalent bond [2]

(ii) $\overset{\delta+}{S} — \overset{\delta-}{F}$ [1]

(iii)

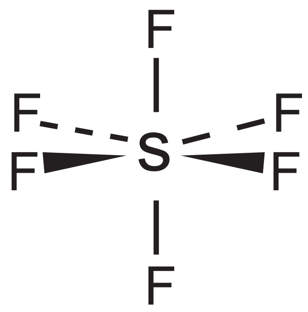


[2]

(iv) does not apply to sulphur (12 electrons in outer shell) [1]
does apply to fluorine (8 electrons in outer shell)/octet rule – has 8 e^- in outer shell [1]

AVAILABLE MARKS

(v)



[1]

Octahedral [1]
(six) bonding pairs [1] repel equally [1] [2]

(vi) dipoles [1] cancel [1] [2]

(c) **(i)** coordinate/dative (covalent) bond [1]

(ii) both shared electrons come from nitrogen/lone pair of electrons on nitrogen shared (donated) [1] 22

13 **(a)** **(i)** loss of electrons [1]

(ii) decrease in oxidation state/number [1]

(b) **(i)** +5 [1]

(ii) +2 [1]

(c) **(i)** $2I^- \rightarrow I_2 + 2e^-$ [1]

(ii) $2HNO_3 + 6H^+ + 6I^- \rightarrow 2NO + 4H_2O + 3I_2$ [2]

(d) reference to **covalent bonding**/**crystalline**
molecular covalent **structure**/**diatomic**/**I_2**
van der Waals' attractions between molecules
1 mark for each two bold points mentioned – Max [3] marks

more soluble in hexane than water [1]
since iodine and hexane are non-polar [1] and water is polar [1]
like dissolves like [1]
Max [4] marks Max [5] marks

Quality of written communication [2] 14

					AVAILABLE MARKS
14	**(a)**	**(i)**	$Cl_2 + 2NaOH \rightarrow NaCl + NaOCl + H_2O$	[2]	
		(ii)	simultaneous reduction and oxidation [1] of the same element/in the same reaction [1]	[2]	
	(b)	**(i)**	$H_2 + Cl_2 \rightarrow 2HCl$	[2]	
		(ii)	HF – hydrogen bonding/polar HCl – polar/less polar than HF Greater energy needed to separate molecules	[1] [1] [1]	
		(iii)	H−Cl bond stronger than H−I bond does not break as easily when heated	[1] [1]	
	(c)	**(i)**	white [1] precipitate [1]	[2]	
		(ii)	$Ag^+ (aq) + Cl^- (aq) \rightarrow AgCl (s)$ Equation [1] State symbols [1]	[2]	
		(iii)	(white) precipitate dissolves	[1]	16

					AVAILABLE MARKS
15	**(a)**	**(i)**	contains no water (of crystallisation)	[1]	
		(ii)	0.0375	[1]	
		(iii)	6.07g	[1]	
		(iv)	0.337	[1]	
		(v)	9	[1]	
	(b)	**(i)**	solution of known concentration	[1]	
		(ii)	$Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$	[2]	
		(iii)	methyl orange	[1]	
	(c)	**(i)**	yellow/orange [1] red/pink [1]	[2]	
		(ii)	0.00285	[1]	
		(iii)	0.001425	[1]	
		(iv)	0.01425	[1]	
		(v)	1.5105	[1]	
		(vi)	2.06	[1]	
		(vii)	0.1144	[1]	
		(viii)	8	[1]	18
				Section B	**80**
				Total	**100**

New Specification

Rewarding Learning

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
2010

Chemistry

Assessment Unit AS 2

assessing

Module 2: Organic, Physical and Inorganic Chemistry

[AC121]

THURSDAY 17 JUNE, AFTERNOON

MARK SCHEME

Section A

			AVAILABLE MARKS
1	D		
2	D		
3	D		
4	B		
5	B		
6	C		
7	B		
8	D		
9	C		
10	A		
[2] for each correct answer		[20]	20
		Section A	**20**

Section B

AVAILABLE MARKS

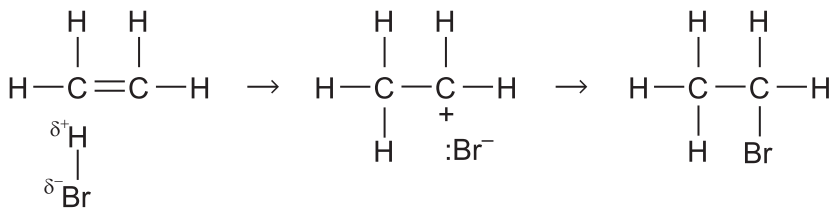
11

Element	Moles	Ratio
C	22.2 ÷ 12 = 1.85	2
H	3.7 ÷ 1 = 3.70	4
Br	74.1 ÷ 80 = 0.93	1

Moles [1]
Empirical Formula = C_2H_4Br [1]
216 ÷ 108 = 2 Molecular Formula = $C_4H_8Br_2$ [1] 3

12 **(a)** **(i)** Electrophilic [1] addition [1] [2]

(ii)



Dipole/H in HBr approaches the C═C [1]
Structure of carbocation intermediate [1]
bromide ion(Br^-) [1]

(b) **(i)** esterification [1]

(ii) ethanoyl chloride [1]

(iii) A = Ethanoic acid [1]
B = Ethyl ethanoate [1] 9
C = Ethanol

13 (a) (i) Either: $C_{17}H_{36} + 17\frac{1}{2}O_2 \rightarrow 17CO + 18H_2O$
Or: $2C_{17}H_{36} + 35O_2 \rightarrow 34CO + 36H_2O$ [2]

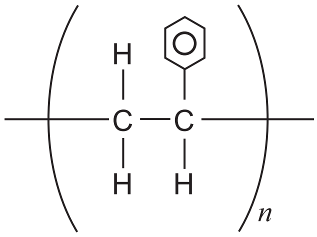
AVAILABLE MARKS

(ii) catalytic converter [1]

(b) Either: $C_{17}H_{36} \rightarrow 2C_2H_4 + C_3H_6 + C_{10}H_{22}$
Or: $C_{17}H_{36} \rightarrow 4C_2H_4 + 2C_3H_6 + C_3H_8$ [2]

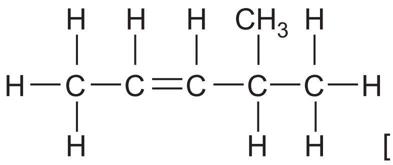
(c) (i) addition [1] polymerisation [1] [2]

(ii)



[1]

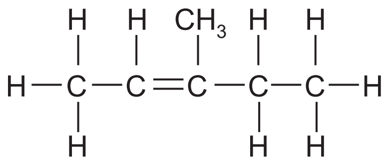
(d) (i) Either:



[1]

4-methylpent-2-ene [1]

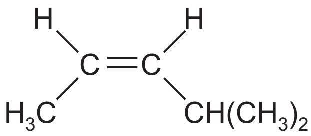
Or:

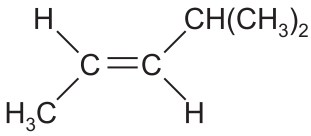


[1]

3-methylpent-2-ene [1] [2]

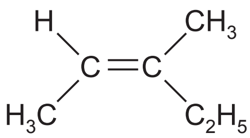
(ii) Either:

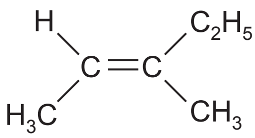




Distinction between *cis* and *trans* not necessary

Or:





Distinction between *E* and *Z* not necessary [2] 12

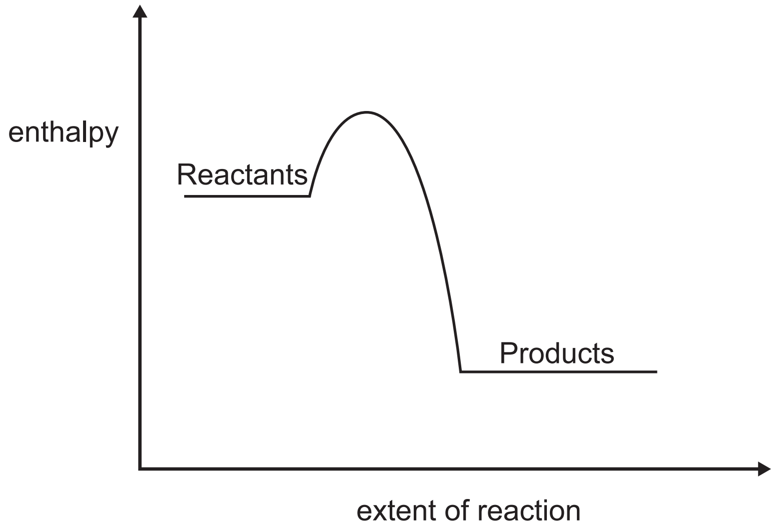
AVAILABLE MARKS

14 (a) (i) Enthalpy change(s) [1]
is/are independent of the route taken [1]

(ii) Enthalpy change which occurs when 1 mole of a substance undergoes complete combustion in oxygen under standard conditions. [2]

(iii) $(-394) + (-286 \times 2) + (+75)$ [1]
$-891\ \text{kJ mol}^{-1}$ [1]

(b) (i)



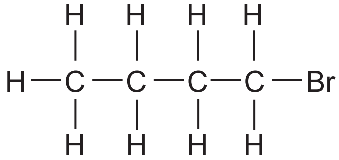
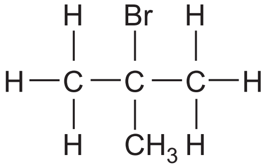
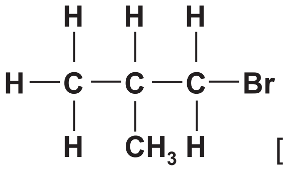
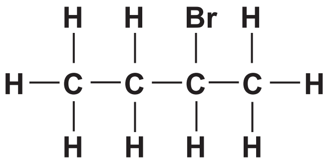
[1]

(ii) Energy required to break one mole of a given bond averaged over many compounds. Error [–1] [2]

(iii) $-698 = 4(\text{C—H}) - 2346$
$4(\text{C—H}) = 1648$
$412\ \text{kJ mol}^{-1}$ [3] 12

				AVAILABLE MARKS
15	**(a)**	Rate of forward reaction = Rate of reverse reaction The amount of any given reactant or product remains constant	[1] [1]	
	(b)	Increase 3 moles (g) LHS 2 moles (g) RHS Equilibrium shifts to RHS to oppose the change	[1] [1]	
	(c)	Decrease (Forward) reaction is exothermic Equilibrium shifts to LHS to oppose the change/cool the system	[1] [1]	
	(d)	No effect on yield Increases the rate of the forward and reverse reactions equally	[1] [1]	
	(e)	**(i)** 400–500 °C [1] 200–300 atm [1]	[2]	
		(ii) Compromise between rate and yield [1]	[1]	11

16 (a)

Structure and Name	Classification
 1-bromobutane	P
 2-bromo-2-methylpropane	**T** [1]
 [1] 1-bromo-2-methylpropane	**P** [1]
 [1] **2-bromobutane** [1]	**S** [1]

[6]

AVAILABLE MARKS

AVAILABLE MARKS

(b) (i) Distil and collect the 1-bromobutane at its boiling point [1]
Shake crude product with a solution of sodium carbonate in a separating funnel [1]
Release pressure [1]
Separate the organic layer [1]
Shake with a drying agent, e.g. anhydrous calcium chloride [1]
Filter [1]
(max = [4])

Quality of written communication

2 marks The candidate expresses ideas clearly and fluently through well-linked sentences and paragraphs. Arguments are generally relevant and well-structured. There are few errors of grammar, punctuation and spelling.

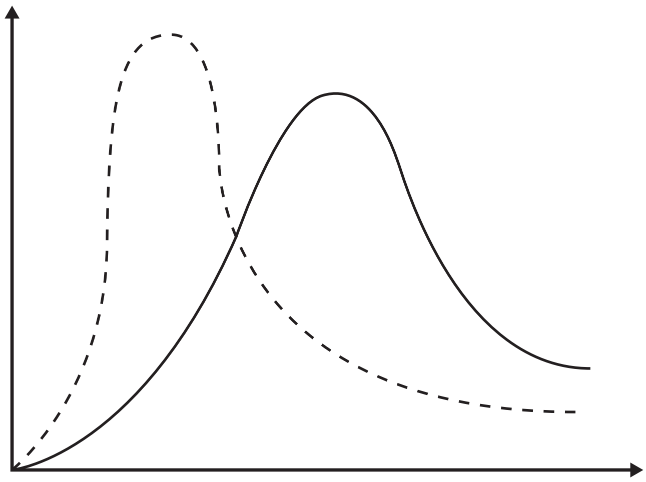
1 mark The candidate expresses ideas clearly, if not always fluently Arguments may sometimes stray from the point. There may be some errors of grammar, punctuation and spelling, but not such as to suggest a weakness in these areas.

0 marks The candidate expresses ideas satisfactorily, but without precision. Arguments may be of doubtful relevance or obscurely presented. Errors in grammar, punctuation and spelling are sufficiently intrusive to disrupt the understanding of the passage.

Quality of written communication [2]

(ii) $11.10 \div 74 = 0.15$ moles of butan-1-ol [1]
$12.33 \div 137 = 0.09$ moles of 1-bromobutane [1]
$(0.09 \div 0.15) \times 100 = 60\%$ [1]

(c) Add an aqueous solution of silver nitrate [1]
A cream precipitate forms [1] 17

17	(a)	(i)	y axis = number of molecules x axis = (kinetic) energy	[1] [1]	
		(ii)	Either: all molecules have some energy/are moving Or: no molecules have zero (kinetic) energy	[1]	
		(iii)	 maximum to LHS of original and higher lower number of molecules at higher energies	[1] [1]	
		(iv)	Number of molecules with energy greater than the activation energy decreases as temperature drops	[2]	
	(b)	(i)	vanadium(V) oxide	[1]	
		(ii)	nickel	[1]	
		(iii)	iron	[1]	
		(iv)	phosphoric acid	[1]	
	(c)	(i)	provides an alternative pathway with lower activation energy	[1] [1]	
		(ii)	moles of hydrogen peroxide = $2 \times 0.05 = 0.1$ mole 2:1 ratio so moles of oxygen = 0.05 Volume (dm^3) = $0.05 \times 24 = 1.2$ dm^3 (1200 cm^3)	[1] [1] [1]	16

AVAILABLE MARKS

Section B	**80**
Total	**100**

New Specification

Rewarding Learning

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
2010

Chemistry

Assessment Unit AS 3

assessing

Module 3: Practical Examination 1

[AC131]

FRIDAY 14 MAY, MORNING

MARK SCHEME

CHEMISTRY

Assessment Unit AS 3

assessing

Module 3: Practical Examination 1

AVAILABLE MARKS

Annotation

1. Please do all marking in red ink.
2. All scripts are checked for mathematical errors. Please adopt the system of one tick (✓) equals [1] mark e.g. if you have awarded 4 marks for part of a question then 4 ticks (✓) should be on this candidate's answer.
3. As candidates have access to scripts please do not write any inappropriate comments on their scripts.

Mark Scheme

Section A

1 **(a)** Rinse the pipette with oven cleaner [1]
Pipette $25cm^3$ of the oven cleaner into the $500cm^3$ volumetric flask and make up to mark using distilled water [1]
(Stopper and) shake the flask/invert [1]
Using a pipette filler [1]
Rinse the pipette with the diluted solution and transfer $25cm^3$ conical flask [1]
(1 mark each to a maximum of four) [4]

(b) Table [1]
Significant figures [2]
Calculation of the average titre [2]
Titration consistency [3]
Agreement with supervisor's titre [4] [12]

NOTES

Table:
Table should include initial burette reading, final burette reading, and volume delivered. The average titre should be calculated and the units included. Units missing (-1)

Significant figures:
All burette readings should be to at least one decimal place – each mistake is penalised by 1 mark.
(However initial burette readings of 0 are penalised once only.)
If used, the second decimal place position should be 0 or 5 only – other values will be penalised by 1 mark.

Average titre
Accurate titrations only should be used.
The use of a rough value is (-1).
The average value can be two decimal places, e.g. 25.37
An incorrect calculation is 0.
Mark denied if:
(i) only one accurate titration done
(ii) if titre not calculated correctly

Titration consistency

This is the difference between the first and second accurate readings

Difference	Mark
±0.1	[3]
±0.2	[2]
±0.3	[1]
±0.4	[4]

Titration agreement with the supervisor

Difference	Mark
±0.1	[4]
±0.2	[3]
±0.3	[2]
±0.4	[1]
±0.5	[0]

Please note that the supervisor's titre should be recorded at the bottom of page 3 in the candidate's script in RED INK.
The marks for table, significant figures etc should be recorded on the left-hand side of the candidate's table of results.

AVAILABLE MARKS

(c) yellow to orange/red [1]

(d) $NaOH + HCl \longrightarrow NaCl + H_2O$ [1]

(e) **(i)** Correct calculation using average titre i.e. $\frac{\text{titre} \times 0.1}{1000}$ [1]

(ii) Use of 1:1 ratio [1]

(iii) Moles of sodium hydroxide in 500cm^3 of diluted oven cleaner (×20) [1]

(iv) Moles of sodium hydroxide in 25cm^3 of undiluted oven cleaner (same value as **(iii)**) [1]

(v) Concentration of sodium hydroxide in oven cleaner (mol dm^{-3}) (×40) [1]

(vi) Concentration in g dm^{-3} (× 40) [1]

(vii) (gdm^{-3}/1000) (× 100) [1]

In part **(e)**, carry error through (c.e.t.) if appropriate. 25

2 Observation and deduction

Safety glasses should be worn at all times and care should be taken during this practical examination.

(a) You are provided with a mixture of two salts, labelled A, which have a common cation.
Carry out the following experiments on the mixture.
Record your observations and deductions in the spaces below and identify the two salts.

Please note that the total marks for each box should be written in the box and not in the examiner's column.

Experiment	Observations	Deductions
1) Make a solution of A by dissolving a spatula of A in a test tube half-full of water. Warm gently. Transfer 1cm^3 of the solution into each of two separate test tubes. (a) Add a few drops of sodium hydroxide solution to the first test tube. Then add a further 3cm^3 of the sodium hydroxide solution to the test tube. (b) Add a few drops of ammonia solution to the second test tube. Then add a further 3cm^3 of the ammonia solution to the test tube.	*Colourless solution [1]* *White precipitate [1]* *Insoluble in excess [1]* *White precipitate [1]* *Insoluble in excess [1]* [5]	*Possibly magnesium, Zinc or Aluminium ion [1]* *Possibly magnesium / ion or compounds [1]* *Confirms the presence of magnesium ion or compounds [1]* [3]
2) Make a solution of A by dissolving half a spatula – measure of A in a test tube half – full of nitric acid solution. Warm gently. Transfer 1cm^3 of the solution into each of two separate test tubes. (a) (i) Add a few drops of silver nitrate solution to the first test tube. (ii) Then add about 2cm^3 of dilute ammonia solution to the same test tube. (b) Add a few drops of barium chloride solution to the second test tube.	*No effervescence [1]* *White precipitate [1]* *Precipitate dissolves [1]* *White precipitate [1]* [4]	*Not a carbonate/hydrogen carbonate [1]* *Possibly chloride ion [1]* *Chloride ion present [1]* *Sulphate ion present [1]* [4]

Name the two salts present in A:

Magnesium chloride [1]

Magnesium sulphate [1]

An incorrect deduction can be carried through to naming the salts.
A deduction based on an incorrect observation can be credited on the basis of carry error through (c.e.t.).
In **2 (a) (i)** and **(ii)** mark denied if 'chlorine' is given.

(b) You are provided with an organic liquid labelled B. Carry out the following experiments on the liquid. Record your observations and deductions in the spaces below.

Experiment	Observations	Deductions
1) Place 10 drops of B in a test tube and add 1cm^3 of water.	*Forms one layer/mixes completely/dissolves/ colourless solution [1]* [1]	*Soluble/miscible with water [1]* *Forms H-bonds with water [1]* [2]
2) Place 10 drops of B on a watch glass placed on a heat proof mat and ignite it using a splint.	*Yellow/blue flame [1]* *Clean flame/little or no smoke [1]* [2]	*Low carbon content/saturated compound/C:H low/short chain/contains oxygen [1]* [1]
3) Add approximately 10 drops of B to a test tube one quarter full of bromine water and mix well.	*Yellow/orange colour remains [1]* [1]	*Alkene present/Saturated/ no C=C bonds present [1]* [1]
4) Add ten drops of B to 2cm^3 of acidified potassium dichromate solution in a test tube. Warm the mixture gently.	*Orange to green [1]* *Change of smell [1]* [2]	*Primary or secondary alcohol / Not a tertiary alcohol [1]* *Can be oxidised/reducing agent / aldehyde or ketone formed [1]* [2]

AVAILABLE MARKS

Based on the tests above, suggest:

a functional group which may be present in B.

-OH [1]

a functional group which is absent in B.

C=C [1]

Mark to a maximum of 29 marks for question 2 29

N.B. Liquid is ethanol

Section A Total **54**

Section B

AVAILABLE MARKS

3 **(a)** **(i)** Mass of the container
Mass of the container + lead oxide [2]

(ii) Carry out in a fume cupboard/use gloves/tongs/allow to cool [1]
Carbon monoxide is poisonous/apparatus is hot [1] [2]

(iii) Heat and reweigh [1] until the mass is constant [1] [2]

(iv) Prevents reoxidising the lead/stops air re-entering [1]
Allows the metal to cool/apparatus still hot [1] [2]

(b) **(i)** 1.39-1.26 = 0.13g [1]

(ii) $0.13/16 = 8.125 \times 10^{-3}$ [1]

(iii) $1.26/207 = 6.087 \times 10^{-3}$ [1]

(iv) $Pb:O = 6.087 \times 10^{-3} : 8.125 \times 10^{-3}$
Pb:O = 3:4
Pb_3O_4
(-1 for each mistake) [2]

(c) **(i)** $2PbS + 3O_2 \rightarrow 2PbO + 2SO_2$ [2]

(ii) $2PbO + C \rightarrow 2Pb + CO_2$ [1]

(iii) $20.32/100 \times 8.5 = 1.73g$
$207/239 \times 100 = 86.61\%$
$1.73/100 \times 86.61 = 1.50g$
$1.500/20.32 \times 100 = 7.38\%$ [4] 20

In parts **(b)** and **(c)**, carry error through if appropriate.

4 **(a)** Repeated boiling and condensing of a reaction mixture without loss of reactants [1]

(b) **(i)** Place sample in a separating funnel
Add sodium hydrogencarbonate / carbonate solution and shake [2]

(ii) Add suitable named anhydrous solid to the sample [1]
Filter/decant [1] [2]

(iii) Distil the sample
Collect the sample boiling between 210 - 212°C [2] 7

			AVAILABLE MARKS
5 (a)	Dip nichrome wire/silica rod into concentrated hydrochloric acid and then into the solid [1] Heat in a blue Bunsen flame [1] Orange/yellow flame indicates the presence of sodium [1]	[3]	
(b)	Add dilute hydrochloric acid to the solid [1] Bubble the gas given off through limewater [1] If the limewater turns milky/cloudy then carbonate/hydrogencarbonate ions are present [1]	[3]	
(c)	Make a solution of the solid [1] Add a solution of magnesium ions [1] A white precipitate indicates a carbonate, no precipitate indicates a hydrogencarbonate [1]	[3]	9
		Section B Total	**36**
		Paper Total	**90**

New Specification

Rewarding Learning

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
2010

Chemistry

Assessment Unit AS 3

assessing

Module 3: Practical Examination 2

[AC132]

TUESDAY 18 MAY, AFTERNOON

MARK SCHEME

CHEMISTRY

Assessment Unit AS 3

assessing

Module 3: Practical Examination 2

Mark Scheme

AVAILABLE MARKS

Annotation

1. Please do all marking in red ink.
2. All scripts are checked for mathematical errors. Please adopt the system of one tick (✓) equals [1] mark e.g. if you have awarded 4 marks for part of a question then 4 ticks (✓) should be on this candidate's answer.
3. As candidates have access to scripts please do not write any inappropriate comments on their scripts.

Section A

1 **(a)** Rinse the pipette with distilled water and then drain cleaner [1]
Pipette $25cm^3$ of the oven cleaner into the $500cm^3$ volumetric flask and make up to mark using distilled water [1]
(Stopper and) shake the flask/invert [1]
Using a pipette filler [1]
Rinse the pipette with the diluted solution and transfer $25cm^3$ conical flask [1]
(1 mark each to a maximum of four) [4]

(b) Table [1]
Significant figures [2]
Calculation of the average titre [2]
Titration consistency [3]
Agreement with supervisor's titre [4] [12]

NOTES

Table:
Table should include initial burette reading, final burette reading, and volume delivered. The average titre should be calculated and the units included. Units missing (-1).

Significant figures:
All burette readings should be to at least one decimal place – each mistake is penalised by 1 mark.
(However initial burette readings of 0 are penalised once only.)
If used, the second decimal place position should be 0 or 5 only – other values will be penalised by 1 mark.

Average titre
Accurate titrations only should be used.
The use of a rough value is (-1).
The average value can be two decimal places, e.g. 25.37
An incorrect calculation is 0.
Mark denied if:
(i) only one accurate titration done
(ii) if titre not calculated correctly

AVAILABLE MARKS

Titration consistency

This is the difference between the first and second accurate readings

Difference	Mark
±0.1	[3]
±0.2	[2]
±0.3	[1]
±0.4	[0]

Titration agreement with the supervisor

Difference	Mark
±0.1	[4]
±0.2	[3]
±0.3	[2]
±0.4	[1]
±0.5	[0]

Please note that the supervisor's titre should be recorded at the bottom of page 3 in the candidate's script in RED INK.
The marks for table, significant figures etc should be recorded on the left-hand side of the candidate's table of results.

(c) yellow to orange/red [1]

(d) $KOH + HCl \longrightarrow KCl + H_2O$ [1]

(e) **(i)** Correct calculation using average titre i.e. $\frac{\text{titre} \times 0.1}{1000}$ [1]

(ii) Use of 1:1 ratio [1]

(iii) Moles of potassium hydroxide 500cm^3 of diluted drain cleaner (×20) [1]

(iv) Moles of potassium hydroxide in 25cm^3 of undiluted drain cleaner (same value as **(iii)**) [1]

(v) Concentration of potassium hydroxide in drain cleaner (mol dm^{-3}) (×40) [1]

(vi) Concentration in mol dm^{-3} (× 56) [1]

(vii) (gdm^{-3}/1000) × 100 [1]

In part **(e)**, carry error through (c.e.t.) if appropriate. 25

2 Observation and deduction

AVAILABLE MARKS

Safety glasses should be worn at all times and care should be taken during this practical examination.

(a) You are provided with a mixture of two salts, labelled X, which have a common cation. Carry out the following experiments on the mixture.
Record your observations and deductions in the spaces below and identify the two salts.

Please note that the total marks for each box should be written in the box and not in the examiner's column.

Experiment	Observations	Deductions
1) Make a solution of X by dissolving half a spatula-measure of X in a test tube half-full of water. Transfer $1cm^3$ of this solution into each of two separate test tubes. (a) Add a few drops of sodium hydroxide solution to the first test tube. Then add a further $10cm^3$ of the sodium hydroxide solution to the test tube. (b) Add a few drops and of ammonia solution to the second test tube. Then add a further $5cm^3$ of the ammonia solution to the test tube.	*Colourless solution [1]* *White precipitate [1]* *soluble in excess [1]* *White precipitate [1]* *soluble in excess [1]* [5]	*Possibly Magnesium, Zinc or Aluminium ion [1]* *Possibly zinc or aluminium ion or compounds [1]* *Confirms the presence of zinc ion or compounds [1]* [3]
2) Make a solution of X by dissolving half a spatula-measure of X in a test tube half-full of nitric acid solution. Transfer $1cm^3$ of this solution into each of two separate test tubes. (a) (i) Add a few drops of silver nitrate solution to the first test tube. (ii) Then add about $5cm^3$ of dilute ammonia solution to the same test tube. (b) Add a few drops of barium chloride solution to the second test tube.	*No effervescence [1]* *White precipitate [1]* *Precipitate dissolves [1]* *White precipitate [1]* [4]	*Not a carbonate/hydrogen carbonate [1]* *Possibly chloride ions [1]* *Chloride ion present [1]* *Sulphate ion present [1]* [4]

Name the two salts present in X:

Zinc chloride [1]

Zinc sulphate [1]

An incorrect deduction can be carried through to naming the salts.
A deduction based on an incorrect observation can be credited on the basis of carry error through (c.e.t.).
In **2 (a) (i)** and **(ii)** mark denied if 'chlorine' is given.

(b) You are provided with an organic liquid labelled Y. Carry out the following experiments on the liquid. Record your observations and deductions in the spaces below.

Experiment	Observations	Deductions
1) Place 10 drops of Y in a test tube and add 1cm^3 of water.	*Immiscible/2 layers [1]* [1]	*No H-bonds/No -OH present/non-polar [1]* *Less dense than water [1]* [2]
2) Place 10 drops of Y on a watch glass placed on a heat proof mat and ignite it using a splint.	*Smoky [1]* *Yellow flame [1]* [2]	*Unsaturated/high carbon content [1]* [1]
3) Add approximately 10 drops of Y to a test tube one quarter full of bromine water and mix well.	*Yellow/orange/brown [1]* *Turns colourless [1]* [2]	*Alkene present/ Unsaturated/C=C bonds present [1]* [1]
4) Add 10 drops of Y to 2cm^3 of acidified potassium dichromate solution in a test tube. Warm the mixture gently.	*Remains orange /no change [1]* [1]	*Possibly a tertiary alcohol/ Not a primary or secondary alcohol [1]* *Not reducing agent/not oxidation [1]* [2]

Based on the tests above, suggest:

AVAILABLE MARKS

a functional group which may be present in Y.

C=C [1]

a functional group which is absent in Y.

-OH [1]

Mark to a maximum of 29 marks for question 2 29

N.B. The Liquid is turpentine

Section A Total **54**

Section B

AVAILABLE MARKS

3 **(a)** **(i)** Mass of the container
Mass of the container + copper oxide [2]

(ii) Flush out the apparatus with hydrogen/use gloves/tongs/allow to cool [1]
Hydrogen – air mixture is explosive/apparatus is hot [1] [2]

(iii) Heat and reweigh [1] until the mass is constant [1] [2]

(iv) Prevents reoxidising the copper/stops air re-entering [1]
Allows the metal to cool/apparatus still hot [1] [2]

(b) **(i)** 2.16 – 1.92 = 0.24g [1]

(ii) 0.24/16 = 0.015 [1]

(iii) 1.92/64 = 0.030 [1]

(iv) Cu:O = 0.030:0.015
Cu:O = 2:l
Cu_2O
(-1 for each mistake) [2]

(c) **(i)** $2Cu_2S + 3O_2 \rightarrow 2Cu_2O + 2SO_2$ [2]

(ii) $2Cu_2O + Cu_2S \rightarrow 6Cu + SO_2$ [1]

(iii) 25.86/100 × 34.8 = 8.999g
64/184 × 100 = 34.78%
8.999/100 × 34.78 = 3.130g
3.130/25.86 × 100 = 12.1% [4] 20

In parts **(b)** and **(c)**, carry error through if appropriate.

4 **(a)** Repeated boiling and condensing of a reaction mixture without loss of reactants [1]

(b) **(i)** Place sample in a separating funnel
Add sodium hydrogencarbonate / carbonate solution and shake [2]

(ii) Add suitable named anhydrous solid to the sample
Filter/decant [1] [2]

(iii) Distil the sample
Collect the sample boiling between 90 – 92°C [2] 7

		AVAILABLE MARKS
5 **(a)** Dip nichrome wire/silica rod into concentrated hydrochloric acid and then into the solid [1] Heat in a blue Bunsen flame [1] lilac flame indicates the presence of potassium [1]	[3]	
(b) Add dilute nitric acid to solid followed by silver nitrate solution [1] White precipitate shows chloride [1] Yellow precipitate shows iodide [1]	[3]	
(c) Silver chloride is soluble in dilute and concentrated ammonia solution [1] forming a colourless solution [1] Silver iodide is insoluble in dilute and concentrated ammonia solution [1]	[3]	9
	Section B Total	**36**
	Paper Total	**90**

Rewarding Learning

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
January 2011

Centre Number

71	

Candidate Number

Chemistry

Assessment Unit AS 1

assessing

Basic Concepts in Physical and Inorganic Chemistry

[AC111]

AC111

THURSDAY 13 JANUARY, MORNING

TIME

1 hour 30 minutes.

INSTRUCTIONS TO CANDIDATES

Write your Centre Number and Candidate Number in the spaces provided at the top of this page.
Answer **all sixteen** questions.
Answer **all ten** questions in **Section A**. Record your answers by marking the appropriate letter on the answer sheet provided. Use only the spaces numbered 1 to 10. Keep in sequence when answering.
Answer **all six** questions in **Section B**. Write your answers in the spaces provided in this question paper.

INFORMATION FOR CANDIDATES

The total mark for this paper is 100.
Quality of written communication will be assessed in question **14(d)**.
In Section A all questions carry equal marks, i.e. **two** marks for each question.
In Section B the figures in brackets printed down the right-hand side of pages indicate the marks awarded to each question or part question.
A Periodic Table of Elements (including some data) is provided.

For Examiner's use only	
Question Number	**Marks**
Section A	
1–10	
Section B	
11	
12	
13	
14	
15	
16	
Total Marks	

Section A

For each of the following questions only **one** of the lettered responses (A–D) is correct.

Select the correct response in each case and mark its code letter by connecting the dots as illustrated on the answer sheet.

1 An atom in which the number of protons is greater than the number of neutrons is

A ^{2}H.
B ^{3}He.
C ^{10}B.
D ^{39}K.

2 Which one of the following is a correct description of electronic transitions in a given series in the atomic emission spectrum of hydrogen?

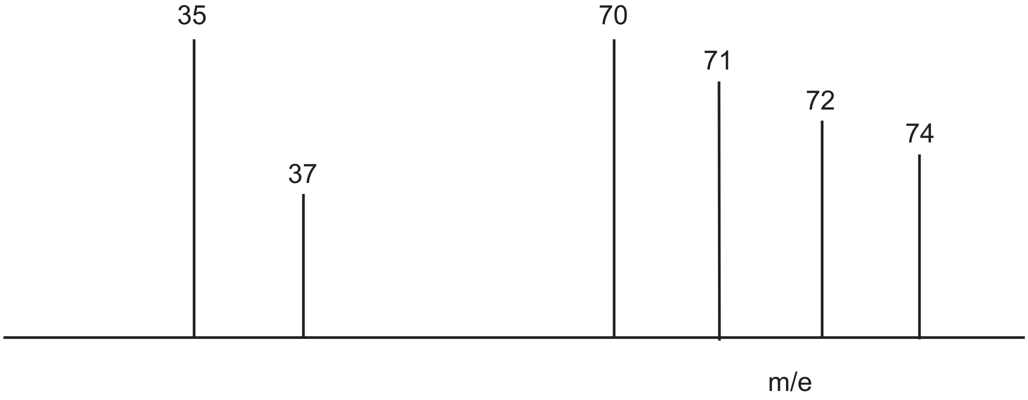
A They all start from the ground state.
B They all end at the ground state.
C They all start from one particular energy level.
D They all end at one particular energy level.

3 Which one of the following lists the first ionisation energies (in kJ mol^{-1}) of the elements magnesium, aluminium, silicon, phosphorus and sulfur in this order?

A	496	736	577	786	1060
B	577	786	1060	1000	1260
C	736	577	786	1060	1000
D	786	1060	1000	1260	1520

4 The mass spectrum of molecular chlorine, Cl_2, is shown below. An additional peak is in the spectrum which should not be present.

Which one of the following peaks should **not** be present?



A 35
B 71
C 72
D 74

5 A solid melts sharply at 100–101 ºC. It does not conduct electricity even when molten. It dissolves in hydrocarbon solvents. The solid has

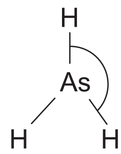
A an atomic structure.
B a giant covalent structure.
C an ionic structure.
D a molecular covalent structure.

6 Which one of the following gaseous hydrides most readily decomposes into its elements on contact with a hot glass rod?

A ammonia
B hydrogen fluoride
C hydrogen iodide
D steam

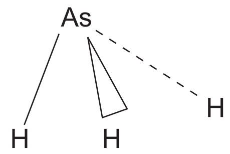
7 Arsine, AsH_3, is a molecular hydride of arsenic which is found in Group V of the Periodic Table. Which one of the following is the structure of arsine in the vapour state?

A



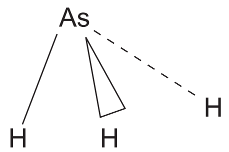
Bond angles 120º

B



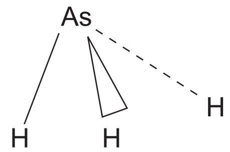
Bond angles greater than 109.5º

C



Bond angles less than 109.5º

D



Bond angles 109.5º

8 50 cm^3 of 0.20 mol dm^{-3} sulphuric acid is exactly neutralised by

A 100 cm^3 of 0.40 mol dm^{-3} potassium hydroxide solution.
B 25 cm^3 of 0.20 mol dm^{-3} potassium hydroxide solution.
C 50 cm^3 of 0.20 mol dm^{-3} potassium hydroxide solution.
D 100 cm^3 of 0.20 mol dm^{-3} potassium hydroxide solution.

9 Which one of the following is the number of electrons which have approximately the same mass as that of a proton?

A 20
B 200
C 2000
D 20000

10 Which one of the following oxides is not polar?

A CO
B CO_2
C H_2O
D NO

Section B

Examiner Only	
Marks	Remark

Answer **all six** questions in this section.

11 Nitrogen dioxide, NO_2, is one of the components of photochemical smog. The energy required to dissociate this molecule into NO molecules and O atoms is 305 kJ mol^{-1}.
Use the following headings to calculate the frequency of radiation required to cause the dissociation.

(a) Convert 305 kJ into Joules

____ [1]

(b) Calculate the number of Joules required to dissociate one molecule of nitrogen dioxide.

____ [1]

(c) Use the equation E = hf to convert the value in Joules into a frequency and state the units.

____

____ [1]

12 The mineral beryl, $Be_3Al_2Si_6O_{18}$, is the principal source of beryllium. Although there are minerals richer in beryllium they are scarce and costly.

(a) Calculate the percentage, by mass, of beryllium in beryl.

____

____

____ [3]

(b) The metal beryllium is obtained either by the electrolysis of a fused mixture of beryllium and potassium chlorides at 350 °C or by the reduction of beryllium fluoride with magnesium.

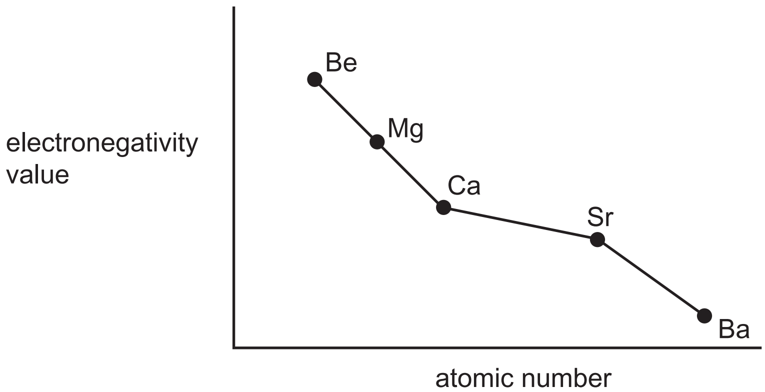
(i) Write an equation for the formation of beryllium from beryllium ions.

____ [1]

(ii) Write an equation for the formation of beryllium from the reduction of beryllium fluoride with magnesium.

____ [1]

(c) The first element in a Group often has more distinctive properties than the elements in the rest of the Group. This is often as a result of the difference in electronegativities. The electronegativity values of the Group II elements are shown below.



(i) Explain the meaning of the term **electronegativity**.

____

____ [2]

Examiner Only	
Marks	Remark

(ii) Using electronegativity suggest why beryllium chloride is a covalent molecule and barium chloride is ionic.

____

____ [2]

(iii) State **two** physical properties which could be used to distinguish these two chlorides.

____

____ [2]

(d) Beryllium chloride may be prepared by the action of chlorine or hydrogen chloride on the metal.

(i) Write the equation for the reaction of beryllium with hydrogen chloride.

____ [1]

(ii) Draw a dot and cross diagram to show the formation of beryllium chloride from beryllium and chlorine atoms. Use only the outer electrons of each atom.

[3]

(iii) State the octet rule.

____

____ [2]

(iv) Beryllium chloride can be said to obey the octet rule and also not to obey the octet rule. Explain this contradiction.

____

____ [2]

Examiner Only	
Marks	Remark

(v) Draw the shape of a beryllium chloride molecule.

Examiner Only	
Marks	Remark

[1]

(vi) State the shape of the beryllium chloride molecule.

____ [1]

(vii) Explain the shape of the beryllium chloride molecule.

____

____ [2]

Examiner Only	
Marks	Remark

13 Chlorine forms a series of oxides some of which are listed below.

chlorine monoxide	Cl_2O
chlorine dioxide	ClO_2
chlorine hexoxide	Cl_2O_6
chlorine heptoxide	Cl_2O_7

(a) Deduce the systematic name for chlorine heptoxide using the oxidation number of chlorine.

____ [1]

(b) Chlorine dioxide dissolves in water to form a solution which eventually forms a mixture of chloric and hydrochloric acids.

$$6ClO_2 + 3H_2O \rightarrow 5HClO_3 + HCl$$

The chlorine atoms in chlorine dioxide undergo disproportionation in this reaction.

(i) Explain the meaning of the term **disproportionation**.

____

____ [1]

(ii) Calculate the oxidation number of chlorine in the reactant and in the products of this reaction and use them to confirm that the reaction is a disproportionation reaction.

____

____

____ [3]

(c) Chlorine gas dissolves in water to the extent of 0.8 g in 100 cm^3 at atmospheric pressure and 20 °C.

(i) Calculate the molarity of the chlorine water, Cl_2 (aq), produced.

____

____

____ [2]

(ii) Name another solvent in which chlorine will readily dissolve.

____ [1]

Examiner Only	
Marks	Remark

14 An experiment was set up to investigate the displacement reactions of the halogens.

Solutions of sodium halides were prepared and reacted with other halogens. The results table is shown below.

	sodium iodide (aq)	**sodium bromide (aq)**	**sodium chloride (aq)**
iodine solution		**X**	**X**
bromine solution	✓		
chlorine solution			

✓ means that a reaction took place
X means that no reaction took place

(a) Complete the **three** remaining places in the table. [2]

(b) (i) Both bromine and iodine solutions are coloured. Describe the observations which would indicate that a reaction took place when aqueous sodium iodide is added to a bromine solution.

____

____ [2]

(ii) Write the ionic equation for the reaction between bromine solution and aqueous sodium iodide.

____ [1]

(c) (i) Describe what is observed when chlorine solution is added to aqueous sodium bromide.

____

____ [2]

(ii) Write the equation for the reaction between chlorine solution and aqueous sodium bromide.

____ [1]

Examiner Only	
Marks	Remark

(d) If you had poured solutions of sodium iodide, bromide and chloride into beakers A, B and C and forgotten to label them, describe how, using aqueous silver nitrate and both dilute and concentrated ammonia solutions, you would determine which sodium salt was in which beaker. Each beaker must be tested.

____

____

____

____

____

____

____

____

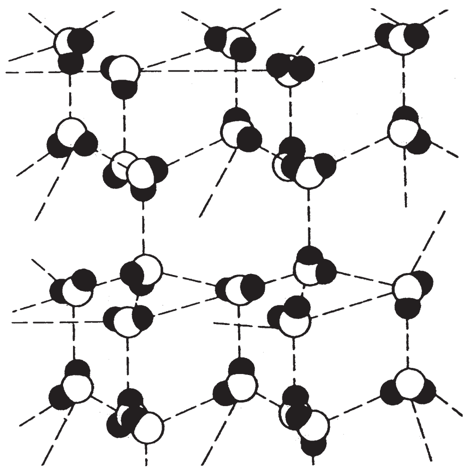
____

____ [6]

Quality of written communication [2]

Examiner Only	
Marks	Remark

15 The structure of ice is shown below. The water molecules are held together by hydrogen bonds which are a type of intermolecular force.



(a) Name **two** other types of intermolecular force.

____ [2]

(b) **(i)** Explain how hydrogen bonding takes place between the water molecules in ice.

____

____ [2]

(ii) Explain, using the structure above, why ice is less dense than water.

____

____

____ [2]

(c) Although water is capable of forming hydrogen bonds it does not form long chains of "polywater" at room temperature. However, in the liquid state, molecules such as hydrogen fluoride do form very short chains. Suggest why water does not form chains and liquid hydrogen fluoride does.

____

____

____ [2]

Examiner Only

Marks	Remark

(d) Ammonia is another substance that can form hydrogen bonds. However, ammonia has a pyramidal structure.

(i) Draw two molecules of ammonia and show the hydrogen bond between the two molecules.

[2]

(ii) Explain why when ammonia reacts with a hydrogen ion it loses the ability to form hydrogen bonds.

____

____ [1]

(e) Explain why ammonia is extremely soluble in water.

____

____ [2]

Examiner Only	
Marks	Remark

16 Lithium exists in nature as two isotopes, ^{6}Li and ^{7}Li. The composition of a sample of lithium in nature is shown in the table below.

isotope	**% abundance**
lithium 6	7.42
lithium 7	92.58

(a) Draw the structure of a ^{7}Li atom, labelling all the sub-atomic particles.

[3]

(b) State and explain to which of the s, p or d blocks lithium belongs.

____

____ [2]

(c) Calculate the relative atomic mass of lithium to two decimal places.

____

____

____ [3]

Examiner Only

Marks	Remark

(d) Lithium sulphate is readily soluble in water and crystallises from solution as the hydrate.

(i) Explain what is meant by the term **water of crystallisation**.

____

____ [1]

(ii) Write the formula of anhydrous lithium sulphate.

____ [1]

(iii) Calculate the formula of hydrated lithium sulphate if 3.76 g of the hydrated lithium salt produces 3.23 g of anhydrous lithium sulphate on heating.

____

____

____

____

____ [3]

(e) Lithium sulphate can be used in a flame test. Explain how a flame test could be carried out and state the expected colour of the flame.

____

____

____

____

____ [4]

THIS IS THE END OF THE QUESTION PAPER

Examiner Only	
Marks	Remark

Rewarding Learning

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
January 2011

Chemistry

Assessment Unit AS 1

assessing

Basic Concepts in Physical and Inorganic Chemistry

[AC111]

THURSDAY 13 JANUARY, MORNING

MARK SCHEME

Section A

			AVAILABLE MARKS
1	B		
2	D		
3	C		
4	B		
5	D		
6	C		
7	C		
8	D		
9	C		
10	B		
[2] for each correct answer		[20]	20
		Section A	**20**

Section B

AVAILABLE MARKS

11 **(a)** $305 \times 10^3 = 3.05 \times 10^5$ J [1]

(b) $\frac{3.05 \times 10^5}{6.02 \times 10^{23}} = 5.07 \times 10^{-19}$ J [1]

(c) $0.507 \times 10^{-18} = 6.63 \times 10^{-34} \times f$
$f = 0.076 \times 10^{16}$
$= 7.6 \times 10^{14}$ s^{-1}/Hz [1] 3

12 **(a)** $Be_3Al_2Si_6O_{18} = 3 \times 9 + 2 \times 27 + 6 \times 28 + 18 \times 16$
$= 27 + 54 + 168 + 288$
$= 537$
% Be $= \frac{27}{537} \times 100 = 5.03\%$ [3]

(b) **(i)** $Be^{2+} + 2e^- \rightarrow Be$ [1]

(ii) $BeF_2 + Mg \rightarrow Be + MgF_2$ [1]

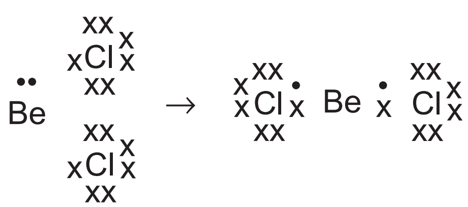
(c) **(i)** ability of an atom in a covalent bond to attract (bonding) electrons [2]

(ii) Be and Cl have similar EN values [1]
Ba and Cl have (very) different EN values [1] [2]

(iii) melting point, boiling point, "reaction" with water, (electrical) conductivity etc.
2 from list [2]

(d) **(i)** $Be + 2HCl \rightarrow BeCl_2 + H_2$ [1]

(ii)



[3]

(iii) 8 electrons around an atom in outer shell [2]

(iv) 8 electrons around Cl [1]
4 electrons around Be [1] [2]

(v) Cl — Be — Cl [1]

(vi) linear/straight [1]

(vii) bond electrons repel [1] to minimise forces [1] [2] 23

		AVAILABLE MARKS
13 **(a)** chlorine(VII) oxide	[1]	
(b) **(i)** atom raises and lowers its oxidation number during a chemical reaction	[1]	
(ii) ClO_2 +4 $HClO_3$ +5 HCl –1 [2] ∵ +4 → +5 oxidation +4 → –1 reduction [1]	[3]	
(c) **(i)** $Cl_2 = 2 \times 35.5 = 71$ $0.8\,g = \frac{0.8}{71} = 0.01127$ mol ∵ 0.113 M	[2]	
(ii) e.g. hexane	[1]	8
14 **(a)** ✓ ✓ ✗ (Note that the marking of the colour changes in this question will be subject to the application of the "colour changes" scheme.)	[2]	
(b) **(i)** compare colours with original solutions it should go darker colour of iodine is darker than bromine	[2]	
(ii) $Br_2 + 2I^- \rightarrow 2Br^- + I_2$ no state symbols required	[1]	
(c) **(i)** colourless solution [1] orange/yellow/brown colour produced [1]	[2]	
(ii) $Cl_2 + 2NaBr \rightarrow 2NaCl + Br_2$ no state symbols required	[1]	
(d) iodide: yellow ppt insoluble in (both dilute and conc.) ammonia solution [2] bromide: cream ppt soluble in conc NH_3 [2] chloride: white ppt soluble in dil NH_3 [2]	[6]	
Quality of written communication	[2]	16

AVAILABLE MARKS

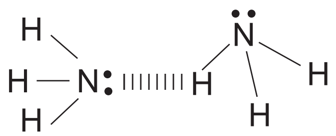
15 **(a)** van der Waals [1]
dipole – dipole [1] [2]

(b) **(i)** attraction between lone pair on O and a H atom on another water molecule [2]

(ii) distance between water molecules in ice greater than in water [1]
open structure [1] longer/expanded H-bonds [1] fixed H-bonds [1]
any 2 [2]

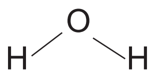
(c) F more electronegative than oxygen/H—F bond more polarised [1]
F IIIIII H bond stronger [1]
movement of water molecules breaks H bonds [1] [2]

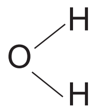
(d) **(i)**



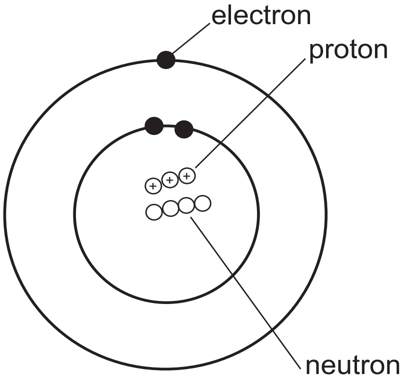
[2]

(ii) lone pair removed/forms a dative bond [1]

(e) N**:** forms hydrogen bonds with  [1]

and N—H forms hydrogen bonds with  [1] [2] 13

16 (a)



[3]

(b) s block [1]
outer electron in s shell [1]

(c) 7.42 3 6 5 44.52

92.58 3 7 5 648.06

Total 5 692.58

4 100 5 6.9258

5 6.93 [3]

(d) **(i)** water chemically bonded in salt [1]

(ii) Li_2SO_4 [1]

(iii) Li_2SO_4
5 2 3 7 1 32 1 64 5 110

moles 5 $\frac{3.23}{110}$ 5 0.029

H_2O
5 2 1 16 5 18

moles 5 $\frac{0.53}{18}$ 5 0.029

{ $Li_2SO_4 \cdot H_2O$ [3]

(e) conc hydrochloric acid
blue Bunsen flame
nichrome (platinum) wire
crimson flame [4]

AVAILABLE MARKS
17

Section B **80**

Total **100**

Rewarding Learning

Centre Number	
71	

Candidate Number

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
2011

Chemistry

Assessment Unit AS 1

assessing

Basic Concepts in Physical and Inorganic Chemistry

[AC112]

AC112

WEDNESDAY 15 JUNE, AFTERNOON

TIME

1 hour 30 minutes.

INSTRUCTIONS TO CANDIDATES

Write your Centre Number and Candidate Number in the spaces provided at the top of this page.
Answer **all sixteen** questions.
Answer **all ten** questions in **Section A**. Record your answers by marking the appropriate letter on the answer sheet provided. Use only the spaces numbered 1 to 10. Keep in sequence when answering.
Answer **all six** questions in **Section B**. Write your answers in the spaces provided in this question paper.

INFORMATION FOR CANDIDATES

The total mark for this paper is 100.
Quality of written communication will be assessed in question **15(f)**.
In Section A all questions carry equal marks, i.e. **two** marks for each question.
In Section B the figures in brackets printed down the right-hand side of pages indicate the marks awarded to each question or part question.
A Periodic Table of Elements (including some data) is provided.

For Examiner's use only	
Question Number	Marks
Section A	
1-10	
Section B	
11	
12	
13	
14	
15	
16	
Total Marks	

Section A

For each of the following questions only **one** of the lettered responses (A–D) is correct.

Select the correct response in each case and mark its code letter by connecting the dots as illustrated on the answer sheet.

1 An element which forms an ion smaller than its atom is

A chlorine.
B potassium.
C oxygen.
D sulfur.

2 A compound which does **not** consist of individual molecules is

A beryllium chloride.
B calcium chloride.
C hydrogen chloride.
D phosphorus trichloride.

3 Which one of the following elements contains the same number of electrons as an ion of magnesium, Mg^{2+}?

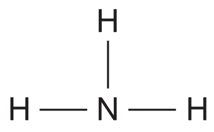
A calcium
B fluorine
C neon
D sodium

4 Which one of the following is the mass of calcium carbonate which will exactly neutralise 500 cm^3 of 0.1 M hydrochloric acid?

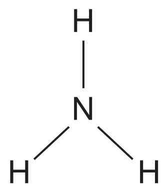
A 1.25 g
B 2.50 g
C 12.50 g
D 25.0 g

5 Which one of the following is the shape of the ammonia molecule?

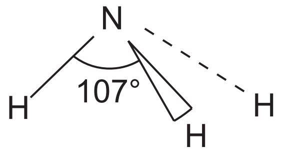
A



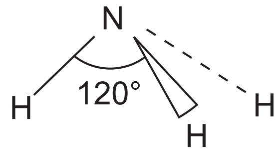
B



C



D



6 If 30 g of water were completely converted into hydrogen and oxygen which one of the following would be the total mass of gases produced?

A 10 g
B 30 g
C 45 g
D 90 g

7 Which one of the following can **not** be used to obtain hydrogen chloride in the laboratory?

A burning hydrogen in chlorine
B heating concentrated hydrochloric acid
C the reaction of chlorine with methane
D bubbling chlorine through hexane at room temperature

8 The element europium reacts with hydrogen to form europium hydride. Atoms of europium have their outer electrons in levels 5 and 6 i.e. $5s^2\ 5p^6\ 6s^2$. Which one of the following formulae resembles europium hydride?

A AsH_3
B CH_4
C CaH_2
D SnH_4

9 Which one of the following atoms in the ground state contains **no** unpaired electrons?

A argon
B fluorine
C potassium
D sulfur

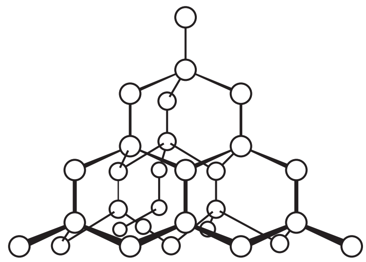
10 Which one of the following species contains a coordinate bond?

A NH_3
B NH_4^+
C NH_2^-
D NH^{2-}

Section B

Answer **all six** questions in this section.

11 The molecule drawn below is that of a giant covalent structure. All the atoms are of the same element.



(a) Name the substance.

____ [1]

(b) Explain whether the substance is hard or soft.

____

____ [2]

(c) Explain whether the substance conducts electricity or not.

____

____ [2]

Examiner Only	
Marks	Remark

Examiner Only	
Marks	Remark

12 Barium chloride crystallises from water to form a hydrate with the formula $BaCl_2.xH_2O$. On heating, the hydrate loses water to form anhydrous barium chloride. A solution of barium chloride is colourless.

(a) Barium chloride solution reacts with aqueous silver nitrate to form silver chloride.

(i) Write the equation for the reaction.

____ [1]

(ii) Write the ionic equation for the reaction.

____ [1]

(iii) Describe what is observed during the reaction.

____ [1]

(b) 3.05 g of $BaCl_2.xH_2O$ were dissolved in water to make 250 cm^3 of solution in a graduated flask. 20 cm^3 of this solution were titrated with 0.1 M silver nitrate solution. It was found that 20.0 cm^3 were required.

(i) How many moles of silver ions were added during the titration?

____ [1]

(ii) How many moles of chloride ions were there in 20 cm^3 of the barium chloride solution?

____ [1]

(iii) How many moles of anhydrous barium chloride were there in 250 cm^3 of the solution?

____ [1]

(iv) What is the relative formula mass of the hydrated barium chloride?

____ [1]

(v) What is the relative formula mass of anhydrous barium chloride?

____ [1]

(vi) Calculate the value of x in $BaCl_2.xH_2O$

____ [1]

(c) Write the equation for the action of heat on the hydrated barium chloride.

____ [1]

Examiner Only	
Marks	Remark

Examiner Only	
Marks	Remark

13 Astatine, the last element of the halogen group, was synthesised in 1940. Since then it has been stated that it is the rarest naturally occurring element on Earth with an estimated 30 g of astatine existing at any one time. It was named from the Greek word for "unstable".

(a) The longest living isotope of astatine is astatine-210, ^{210}At. However, half of this isotope disappears after about 8 hours.

(i) Define the meaning of the term **isotope**.

____

____ [2]

(ii) Name and calculate the numbers of the individual sub-atomic particles in one atom of astatine-210.

____

____ [3]

(b) Using Avogadro's number calculate the number of astatine atoms that exist in 30 g of astatine. Assume that all of the atoms are of astatine-210.

____

____ [2]

(c) Astatine was predicted to exist by Mendeleev in his original Periodic Table and was given the name eka-iodine.

Complete the table below by predicting some of the properties of astatine.

Property	Result for Astatine
formula of an astatine molecule	
physical state at room temperature	
colour of astatine at room temperature	
colour of astatine vapour/gas	
solubility in water (yes or no)	
solubility in hexane (yes or no)	

[6]

(d) Write the equation for the reaction of iodine with sodium astatide.

____ [1]

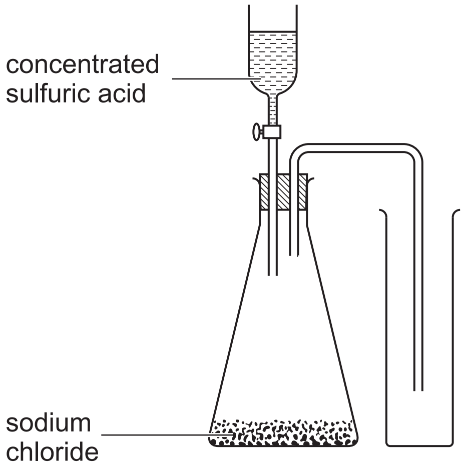
(e) Astatine was first made by bombarding bismuth with alpha particles. Explain how you would show, using mass spectrometry, that astatine had actually been formed.

____

____ [1]

Examiner Only	
Marks	Remark

14 The diagram below shows a common method of preparing hydrogen chloride gas in the laboratory.



(a) Write the equation for the reaction of sodium chloride with concentrated sulfuric acid.

____ [2]

(b) The hydrogen chloride is collected by “downward delivery” in which air, a mixture of oxygen and nitrogen, is displaced upwards.

(i) Calculate the relative molecular masses of oxygen, nitrogen and hydrogen chloride.

oxygen ____

nitrogen ____

hydrogen chloride ______________________________________ [2]

(ii) Use the values of the calculated relative molecular masses to explain why hydrogen chloride is collected by downward delivery.

____

____ [2]

Examiner Only	
Marks	Remark

Examiner Only	
Marks	Remark

(c) Sometimes dry hydrogen chloride gas is required.

(i) Explain why sodium hydroxide would be inappropriate to use as a drying agent.

____ [1]

(ii) Suggest the name of a substance which could be used.

____ [1]

(d) Explain whether concentrated sulfuric acid could be reacted with sodium bromide in the preparation of hydrogen bromide using this method.

____ [1]

(e) Which one of the following acids, hydrogen chloride, hydrogen bromide and hydrogen iodide is the strongest?

____ [1]

(f) How could you prove that a gas jar you believed contained hydrogen chloride actually contained the gas.

____

____ [2]

Examiner Only
Marks | Remark

15 Some properties of the Group I elements from sodium to caesium are shown in the table below.

metal	ionic radius/nm	first ionisation energy/kJ mol^{-1}	melting point/K
sodium	0.102	496	371
potassium	0.138	419	336
rubidium	0.149	403	312
caesium	0.170	376	302

(a) Explain why all of the Group I elements are described as being s-block elements.

______________________________ [1]

(b) Explain why the ionic radius increases down the group.

____

______________________________ [1]

(c) The first ionisation energy of the Group I elements may be determined using spectroscopic methods.

(i) If the frequency of the radiation needed to remove the outermost electron from a sodium atom is 1.25×10^{15} s^{-1} calculate the first ionisation energy of sodium in kJ per mole.

____

____

____

______________________________ [3]

(ii) Write the equation, using state symbols, for the first ionisation energy of sodium.

______________________________ [2]

Examiner Only	
Marks	Remark

(iii) Give **two** reasons to explain why potassium has a lower first ionisation energy than sodium.

____

____

____ [2]

(iv) Why is the second ionisation energy of all the Group I metals very much higher than the first?

____

____ [1]

(d) All of the Group I metals exhibit metallic bonding.

(i) Using a labelled diagram, explain what is meant by the term **metallic bonding**.

____

____ [3]

(ii) Suggest why the melting point of the metals decreases from sodium to caesium.

____

____ [1]

(iii) Using the concept of metallic bonding suggest why calcium should be a better electrical conductor than potassium.

____

____ [1]

Examiner Only	
Marks	Remark

(e) All of the Group I metals react with halogens to form ionic metal halides. Using outer electrons only draw diagrams to explain the formation of caesium chloride from caesium atoms and chlorine atoms.

[3]

(f) All of the Group I elements produce characteristic flame colours in a Bunsen burner flame which can be used to identify them.

(i) Describe how you would carry out a flame test.

____

____

____

____

____

____ [4]

Quality of written communication [2]

(ii) How would you distinguish between sodium chloride and potassium chloride using a flame test?

____

____ [1]

Examiner Only	
Marks	Remark

16 Carbon dioxide is the most frequently found oxide of carbon in nature. It is a colourless gas with a faint taste and smell. The structure of the molecule can be readily deduced by the application of the octet rule. Even though carbon and oxygen have different electronegativities the molecule does not have a permanent dipole.

(a) Explain the term **octet rule**.

____

____ [2]

(b) Explain the term **electronegativity**.

____

____ [2]

(c) **(i)** Using outer electrons only draw the dot and cross structure of carbon dioxide.

[2]

(ii) Draw and name the shape of a carbon dioxide molecule.

____ [2]

(iii) Explain why carbon dioxide has the shape you have drawn.

____

____ [2]

(d) Carbon and oxygen have different electronegativities and form a polar bond.

Explain why a carbon dioxide molecule does not have a permanent dipole.

____

____ [1]

(e) Although carbon dioxide does not have a dipole it is very soluble in water. Using intermolecular forces explain this extreme solubility.

____

____ [2]

Examiner Only	
Marks	Remark

THIS IS THE END OF THE QUESTION PAPER

111292

Rewarding Learning

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
2011

Chemistry

Assessment Unit AS 1

assessing

Basic Concepts in Physical and Inorganic Chemistry

[AC112]

WEDNESDAY 15 JUNE, AFTERNOON

MARK SCHEME

Section A

		AVAILABLE MARKS
1	B	
2	B	
3	C	
4	B	
5	C	
6	B	
7	D	
8	C	
9	A	
10	B	
[2] for each correct answer	[20]	20
	Section A	**20**

Section B

AVAILABLE MARKS

11 (a) diamond [1]

(b) hard [1]
strong bonds [1] [2]

(c) no [1]
electrons (in bonds) cannot move [1] [2] 5

12 (a) (i) $BaCl_2 + 2AgNO_3 \rightarrow 2AgCl + Ba(NO_3)_2$ [1]

(ii) $Ag^+ + Cl^- \rightarrow AgCl$ [1]

(iii) White precipitate [1]

(b) (i) $20 \times 10^{-3} \times 10^{-1} = 2 \times 10^{-3}$ mol [1]

(ii) 2×10^{-3} mol [1]

(iii) $\frac{1}{2} \times 2 \times 10^{-3} \times \frac{250}{20} = 1.25 \times 10^{-2}$ [1]

(iv) $0.0125 \equiv 3.05\,g$

$1 \text{ mol} \equiv \frac{3.05}{0.0125} = 244$ [1]

(v) $BaCl_2 = 137 + 71 = 208$ [1]

(vi) $244 - 208 = 36 \quad 36 = H_2O$
$\therefore x = 2$ [1]

(c) $BaCl_2.xH_2O \rightarrow BaCl_2 + xH_2O$ [1] 10
or $BaCl_2.2H_2O \rightarrow BaCl_2 + 2H_2O$

						AVAILABLE MARKS
13	**(a)**	**(i)**	same number of protons [1] different number of neutrons [1]		[2]	
		(ii)	85 protons 85 electrons 125 neutrons [−1] for each wrong part		[3]	
	(b)		$30\,g = \frac{30}{210} = 0.1429$ mol $6.023 \times 10^{23} \times 0.1429 = 0.86 \times 10^{23} = 8.6 \times 10^{22}$		[2]	
	(c)		At_2 solid grey-black/black purple/violet/dark violet no yes	[1] [1] [1] [1] [1] [1]	[6]	
	(d)		$I_2 + 2NaAt \rightarrow 2NaI + At_2$		[1]	
	(e)		new peak at 210 or round about		[1]	15
14	**(a)**		$NaCl + H_2SO_4 \rightarrow NaHSO_4 + HCl$ use of Na_2SO_4 = [1]		[2]	
	(b)	**(i)**	$O_2 = 32 \quad N_2 = 28 \quad HCl = 36.5$ wrong value is [−1]		[2]	
		(ii)	HCl is heavier than N_2 or O_2 [1] hence HCl sinks [1]		[2]	
	(c)	**(i)**	NaOH reacts with HCl/acid + base [1]			
		(ii)	conc H_2SO_4/anhydrous $CuSO_4$ etc. [1]		[2]	
	(d)		no – HBr reacts with H_2SO_4		[1]	
	(e)		hydrogen iodide		[1]	
	(f)		use $AgNO_3(aq)$ or conc $NH_3(aq)$ [1] white ppt/white smoke [1]		[2]	12

				AVAILABLE MARKS
15	**(a)**	outer electrons are s electrons	[1]	
	(b)	increased number of shells	[1]	
	(c)	**(i)** $E = hf = 6.63 \times 10^{-34} \times 1.25 \times 10^{15}$ $= 8.29 \times 10^{-19}$ J For one mole $= 6.023 \times 10^{23} \times 8.29 \times 10^{-19}$ J $= 49.9 \times 10^{4}$J $= 499$	[3]	
		(ii) $Na(g) \rightarrow Na^{+}(g) + e^{-}$	[2]	
		(iii) outer electrons further away from the nucleus [1] shielded by increased shells of electrons [1]	[2]	
		(iv) removal of second electron is from a full shell	[1]	
	(d)	**(i)** e^{-} e^{-} e^{-} e^{-} ⊕ ⊕ ⊕ ⊕ e^{-} e^{-} e^{-} e^{-} ⊕ ⊕ ⊕ ⊕ [2] electrons are delocalised/can move/ (electrostatic) attraction between e^{-} and metal ion [1]	[3]	
		(ii) forces of attraction decreases charge density decreases	[1]	
		(iii) Ca has two outer electrons	[1]	
	(e)	$\dot{Cs} + \overset{xx}{\underset{x}{{}^{x}_{x}Cl^{x}_{x}}} \rightarrow Cs^{+} + \left[\overset{xx}{\underset{\cdot x}{{}^{x}_{x}Cl^{x}_{x}}}\right]^{-}$	[3]	
	(f)	**(i)** nichrome/platinum wire [1] blue flame (of Bunsen) [1] conc. hydrochloric acid [1] place compound on wire/put in blue flame [1]	[4]	
		Quality of written communication	[2]	
		(ii) potassium → lilac **or** sodium → yellow/orange	[1]	25

16 **(a)** 8 electrons [1]
in outer shell [1] [2]

(b) attraction of electrons by an atom [1]
in a covalent bond [1] [2]

(c) **(i)**  [2]

(ii) O ═ C ═ O or O — C — O [1]
linear/straight [1] [2]

(iii) electrons in the bonds [1]
repel as much as possible [1] [2]

(d) the dipoles "cancel" out [1]

(e) attraction between
$\delta+$ on C and $\delta-$ on O in H_2O
or $\delta-$ on O and $\delta+$ on H in H_2O [2]

	AVAILABLE MARKS
	13
Section B	**80**
Total	**100**

Rewarding Learning

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
January 2012

Chemistry

Assessment Unit AS 1

assessing

Basic Concepts in Physical and Inorganic Chemistry

[AC112]

FRIDAY 13 JANUARY, AFTERNOON

AC112

Centre Number

71	

Candidate Number

TIME

1 hour 30 minutes.

INSTRUCTIONS TO CANDIDATES

Write your Centre Number and Candidate Number in the spaces provided at the top of this page.
Answer **all fifteen** questions.
Answer **all ten** questions in **Section A**. Record your answers by marking the appropriate letter on the answer sheet provided. Use only the spaces numbered 1 to 10. Keep in sequence when answering.
Answer **all five** questions in **Section B**. Write your answers in the spaces provided in this question paper.

INFORMATION FOR CANDIDATES

The total mark for this paper is 100.
Quality of written communication will be assessed in question **11**.
In Section A all questions carry equal marks, i.e. **two** marks for each question.
In Section B the figures in brackets printed down the right-hand side of pages indicate the marks awarded to each question or part question.
A Periodic Table of Elements (including some data) is provided.

For Examiner's use only	
Question Number	Marks
Section A	
1–10	
Section B	
11	
12	
13	
14	
15	
Total Marks	

Section A

For each of the following questions only **one** of the lettered responses (A–D) is correct.

Select the correct response in each case and mark its code letter by connecting the dots as illustrated on the answer sheet.

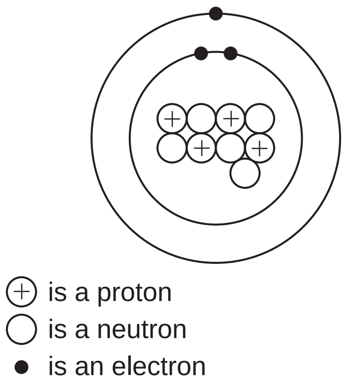
1 Which one of the following bonds is the most polar?

A B-F
B N-F
C C-I
D O-I

2 Which one of the following can **not** form hydrogen bonds?

A H_2O
B H_3O^+
C NH_3
D NH_4^+

3 Which one of the following is the name of the species shown below?



A beryllium atom
B beryllium ion
C lithium atom
D lithium ion

4 Which one of the following solids consists of molecular covalent crystals?

A Diamond
B Graphite
C Ice
D Quartz

5 When excess chlorine is bubbled into hot concentrated alkali which one of the following lists the main products of the reaction?

A Cl^-, ClO^-, H_2O
B Cl^-, ClO_3^-, H_2O
C Cl^-, ClO_4^-, H_2O
D ClO^-, ClO_3^-, H_2O

6 The elements X and Y are in Groups VI and VII respectively of the Periodic Table.

Which one of the following shows the formula and the bond type of the compound that they form?

A XY_2, covalent
B XY_2, ionic
C X_2Y, covalent
D X_2Y, ionic

7 Which one of the following orbitals is occupied by an electron with the energy level $n = 2$?

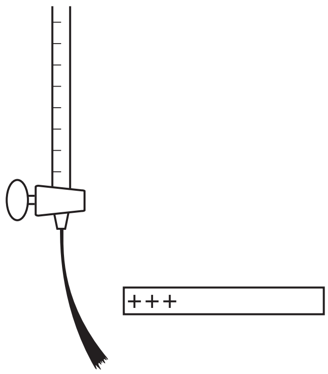
A A dumb-bell shaped orbital
B A spherically shaped orbital
C An s or d orbital
D An s or p orbital

8 A crystalline solid melts sharply at 95 °C. It does not conduct electricity in the solid and liquid states. It dissolves in hexane.

Which one of the following is the structure of the crystal?

A giant molecular
B ionic
C metallic
D molecular covalent

9 The diagram below shows a liquid escaping from a burette and passing a charged glass rod.



Which one of the following liquids will be attracted to the glass rod?

A CCl_4
B $CHCl_3$
C CS_2
D C_5H_{12}

10 The species Ar, K^+ and Ca^{2+} have the same number of electrons. Starting with the smallest, which one of the following is the order in which their radii increase?

A	Ar	Ca^{2+}	K^+
B	Ar	K^+	Ca^{2+}
C	Ca^{2+}	K^+	Ar
D	K^+	Ar	Ca^{2+}

Section B

Answer all **five** questions in this section.

Examiner Only	
Marks	Remark

11 The geometry of covalent inorganic hydrides may be predicted using the electron structures of the molecules.

Draw the shapes of the following hydrides using their outer electron structures. Explain these shapes giving the values of the angles between bonds.

hydrogen chloride	HCl
methane	CH_4
hydrogen sulfide	H_2S
ammonia	NH_3

[6]

Quality of written communication [2]

Examiner Only	
Marks	Remark

12 Solutions of acidified iodide ions are very easily oxidised to produce iodine molecules.
Even oxygen, from the air, will oxidise iodide ions to liberate iodine.
The following half-equations represent the formation of iodine molecules and the conversion of oxygen to water.

Equation 1 $2I^- \rightarrow I_2 + 2e^-$

Equation 2 $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$

(a) (i) Using electron transfer explain which equation represents an oxidation reaction.

____ [1]

(ii) Using electron transfer explain which equation represents a reduction reaction.

____ [1]

(b) Write the equation for the reaction of acidified iodide ions with oxygen.

____ [2]

(c) The following solutions were added to a solution of acidified iodide ions in separate test tubes.

chlorine, iron(**III**) ions, ammonia, sodium hydroxide, sodium chloride.

Which **two** of these solutions would react with acidified iodide ions to produce iodine?

____ [2]

(d) The reaction of oxidising agents with potassium iodide can be used to prepare iodine in the laboratory. When heated with manganese dioxide and concentrated sulfuric acid, potassium iodide liberates iodine.

$$2KI + MnO_2 + 3H_2SO_4 \rightarrow 2KHSO_4 + MnSO_4 + 2H_2O + I_2$$

(i) Using oxidation numbers, explain this redox reaction.

____

____

____ [3]

(ii) What observation would confirm that iodine had been produced?

____ [1]

Examiner Only	
Marks	Remark

13 Phosphorus, P, reacts with bromine at room temperature to form phosphorus tribromide, PBr_3, which is a liquid with boiling point 173 °C. It reacts with water immediately forming hydrogen bromide and phosphoric(III) acid, H_3PO_3. The reaction is used as a method of preparing hydrogen bromide in the laboratory.

Examiner Only	
Marks	Remark

(a) Write the equation for the reaction of bromine with phosphorus.

____ [1]

(b) Calculate the maximum mass of phosphorus tribromide which can be formed when 6.2 g of phosphorus, *which is an excess*, reacts with 8.0 cm^3 of bromine, Br_2. The density of liquid bromine is 3.1 $g\,cm^{-3}$.

mass of bromine, Br_2, in grams

____ [1]

moles of bromine, Br_2

____ [1]

moles of phosphorus, P, in 6.2 g

____ [1]

moles of bromine, Br_2 reacting

____ [1]

moles of phosphorus tribromide formed

____ [1]

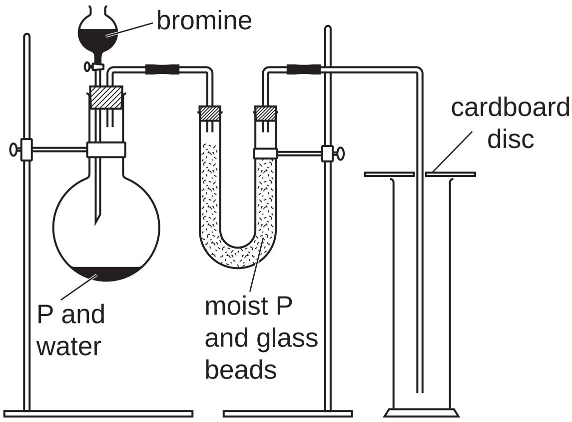
mass of phosphorus tribromide formed

____ [1]

(c) Write the equation for the reaction of phosphorus tribromide with water.

____ [1]

(d) The apparatus shown below was used to prepare hydrogen bromide in the laboratory. Bromine is slowly added to a paste of phosphorus and water. Phosphorus tribromide is first formed and is immediately decomposed by the water present. The gases produced are passed through a U-tube containing glass beads coated in phosphorus.



(i) Suggest why the bromine is not added all at once.

____ [1]

(ii) An excess of water is not used in the experiment. What is the property of hydrogen bromide which is the reason for not using an excess?

____ [1]

(iii) Why is the hydrogen bromide collected as shown and not with the delivery tube pointing upwards?

____ [1]

(e) Hydrogen bromide is a colourless gas but produces fumes in moist air. Why does it fume in moist air?

____ [1]

Examiner Only	
Marks	Remark

Examiner Only	
Marks	Remark

(f) When hydrogen bromide is heated in a loosely corked test tube it produces a very pale red–brown colour after heating for several minutes.

(i) Name the substance responsible for the red–brown colour.

____ [1]

(ii) Explain what would be observed if a test tube of hydrogen iodide was heated.

____ [1]

(iii) Explain what would be observed if a test tube of hydrogen chloride was heated.

____ [1]

(iv) What do these observations indicate about the relative thermal stability of the hydrogen halides?

____

____ [1]

14 Only 0.08% of the Earth's crust consists of carbon yet this element is an essential part of living organisms. It occurs naturally as the isotopes carbon-12 and carbon-13 although there is a radioactive isotope carbon-14. Carbon occurs in nature as two structures known as diamond and graphite.

Examiner Only	
Marks	Remark

(a) Naturally occurring carbon contains 98.89% of carbon-12 and 1.11% carbon-13. Calculate the relative atomic mass of carbon to three decimal places.

____

____

____ [3]

(b) Carbon-14 is not used in the calculation of the relative atomic mass because virtually none of it exists. It decomposes when a neutron in its nucleus changes into an electron and a proton forming a new element.

(i) What are the numbers of electrons, protons and neutrons in the new element?

____ [2]

(ii) Name the element produced when carbon-14 decomposes.

____ [1]

(c) Mass spectrometry uses carbon-12 as the international standard.

(i) What is the purpose of mass spectrometry?

____

____ [2]

(ii) Explain the meaning of the term **carbon-12 standard**.

____

____ [2]

(d) Explain why carbon-12 and carbon-14 are isotopes.

____

____ [2]

(e) Carbon may be produced in the laboratory in many ways. One is to heat cane sugar, $C_{12}H_{22}O_{11}$, with concentrated sulfuric acid. Steam and carbon are produced together with diluted sulfuric acid.

Examiner Only	
Marks	Remark

(i) Write the equation for the reaction. Do not include sulfuric acid in the equation.

______________________________ [1]

(ii) Explain the meaning of the terms **hydrated** and **water of crystallisation**.

____

____

____

______________________________ [2]

(iii) Explain whether the cane sugar is hydrated.

____

______________________________ [1]

(f) Diamond is oxidised when it burns in oxygen at about 700 °C.

(i) Name the product formed from the complete oxidation of diamond.

______________________________ [1]

(ii) Name the product formed from the incomplete oxidation of diamond.

______________________________ [1]

(iii) Explain whether graphite will form the same products when it is burned.

____

______________________________ [2]

Examiner Only	
Marks	Remark

(g) Draw dot and cross diagrams, using outer electrons only, to show the formation of a carbon dioxide molecule from a carbon atom and an oxygen molecule.

[3]

Examiner Only	
Marks	Remark

15 Anhydrous copper(II) chloride, $CuCl_2$, may be prepared by heating copper in chlorine gas. When prepared by dissolving copper(II) oxide in hydrochloric acid, copper(II) chloride crystallises with two molecules of water of crystallisation.

(a) Write the equation for the reaction of copper with chlorine.

______________________________ [1]

(b) Write the equation for the reaction of copper(II) oxide with hydrochloric acid.

______________________________ [1]

(c) Write the formula for hydrated copper(II) chloride.

______________________________ [1]

(d) The purity of the copper(II) oxide may be determined by the process of back titration. Explain, without calculations, how this process would be carried out.

____

____

____

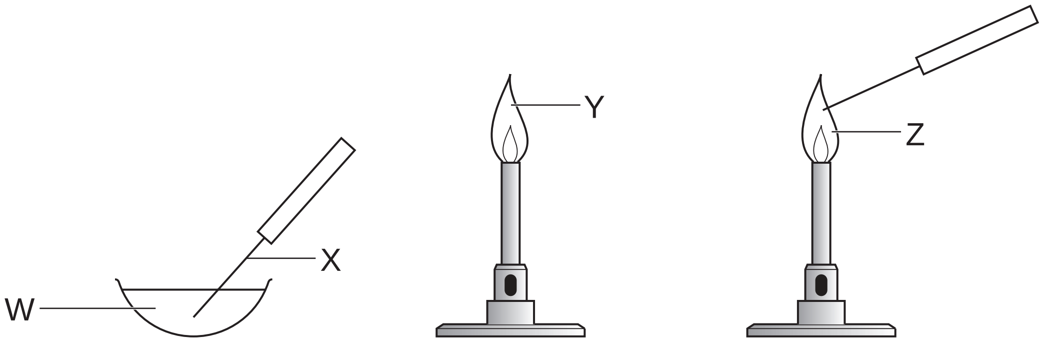
____

______________________________ [4]

Examiner Only	
Marks	Remark

(e) The presence of copper in copper(II) chloride can be shown using a flame test.

(i) The diagram below shows the equipment needed for the test. Identify the acid W, the metal wire X, the colour Y of the flame before the test and the colour Z during the test.



W ______________________________ [1]

X ______________________________ [1]

Y ______________________________ [1]

Z ______________________________ [1]

(ii) State **two** reasons for using W.

____

____

______________________________ [2]

(iii) Explain the origin of the flame colour produced by copper(II) chloride.

____

____

____

______________________________ [3]

Examiner Only	
Marks	Remark

(f) When copper(II) chloride dissolves in water its ions are surrounded by water molecules. The polar water molecules surround both anion and cation.

(i) State the formulae of the ions present in copper(II) chloride.

____ [2]

(ii) Draw a diagram showing two water molecules around the anion.

[1]

(iii) Draw a diagram showing two water molecules around the cation.

[1]

(iv) Explain how you would confirm the presence of chloride ions in the solution.

____

____

____

____ [3]

THIS IS THE END OF THE QUESTION PAPER

Rewarding Learning

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
January 2012

Chemistry

Assessment Unit AS 1

assessing

Basic Concepts in Physical and Inorganic Chemistry

[AC111]

FRIDAY 13 JANUARY, AFTERNOON

MARK SCHEME

Section A

			AVAILABLE MARKS
1	A		
2	D		
3	B		
4	C		
5	B		
6	A		
7	D		
8	D		
9	B		
10	C		
[2] for each correct answer		[20]	20
		Section A	**20**

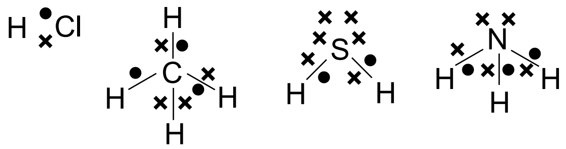
Section B

AVAILABLE MARKS

11 Depending on the response of candidates it is likely that two marking points will be needed for each mark awarded.

shapes	H–Cl	C	S	N
angles	180°	109°	105°	107°
	(accept no angle)			

dot and cross diagram



apply the following to each compound
lone pair v lone pair > lone pair v bond pair > bond pair v bond pair

the electron pairs repel to be as far apart as possible [6]

4 marking points per compound, i.e. shape, electron structure, angle, explanation, i.e. 16 marking points – count number of errors. Apply following:
Even number of errors ÷ 2, subtract this from 6
(Odd number of errors –1) ÷ 2, subtract this from 6

Quality of written communication [2] 8

12 **(a)** **(i)** $2I^- \rightarrow I_2 + 2e$ oxidation because electrons are lost [1]

(ii) $O_2 \rightarrow 2H_2O$ reduction because electrons are gained [1]

(b) $4I^- + 4H^+ + O_2 \rightarrow 2I_2 + 2H_2O$ (electrons left in [–1]) [2]

(c) chlorine [1]
iron(III) ions [1] [2]

(d) **(i)** $I^- = -1$ $MnO_2 = +4$
$I_2 = 0$ $MnSO_4 = +2$ [2]

iodide/iodine is oxidised and manganese is reduced [1] [3]

(ii) violet/purple vapour
or grey/black solid at top of test tube [1] 10

AVAILABLE MARKS

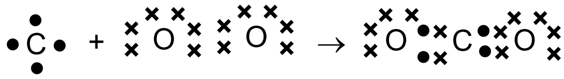
13	**(a)**		$2P + 3Br_2 \rightarrow 2PBr_3$ **or** $P_4 + 6Br_2 \rightarrow 4PBr_3$	[1]	
	(b)		$8.0 \times 3.1 = 24.8$ g	[1]	
			$24.8/160 = 0.155$ mol	[1]	
			$6.2/31 = 0.2$ mol	[1]	
			0.155 mol	[1]	
			0.103 mol	[1]	
			$PBr_3 = 31 + 3 \times 80 = 271$ $0.103 \times 271 = 27.9$ g	[1]	
	(c)		$PBr_3 + 3H_2O \rightarrow 3HBr + H_3PO_3$	[1]	
	(d)	**(i)**	reaction could be too vigorous	[1]	
		(ii)	hydrogen bromide is soluble (in water)	[1]	
		(iii)	hydrogen bromide is heavier (than air)	[1]	
	(e)		dissolves (in water vapour) to form hydrobromic acid	[1]	
	(f)	**(i)**	bromine	[1]	
		(ii)	violet/purple colour	[1]	
		(iii)	nothing observed/stays the same/remains colourless	[1]	
		(iv)	HCl > HBr > HI (mark is dependent on given observations)	[1]	16
14	**(a)**		$98.89 \times 12 = 1186.68$ $1.11 \times 13 = 14.43$ $= 1201.11$ $= 12.011$	[3]	
	(b)	**(i)**	7 electrons 7 protons 7 neutrons	[2]	
		(ii)	nitrogen	[1]	
	(c)	**(i)**	to determine RAM and isotopic abundance/RMM	[2]	
		(ii)	atomic masses or RAM/mol mass/RMM are measured relative to C = 12.000	[2]	
	(d)		same atomic number but different mass numbers	[2]	
	(e)	**(i)**	$C_{12}H_{22}O_{11} \rightarrow 12C + 11H_2O$	[1]	
		(ii)	hydrated: contains water of crystallisation/water present [1] water of crystallisation: water chemically bonded [1]	[2]	
		(iii)	not hydrated, water is formed/no water in the sugar	[1]	

AVAILABLE MARKS

(f) **(i)** carbon dioxide [1]

(ii) carbon monoxide [1]

(iii) yes [1] it is also carbon [1] [2]

(g)  [3] 23

15 **(a)** $Cu + Cl_2 \rightarrow CuCl_2$ [1]

(b) $CuO + 2HCl \rightarrow CuCl_2 + H_2O$ [1]

(c) $CuCl_2.2H_2O$ [1]

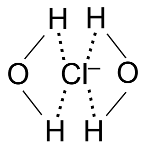
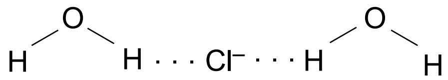
(d) Weigh the CuO, add (known) excess (hydrochloric) acid (to CuO) [1]
titrate excess hydrochloric acid [1]
with (standard) alkali/sodium hydroxide [1]
named indicator, e.g. phenolphthalein/methyl orange [1] [4]

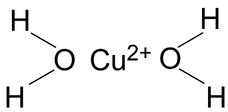
(e) **(i)** W: concentrated hydrochloric acid [1]
X: nichrome/platinum [1]
Y: blue [1]
Z: green-blue [1]

(ii) clean the wire [1]
make the solid stick to the wire/dissolve the solid [1] [2]

(iii) electrons (in the energy levels) raised to higher levels [1]
fall back down [1] to give out light [1] [3]

(f) **(i)** Cu^{2+}; Cl^- [2]

(ii)  **or**  [1]

(iii)  [1]

(iv) add silver nitrate (solution) [1]
(add dilute nitric acid)
white [1]
precipitate/solid [1] [3] 23

Section B **80**

Total **100**

Rewarding Learning

Centre Number

71	

Candidate Number

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
2012

Chemistry

Assessment Unit AS 1

assessing

Basic Concepts in Physical and Inorganic Chemistry

[AC112]

AC112

WEDNESDAY 13 JUNE, MORNING

TIME

1 hour 30 minutes.

INSTRUCTIONS TO CANDIDATES

Write your Centre Number and Candidate Number in the spaces provided at the top of this page.
Answer **all fifteen** questions.
Answer **all ten** questions in **Section A**. Record your answers by marking the appropriate letter on the answer sheet provided. Use only the spaces numbered 1 to 10. Keep in sequence when answering.
Answer **all five** questions in **Section B**. Write your answers in the spaces provided in this question paper.

INFORMATION FOR CANDIDATES

The total mark for this paper is 100.
Quality of written communication will be assessed in question **14(b)(i)**.
In Section A all questions carry equal marks, i.e. **two** marks for each question.
In Section B the figures in brackets printed down the right-hand side of pages indicate the marks awarded to each question or part question.
A Periodic Table of Elements (including some data) is provided.

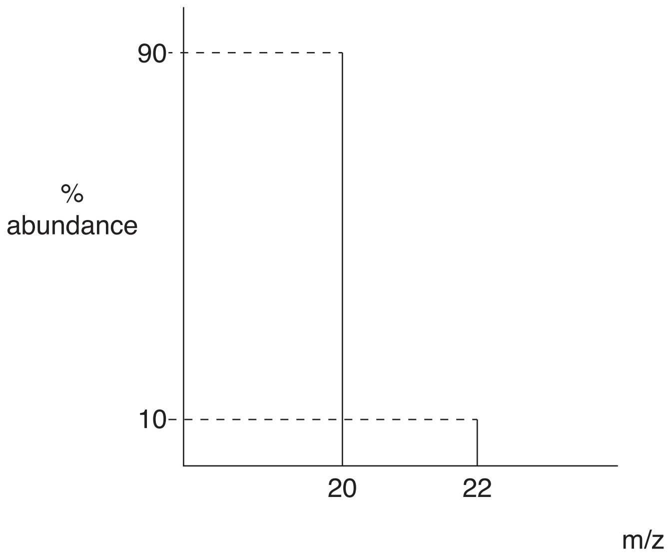
For Examiner's use only	
Question Number	**Marks**
Section A	
1–10	
Section B	
11	
12	
13	
14	
15	
Total Marks	

Section A

For each of the following questions only **one** of the lettered responses (A–D) is correct.

Select the correct response in each case and mark its code letter by connecting the dots as illustrated on the answer sheet.

1 Part of the mass spectrum for an element is shown below:



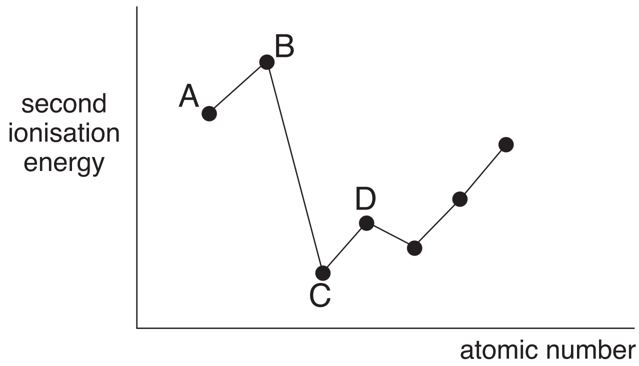
Which one of the following is the relative atomic mass of the element?

A 20.0
B 20.2
C 21.0
D 22.8

2 Which one of the following metal compounds will produce a lilac flame colour?

A barium nitrate
B calcium chloride
C lithium chloride
D potassium sulfate

3 The graph below shows how the **second** ionisation energy of elements varies across a period.



Which one of the elements is an alkali metal?

4 Which one of the following is the oxidation number of nitrogen in the nitrate ion, NO_3^-?

A -1
B -3
C $+5$
D $+7$

5 Which one of the following molecules is the most polar?

A BF_3

B CO_2

C F_2

D NH_3

6 Which one of the following m/z values will **not** appear when a sample of chlorine gas is injected into a mass spectrometer?

A 35.0
B 35.5
C 37.0
D 74.0

7 Which one of the following molecules contains the smallest bond angle?

A $BeCl_2$

B BF_3

C CH_4

D SF_6

8 5.30 g of anhydrous sodium carbonate was dissolved in water and made up to 250 cm^3 in a volumetric flask. Which one of the following is the concentration of **sodium ions** in mol dm^{-3}?

A 0.05
B 0.10
C 0.20
D 0.40

9 Which block in the Periodic Table contains silver?

A d block
B f block
C p block
D s block

10 Which one of the following is involved in metallic bonding?

A electron delocalisation
B electron transitions
C gaining electrons to form ions
D sharing electron pairs

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(Questions continue overleaf)

Section B

Answer **all five** questions in the spaces provided.

Examiner Only	
Marks	Remark

11 The elements magnesium and chlorine are characterised by their atomic numbers. Chlorine has two isotopes each with a different mass number.

(a) Define each of the following in terms of protons, neutrons and electrons.

(i) atomic number

____ [1]

(ii) mass number

____ [1]

(iii) isotopes

____

____ [1]

(b) Magnesium reacts with chlorine to form magnesium chloride.

(i) How many protons, neutrons and electrons are present in each of the following ions?

Ion	Numbers of		
	protons	**neutrons**	**electrons**
$^{24}Mg^{2+}$			
$^{35}Cl^{-}$			

[2]

(ii) Use the boxes below to complete the electronic configuration of the ions:

	1s	2s	2p	3s	3p	
$^{24}Mg^{2+}$	☐	☐	☐☐☐	☐	☐☐☐	[1]
$^{35}Cl^{-}$	☐	☐	☐☐☐	☐	☐☐☐	[1]

(iii) Use a dot and cross diagram to show, using outer electrons only, how magnesium atoms react with chlorine atoms to form magnesium chloride.

[4]

(c) Magnesium forms ions with a double positive charge.

(i) Define the term **second ionisation energy**.

____

____ [2]

(ii) Write an equation, including state symbols, which represents the second ionisation energy of magnesium.

____ [2]

(iii) Give reasons why the third ionisation energy of magnesium is much larger than the second.

____

____

____

____ [3]

(d) The Group II chloride, $SrCl_2$, produces a characteristic red colour in a Bunsen flame. Explain, using energy levels, why this colour is observed.

____

____

____

____ [3]

Examiner Only	
Marks	Remark

12 Avogadro's number has the value 6.02×10^{23}.

Examiner Only	
Marks	Remark

(a) Define the term **Avogadro's number**.

____

____ [2]

(b) X is an oxide of nitrogen.

(i) 2.30 g of X contains 3.01×10^{22} molecules of X. Calculate the molar mass of X.

____

____

____ [2]

(ii) Deduce the formula of X.

____ [1]

(c) Dinitrogen tetroxide (N_2O_4) reacts with water to form nitric acid, (HNO_3) and nitrogen(II) oxide (NO). Write an equation for the reaction.

____ [1]

(d) Dilute nitric acid reacts with magnesium:

$$Mg + 2HNO_3 \rightarrow Mg(NO_3)_2 + H_2$$

(i) Calculate the volume, in cm^3, of $2.0\,mol\,dm^{-3}$ nitric acid required to react with 6.0 g of magnesium.

Number of moles of magnesium

____ [1]

Number of moles of nitric acid

____ [1]

Volume of nitric acid (in cm^3)

____ [1]

(ii) Calculate the mass of magnesium nitrate produced.

Number of moles of magnesium nitrate produced

____ [1]

Mass of magnesium nitrate produced

____

____ [1]

Examiner Only	
Marks	Remark

13 Calcium carbonate is present in eggshells. The percentage of calcium carbonate may be determined by a back titration method. The eggshells are crushed, weighed and then treated with excess dilute hydrochloric acid.

$$CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$$

The unreacted acid is then titrated with standard sodium hydroxide solution.

$$HCl + NaOH \rightarrow NaCl + H_2O$$

(a) **(i)** Explain the term **standard solution**.

____ [1]

(ii) Name a suitable indicator for the titration and state the colour change occurring at the end point.

indicator _______________________________________ [1]

from ____

to ____ [2]

(b) A student weighed out 10.0 g of the crushed eggshells and added 100.0 cm^3 of 2.0 mol dm^{-3} hydrochloric acid. The resultant solution was transferred to a 250 cm^3 volumetric flask and made up to the mark with distilled water. 25.0 cm^3 portions of the solution were titrated with 0.10 mol dm^{-3} sodium hydroxide solution. The average titre was found to be 18.0 cm^3.

(i) Calculate the number of moles of sodium hydroxide used in the titration.

____ [1]

(ii) Calculate the number of moles of hydrochloric acid present in the 25.0 cm^3 portion.

____ [1]

(iii) Calculate the number of moles of hydrochloric acid present in the 250 cm^3 volumetric flask.

____ [1]

Examiner Only	
Marks	Remark

Examiner Only	
Marks	Remark

(iv) Calculate the total number of moles of hydrochloric acid added to the crushed eggshells.

____ [1]

(v) Calculate the number of moles of hydrochloric acid which reacted with the calcium carbonate in the crushed eggshells.

____ [1]

(vi) Calculate the number of moles of calcium carbonate in the crushed eggshells.

____ [1]

(vii) Calculate the mass of calcium carbonate in the crushed eggshells.

____ [1]

(viii) Calculate the percentage, by mass, of calcium carbonate in the crushed eggshells.

____ [1]

14 The halogens are reactive non-metals which often react by gaining electrons to form halide ions.

(a) Complete the table to show the colours and physical states of chlorine, bromine and iodine at room temperature and pressure.

Halogen	**Colour**	**Physical State**
Chlorine		
Bromine		
Iodine		

[3]

(b) Solutions of silver nitrate and ammonia can be used to test for the presence of aqueous halide ions.

(i) Describe how you would use these reagents to distinguish between solutions of sodium chloride, sodium bromide and sodium iodide. State the expected result for each solution.

____

____

____

____

____

____

____ [6]

Quality of written communication [2]

(ii) Give an ionic equation, including state symbols, for the reaction of aqueous sodium iodide with silver nitrate solution.

____ [2]

Examiner Only	
Marks	Remark

(c) Solid samples of sodium chloride, sodium bromide and sodium iodide can be distinguished using concentrated sulfuric acid.

Examiner Only	
Marks	Remark

(i) Write an equation for the reaction of sodium chloride with concentrated sulfuric acid.

____ [2]

(ii) Balance the following half-equation for the reduction of concentrated sulfuric acid to form hydrogen sulfide:

$$H_2SO_4 \quad + \quad H^+ \quad + \qquad\qquad \rightarrow \quad H_2S \quad + \quad H_2O$$

[2]

(iii) Combine the reduction half-equation in **(c)(ii)** with the following oxidation half-equation to produce a balanced redox equation.

$$2I^- \rightarrow I_2 + 2e^-$$

____ [2]

(iv) Give **one** observation which indicates the formation of hydrogen sulfide.

____ [1]

(v) Name **two** other reduction products which are formed when concentrated sulfuric acid is added to sodium iodide.

____

____ [2]

(vi) Suggest why iodide ions are stronger reducing agents than chloride ions.

____

____

____ [2]

(d) **(i)** Write the equation for the reaction of chlorine with hot concentrated sodium hydroxide solution.

____ [2]

(ii) Name the type of redox reaction taking place.

____ [1]

Examiner Only	
Marks	Remark

Examiner Only	
Marks	Remark

15 The bonding and shape of a water molecule determines the properties of water.

(a) Draw a dot and cross diagram to show the bonding in a water molecule.

[2]

(b) **(i)** What is the bond angle in a water molecule?

____ [1]

(ii) State the shape of a water molecule and explain why it adopts this shape.

____

____

____ [3]

(iii) Why is the bond angle of water different to the bond angle in methane?

____

____ [1]

(c) Why does water have a higher boiling point than hydrogen sulfide?

____

____

____ [2]

THIS IS THE END OF THE QUESTION PAPER

Rewarding Learning

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
2012

Chemistry

Assessment Unit AS 1

assessing

Basic Concepts in Physical and Inorganic Chemistry

[AC112]

WEDNESDAY 13 JUNE, MORNING

MARK SCHEME

Section A

		AVAILABLE MARKS
1	B	
2	D	
3	B	
4	C	
5	D	
6	B	
7	D	
8	D	
9	A	
10	A	
[2] for each correct answer	[20]	20
	Section A	**20**

Section B

AVAILABLE MARKS

11 **(a)** **(i)** number of protons [1]

(ii) number of protons + number of neutrons [1]

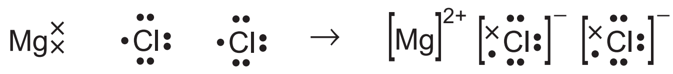
(iii) atoms with the same number of protons but a different number of neutrons [1]

(b) **(i)** 12 12 10
17 18 18 error [–1] [2]

(ii)

	1s	2s	2p			3s	3p			
$^{24}Mg^{2+}$	↑↓	↑↓	↑↓	↑↓	↑↓					[1]
$^{35}Cl^{-}$	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	[1]

(iii)



[4]

[–1] each error

(c) **(i)** The energy required to convert one mole of gaseous ions with a single positive charge into (gaseous) ions with a double positive charge [2]

(ii) $Mg^{+}(g) \rightarrow Mg^{2+}(g) + e^{-}$ [2]

(iii) electron closer to nucleus [1]
less shielded [1]
full shell [1] [3]

(d) electrons promoted from ground state/lower energy level to excited state/higher energy level [1] as the electron drops back down [1] energy given out as (red) light [1] [3] 21

12 **(a)** number of atoms [1] present in 12.000 g of carbon-12 [1] [2]

(b) **(i)** moles of X = 0.05 [1] molar mass = 46 g [1] [2]

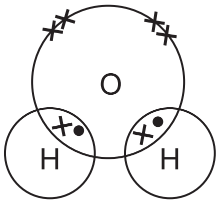
(ii) NO_2 [1]

(c) $3N_2O_4 + 2H_2O \rightarrow 4HNO_3 + 2NO$ [1]

(d) **(i)** moles of Mg = 0.25
moles of nitric acid = 0.50
volume of nitric acid = 250 cm^3 each error [–1] [3]

(ii) moles of magnesium nitrate formed 0.25
molar mass = 148
mass (g) = 0.25 × 148 = 37g each error [–1] [2] 11

					AVAILABLE MARKS
13	**(a)**	**(i)**	solution of known concentration	[1]	
		(ii)	e.g. phenolphthalein/methyl orange/colourless [1] to pink [1] or red/pink [1] to yellow/orange [1]	[3]	
	(b)	**(i)**	0.0018/1.8 × 10^{-3}	[1]	
		(ii)	0.0018/1.8 × 10^{-3}	[1]	
		(iii)	0.018/1.8 × 10^{-2}	[1]	
		(iv)	0.2	[1]	
		(v)	0.2 – 0.018 = 0.182	[1]	
		(vi)	0.091	[1]	
		(vii)	9.1 g	[1]	
		(viii)	91%	[1]	12
14	**(a)**		green/green-yellow/yellow-green — gas red-brown — liquid grey-black/black — solid [–1] for each error	[3]	
	(b)	**(i)**	white precipitate for solution of NaCl [1] dissolves in dilute ammonia [1] cream precipitate for solution of NaBr [1] dissolves in concentrated ammonia [1] yellow precipitate for solution of NaI [1] does not dissolve in concentrated ammonia [1]	[6]	
			Quality of written communication	[2]	
		(ii)	$Ag^+(aq) + I^-(aq) \rightarrow AgI(s)$	[2]	
	(c)	**(i)**	$NaCl + H_2SO_4 \rightarrow NaHSO_4 + HCl$	[2]	
		(ii)	$H_2SO_4 + \mathbf{8}\,H^+ + \mathbf{8e^-} \rightarrow H_2S + \mathbf{4}\,H_2O$ unbalanced [–1]	[2]	
		(iii)	$H_2SO_4 + 8\,H^+ + 8I^- \rightarrow 4\,I_2 + H_2S + 4\,H_2O$ unbalanced [–1]	[2]	
		(iv)	smell of rotten eggs	[1]	
		(v)	sulfur dioxide, sulfur	[2]	
		(vi)	the outer electrons in the iodide ions are further from the nucleus/ more shielded [1] iodide ions lose electrons more easily (than chloride ions) [1]	[2]	
	(d)	**(i)**	$3Cl_2 + 6NaOH \rightarrow 5NaCl + NaClO_3 + 3H_2O$	[2]	
		(ii)	disproportionation	[1]	27

		AVAILABLE MARKS
15 **(a)** 	[2]	
(b) **(i)** 104–105°	[1]	
(ii) V-shaped/angular/bent [1] repulsion [1] between lone (electron) pairs and bond pairs [1]	[3]	
(iii) no lone pairs in methane/only bond pairs in methane [1]	[1]	
(c) water molecules held together by hydrogen bonding [1] which is stronger than the intermolecular/polar forces between hydrogen sulfide molecules [1]	[2]	9
	Section B	**80**
	Total	**100**

Rewarding Learning

Centre Number

71	

Candidate Number

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
January 2013

Chemistry

Assessment Unit AS 1

assessing

Basic Concepts in Physical and Inorganic Chemistry

[AC112]

AC112

THURSDAY 10 JANUARY, MORNING

TIME

1 hour 30 minutes.

INSTRUCTIONS TO CANDIDATES

Write your Centre Number and Candidate Number in the spaces provided at the top of this page.
Answer **all sixteen** questions.
Answer **all ten** questions in **Section A**. Record your answers by marking the appropriate letter on the answer sheet provided. Use only the spaces numbered 1 to 10. Keep in sequence when answering.
Answer **all six** questions in **Section B**. Write your answers in the spaces provided in this question paper.

INFORMATION FOR CANDIDATES

The total mark for this paper is 100.
Quality of written communication will be assessed in question **11(b)**.
In Section A all questions carry equal marks, i.e. **two** marks for each question.
In Section B the figures in brackets printed down the right-hand side of pages indicate the marks awarded to each question or part question.
A Periodic Table of Elements (including some data) is provided.

For Examiner's use only	
Question Number	Marks
Section A	
1–10	
Section B	
11	
12	
13	
14	
15	
16	
Total Marks	

Section A

For each of the following questions only **one** of the lettered responses (A–D) is correct.

Select the correct response in each case and mark its code letter by connecting the dots as illustrated on the answer sheet.

1 Which one of the following is the electronic configuration for the Fe^{2+} ion?

A $1s^2 2s^2 2p^6 3s^2 3p^6 3d^4 4s^2$

B $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$

C $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^0$

D $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2$

2 Which one of the following lists the first four ionisation energies of a Group II element?

A 584, 1823, 2751, 11584

B 744, 1457, 7739, 10547

C 793, 1583, 3238, 4362

D 1018, 1909, 2918, 4963

3 Which one of the following is the mass of zinc chloride produced when 8.1 g of zinc oxide, ZnO, is added to 150.0 cm^3 of 1.0 mol dm^{-3} hydrochloric acid?

A 10.2 g

B 13.6 g

C 20.4 g

D 40.8 g

4 Which one of the following substances has coordinate bonding in its structure?

A Ammonia

B Ammonium chloride

C Carbon dioxide

D Water

5 Which one of the following, in the liquid state, has van der Waals' forces and permanent dipole attractions but **not** hydrogen bonds between the molecules?

A CH_4

B CO

C H_2O

D O_2

6 Chlorine does **not** undergo disproportionation when reacted with

A cold dilute sodium hydroxide solution.

B hot concentrated sodium hydroxide solution.

C potassium bromide solution.

D water.

7 Which one of the following solids will react with concentrated sulfuric acid to give hydrogen sulfide?

A Calcium bromide

B Magnesium iodide

C Potassium chloride

D Sodium fluoride

8 For which one of the following titrations would phenolphthalein be a suitable indicator?

A Ethanoic acid and sodium carbonate

B Ethanoic acid and sodium hydroxide

C Hydrochloric acid and aqueous ammonia

D Hydrochloric acid and sodium carbonate

9 Which one of the following increases as Group VII is descended?

A Atomic radius

B Electronegativity

C First ionisation energy

D Reactivity

10 Which one of the following is the bond angle in a water molecule?

A 104.5°

B 107.0°

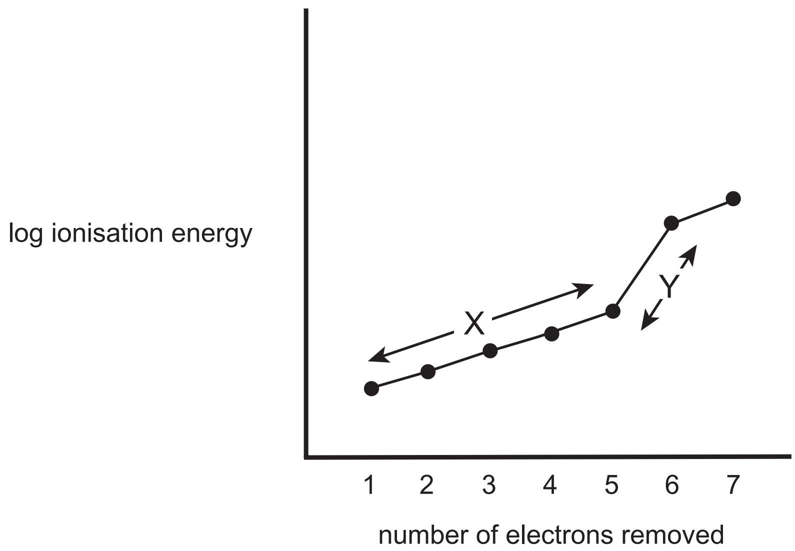
C 109.5°

D 112.0°

Section B

Answer all **six** questions in this section.

11 The graph below shows the successive ionisation energies of an element when all its electrons are removed.



(a) Name the element that gives rise to this graph.

____ [1]

(b) (i) Explain why the ionisation energies increase in section **X**.

____

____

____ [2]

(ii) Explain the large difference in ionisation energies in section **Y**.

____

____

____

____ [2]

Quality of written communication [2]

Examiner Only	
Marks	Remark

12 Caesium is a very reactive metal that has sky blue lines in the visible region of its emission spectrum (from the Latin *caesius* meaning "sky blue").

Examiner Only	
Marks	Remark

(a) (i) How are the lines in the emission spectrum of caesium formed?

____

____

____ [3]

(ii) Why do the lines in the emission spectrum of caesium converge?

____

____ [1]

(b) Caesium has one of the lowest first ionisation energies of all the elements in the Periodic Table.

(i) Write an equation, including state symbols, for the first ionisation of caesium.

____ [2]

(ii) The frequency at the convergence limit of caesium is 9.41×10^{14} Hz. Calculate the first ionisation energy of caesium in $kJ\,mol^{-1}$.

____

____

____

____ [3]

(iii) Give **two** reasons why the first ionisation energy of caesium is low.

____

____

____ [2]

(c) Caesium is so reactive that it will react with gold to form caesium auride, CsAu. The gold can be obtained from caesium auride by reacting it with water. Caesium hydroxide and hydrogen are the other products. Write an equation for the reaction of caesium auride with water.

____ [2]

Examiner Only	
Marks	Remark

Examiner Only	
Marks	Remark

13 Fluorine is the most electronegative element in the Periodic Table. Although the element is extremely reactive, fluoride ions can be safely added to water supplies and toothpastes.

(a) What is the meaning of the term **electronegativity**?

____

____ [2]

(b) Fluorine reacts with boron to form boron trifluoride.

(i) Draw a dot and cross diagram, using outer shell electrons only, to show the bonding in boron trifluoride.

[2]

(ii) State and explain the shape of a boron trifluoride molecule.

Shape: _______________________________________

Explanation: __________________________________

____

____ [3]

(iii) State the **octet rule**.

____

____ [2]

(iv) Explain whether or not the elements in boron trifluoride obey the octet rule.

____

____

____ [2]

Examiner Only	
Marks	Remark

(c) Sodium fluoride is added to toothpaste to strengthen tooth enamel.

(i) The data on a 50 g tube of toothpaste states that it contains "1450 ppm fluoride"; ppm means "parts per million" i.e. there would be 1450 g of fluoride ions in 10^6 (1,000,000) g of toothpaste. Use the following headings to work out the concentration of sodium fluoride in the toothpaste. The density of the toothpaste is 1.6 g cm^{-3}.

Mass of fluoride ions in the 50 g tube

____

Number of moles of fluoride ions in the 50 g tube

____

Number of moles of sodium fluoride in the 50 g tube

____

Volume of toothpaste in the 50 g tube

____

Concentration of sodium fluoride in the toothpaste with units

____

____ [6]

(ii) Sodium fluoride is also added to some public water supplies. Why might some people be opposed to this?

____

____ [1]

(iii) Giving practical details, describe how you could prove that a sample of solid sodium fluoride contains sodium ions.

____

____

____

____ [4]

14 Halide ions can be displaced from aqueous solutions of their salts using a more reactive halogen.

Examiner Only	
Marks	Remark

(a) A student bubbled excess chlorine into sodium bromide solution and the following reaction took place.

$$Cl_2(g) + 2NaBr(aq) \rightarrow 2NaCl(aq) + Br_2(aq)$$

(i) What change would be observed in the solution?

____ [2]

(ii) With reference to oxidation numbers explain why this is a redox reaction.

____

____

____ [3]

(iii) Describe how you could prove that there were no bromide ions remaining.

____

____

____

____ [4]

(b) Hydrogen halides are gases which are very soluble in water forming acidic solutions.

(i) State and suggest an explanation for the strength of hydrofluoric acid relative to the other hydrogen halides.

____

____

____ [2]

(ii) Silicon dioxide is used to make glass. Hydrofluoric acid has to be stored in plastic containers as it reacts with the "silicon dioxide" in glass to produce silicon tetrafluoride and water. Write the equation for this reaction.

____ [2]

15 Magnesium is an s-block metal which exists as the three isotopes, ^{24}Mg, ^{25}Mg and ^{26}Mg in a ratio of 8:1:1. It reacts with oxygen to form magnesium oxide which has a wide variety of uses including cement manufacture and heartburn medication.

(a) (i) Why is magnesium in the s-block of the Periodic Table?

____ [1]

(ii) Explain the meaning of the term **isotope**.

____

____ [2]

(iii) Calculate the relative atomic mass of magnesium to one decimal place.

____

____

____ [3]

(b) Magnesium oxide can be formed by the combustion of magnesium metal in oxygen.

(i) Draw a dot and cross diagram, using outer shell electrons only, to show how magnesium oxide is formed from a magnesium atom and an oxygen atom.

[3]

(ii) What type of bonding exists in magnesium oxide?

____ [1]

(iii) State **two** physical properties of magnesium oxide.

____

____ [2]

Examiner Only	
Marks	Remark

Examiner Only

Marks	Remark

(c) The amount of magnesium oxide in heartburn tablets can be determined by adding a known excess of hydrochloric acid to the tablets.

$$2HCl(aq) + MgO(s) \rightarrow MgCl_2(aq) + H_2O(l)$$

The amount of unreacted hydrochloric acid is determined by titrating it against sodium hydroxide.

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

(i) What is this method of titration called?

____ [1]

(ii) A student added 80 cm^3 of 2.0 $mol\,dm^{-3}$ hydrochloric acid to five crushed heartburn tablets which contained magnesium oxide. The unreacted acid required 25 cm^3 of 2.0 $mol\,dm^{-3}$ sodium hydroxide for complete neutralisation. Use the headings below to calculate the mass, in milligrams, of magnesium oxide in each tablet.

Number of moles of hydrochloric acid added to the tablets

____

Number of moles of unreacted hydrochloric acid

____

Number of moles of hydrochloric acid which reacted with the magnesium oxide

____

Number of moles of magnesium oxide present in five tablets

____

Mass of magnesium oxide per tablet in milligrams

____

____ [6]

Examiner Only				

16 The table below gives some information about three solids, aluminium, ice and diamond.

	Aluminium	**Ice**	**Diamond**
Density ($g\,cm^{-3}$)	2.70	0.92	3.52
Electrical conductivity	High	Low	Low
Melting point (°C)	660	0	3550

Use your knowledge of bonding and intermolecular forces to answer the following questions.

(a) Why is the density of ice lower than the density of water?

____

____

____ [2]

(b) Explain the high electrical conductivity of aluminium.

____

____

____ [2]

(c) Why does diamond have a high melting point?

____

____

____ [2]

THIS IS THE END OF THE QUESTION PAPER

113554

Rewarding Learning

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
January 2013

Chemistry

Assessment Unit AS 1

assessing

Basic Concepts in Physical and Inorganic Chemistry

[AC112]

THURSDAY 10 JANUARY, MORNING

MARK SCHEME

Section A

		AVAILABLE MARKS
1	C	
2	B	
3	A	
4	B	
5	B	
6	C	
7	B	
8	B	
9	A	
10	A	
[2] for each correct answer	[20]	20
	Section A	**20**

Section B

AVAILABLE MARKS

11 (a) Nitrogen [1]

(b) (i) **X**: proton to electron ratio increasing or effective nuclear charge increasing [1]; as each electron is removed the remaining electrons are held more tightly [1] [2]

(ii) **Y**: electron being removed from a full shell [1] closer to nucleus or less/no shielding represents the change in ionisation energy [1] [2]

Quality of written communication [2] 7

12 (a) (i) Electrons in high energy level [1]
drop back down [1]
emit energy/light [1] [3]

(ii) Energy levels converge [1]

(b) (i) $Cs(g) \rightarrow Cs^{+}(g) + e^{-}$ [2]

(ii) $E = (6.63 \times 10^{-34})(9.41 \times 10^{14}) = 6.239 \times 10^{-19}$
$IE = (6.239 \times 10^{-19})(6.02 \times 10^{23}) \div 1000 = 375.6\ kJ\,mol^{-1}$ [3]

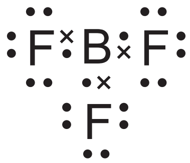
(iii) The electron being removed is far from the nucleus [1]
and experiences a lot of shielding [1] [2]

(c) $2CsAu + 2H_2O \rightarrow 2CsOH + 2Au + H_2$ [2] 13

AVAILABLE MARKS

13 **(a)** The extent to which an atom attracts the (bonding) electrons [1] in a covalent bond [1] [2]

(b) **(i)**



[2]

(ii) Trigonal planar [1]
Bonding electrons [1]
repel to get as far apart as possible [1] [3]

(iii) Atoms bond in order to get eight electrons [1] in the outer shell [1] [2]

(iv) Boron does not obey the octet rule in BF_3 as it has only six electrons [1] in its outer shell;
fluorine does obey [1] [2]

(c) **(i)** $1450/10^6 \times 50 = 0.0725$
$0.0725/19 = 3.816 \times 10^{-3}$
3.816×10^{-3}
$50/1.6 = 31.25\ \text{cm}^3$ **or** $0.03125\ \text{dm}^3$
$3.816 \times 10^{-3} \div 31.25 \times 1000 = 0.122\ \text{mol dm}^{-3}$ **or** $0.000122\ \text{mol cm}^{-3}$
[6]

(ii) Freedom of choice [1]

(iii) Nichrome wire [1]
concentrated hydrochloric acid [1]
place sample in blue Bunsen flame [1]
orange/yellow [1] [4] 22

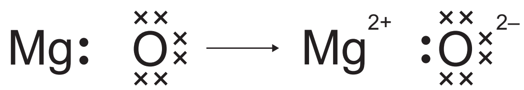
14 **(a)** **(i)** Colourless [1] to orange/yellow/brown [1] [2]

(ii) Cl 0 to –1 [1]
Br –1 to 0 [1]
Br oxidised and Cl reduced [1] [3]

(iii) Silver nitrate solution [1]
(white) precipitate [1]
add dilute ammonia [1]
no precipitate remaining [1] [4]

(b) **(i)** HF is weakest acid [1]
because the HF bond is strongest [1] [2]

(ii) $4HF + SiO_2 \rightarrow SiF_4 + 2H_2O$ [2] 13

					AVAILABLE MARKS
15	**(a)**	**(i)**	Outermost electrons in an s orbital	[1]	
		(ii)	Atoms having the same atomic number [1] different mass numbers [1]	[2]	
		(iii)	RAM = (24 × 0.8) + (25 × 0.1) + (26 × 0.1) = 24.3	[3]	
	(b)	**(i)**		[3]	
		(ii)	Ionic	[1]	
		(iii)	Conducts electricity when molten/aqueous [1] high melting point [1]	[2]	
	(c)	**(i)**	Back [1] titration	[1]	
		(ii)	nHCl = (80 × 2)/1000 = 0.16 nNaOH = (25 × 2)/1000 = 0.05 = moles of unreacted HCl nHCl reacting with MgO = 0.16 – 0.05 = 0.11 nMgO = 0.11 ÷ 2 = 0.055 mass MgO = (0.055/5) × (24 + 16) = 0.44 g = 440 mg	[6]	19
16	**(a)**		Hydrogen bonds in ice are fixed [1] holding water molecules further apart/leading to open structure [1]	[2]	
	(b)		(Three/high number of) delocalised electrons (per atom) [1] move and carry charge [1]	[2]	
	(c)		A lot of heat/energy required [1] to break strong covalent bonds/to break strong bonds in giant covalent structure [1]	[2]	6
				Section B	**80**
				Total	**100**

Rewarding Learning

Centre Number

71	

Candidate Number

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
2013

Chemistry

Assessment Unit AS 1

assessing

Basic Concepts in Physical and Inorganic Chemistry

[AC112]

AC112

WEDNESDAY 12 JUNE, AFTERNOON

TIME

1 hour 30 minutes.

INSTRUCTIONS TO CANDIDATES

Write your Centre Number and Candidate Number in the spaces provided at the top of this page.
Answer **all sixteen** questions.
Answer **all ten** questions in **Section A**. Record your answers by marking the appropriate letter on the answer sheet provided. Use only the spaces numbered 1 to 10. Keep in sequence when answering.
Answer **all six** questions in **Section B**. Write your answers in the spaces provided in this question paper.

INFORMATION FOR CANDIDATES

The total mark for this paper is 100.
Quality of written communication will be assessed in Question **16(b)(iii)**.
In Section A all questions carry equal marks, i.e. **two** marks for each question.
In Section B the figures in brackets printed down the right-hand side of pages indicate the marks awarded to each question or part question.
A Periodic Table of the Elements, containing some data, is included in this question paper.

For Examiner's use only	
Question Number	**Marks**
Section A	
1–10	
Section B	
11	
12	
13	
14	
15	
16	
Total Marks	

8181.04**R**

Section A

For each of the following questions only **one** of the lettered responses (A–D) is correct.

Select the correct response in each case and mark its code letter by connecting the dots as illustrated on the answer sheet.

1 Which one of the following is the formula for nitrogen(I) oxide?

A NO

B NO_2

C N_2O

D N_2O_4

2 Which one of the following is the number of atoms present in 0.25 moles of $C_{12}H_{22}O_{11}$?

A 6.8×10^{24}

B 1.4×10^{25}

C 2.7×10^{25}

D 1.1×10^{26}

3 Which one of the following is a molecular covalent substance?

A CaO

B CO

C Cr_2O_3

D CuO

4 A caesium atom differs from a caesium ion because the atom has a greater

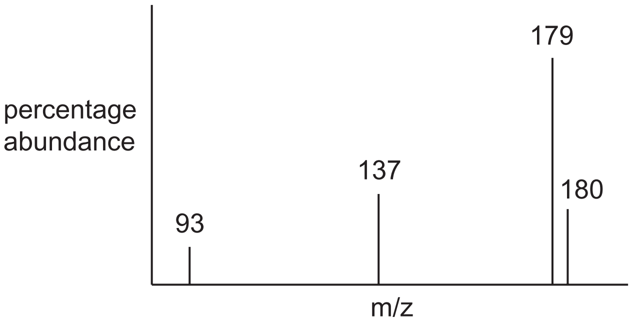
A atomic number.

B mass number.

C number of electrons.

D number of protons.

5 Part of the mass spectrum for aspirin is shown below. Which one of the following numbers is the molecular ion peak?



A 93

B 137

C 179

D 180

6 In which one of the following liquids are the van der Waals forces greatest?

A Argon

B Krypton

C Neon

D Xenon

7 Prozac tablets contain 20 mg of fluoxetine ($C_{17}H_{18}F_3NO$) in each tablet. The number of moles of fluoxetine in each tablet is

A 6.47×10^{-5}

B 1.39×10^{-4}

C 6.47×10^{-2}

D 1.39×10^{-1}

8 Which one of the following does **not** have a total of 14 electrons?

A CO

B Li_2O

C N_2

D S^{2-}

9 Successive ionisation energies for elements X and Y are shown below.

Ionisation energy ($kJ\,mol^{-1}$)	**1st**	**2nd**	**3rd**	**4th**	**5th**	**6th**	**7th**	**8th**
X	578	1817	2745	11 577	14 842	18 379	23 326	27 465
Y	1314	3388	5301	7469	10 990	13 327	71 330	84 078

Which one of the following is the formula for a compound of X and Y?

A XY_2

B X_2Y

C X_2Y_3

D X_3Y_2

10 Hexan-1-ol can be converted to hex-1-ene as follows:

$$C_6H_{13}OH \rightarrow C_6H_{12} + H_2O$$

40.0 g of hexan-1-ol produced 24.7 g of hex-1-ene. Which one of the following is the percentage yield?

A 24.7%

B 50.8%

C 72.0%

D 75.0%

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(Questions continue overleaf)

Section B

Answer **all six** questions in this section.

11 (a) Atoms consist of protons, neutrons and electrons.

(i) Complete the table below giving the properties of a proton, a neutron and an electron.

	Relative mass	**Relative charge**
Proton		
Neutron		
Electron		

[3]

(ii) Element 116, ununhexium, was added to the Periodic Table in June 2011. Complete the table below.

Atomic number	116
Mass number	
Number of protons	
Number of neutrons	177
Number of electrons	

[3]

(b) Iron is the sixth most abundant element in the Universe. It has four isotopes as shown in the table.

Isotope	^{54}Fe	^{56}Fe	^{57}Fe	^{58}Fe
Percentage abundance	5.84	91.76	2.12	0.28

(i) Explain what is meant by the term **isotope**.

____

____ [2]

(ii) Use the table to calculate the relative atomic mass of iron to **two** decimal places.

____

____

____ [2]

Examiner Only	
Marks	Remark

12 The emission spectrum for atomic hydrogen has been used to provide evidence for discrete electron energy levels in atoms.

(a) Complete the diagram to show the electron transitions associated with the first **two** lines of the hydrogen emission spectrum in the visible region.

n = 5 ______________________________

n = 4 ______________________________

n = 3 ______________________________

n = 2 ______________________________

n = 1 ______________________________ [2]

(b) The convergence limit of the hydrogen spectrum in the ultraviolet region is at 3.28×10^{15} Hz. Calculate the ionisation energy of hydrogen in $kJ\,mol^{-1}$.

____

____

______________________________ [3]

(c) The emission spectra of elements give rise to characteristic flame colours. Complete the table below.

Flame colour	Formula of metal ion
Blue-green	
Crimson	
Green	

[3]

Examiner Only	
Marks	Remark

13 Wood vinegar, which contains ethanoic acid, is formed when wood is heated. The percentage by mass of ethanoic acid in wood vinegar can be found by titration with standard sodium hydroxide solution.

Examiner Only	
Marks	Remark

(a) **(i)** What is meant by the term **standard solution**?

____ [1]

(ii) Write the equation for the reaction between ethanoic acid and sodium hydroxide.

____ [1]

(b) 25.0 cm^3 of wood vinegar were diluted to 250 cm^3 in a volumetric flask. 25.0 cm^3 of the diluted wood vinegar required 30.3 cm^3 of 0.1 $mol\,dm^{-3}$ sodium hydroxide solution for neutralisation.

(i) How many moles of sodium hydroxide were required?

____ [1]

(ii) How many moles of ethanoic acid were present in the 25.0 cm^3 of diluted wood vinegar?

____ [1]

(iii) How many moles of ethanoic acid were present in 25.0 cm^3 of undiluted wood vinegar?

____ [1]

(iv) What was the mass of ethanoic acid in the 25.0 cm^3 of undiluted wood vinegar?

____ [1]

(v) What was the percentage of ethanoic acid by mass in the wood vinegar? Assume that the density of wood vinegar is 1.02 $g\,cm^{-3}$.

____

____ [1]

(c) Suggest a suitable indicator for the titration and state the colour change at the end point.

Indicator: ______________________________

Colour change: from ______________________________

to ______________________________ [3]

Examiner Only	
Marks	Remark

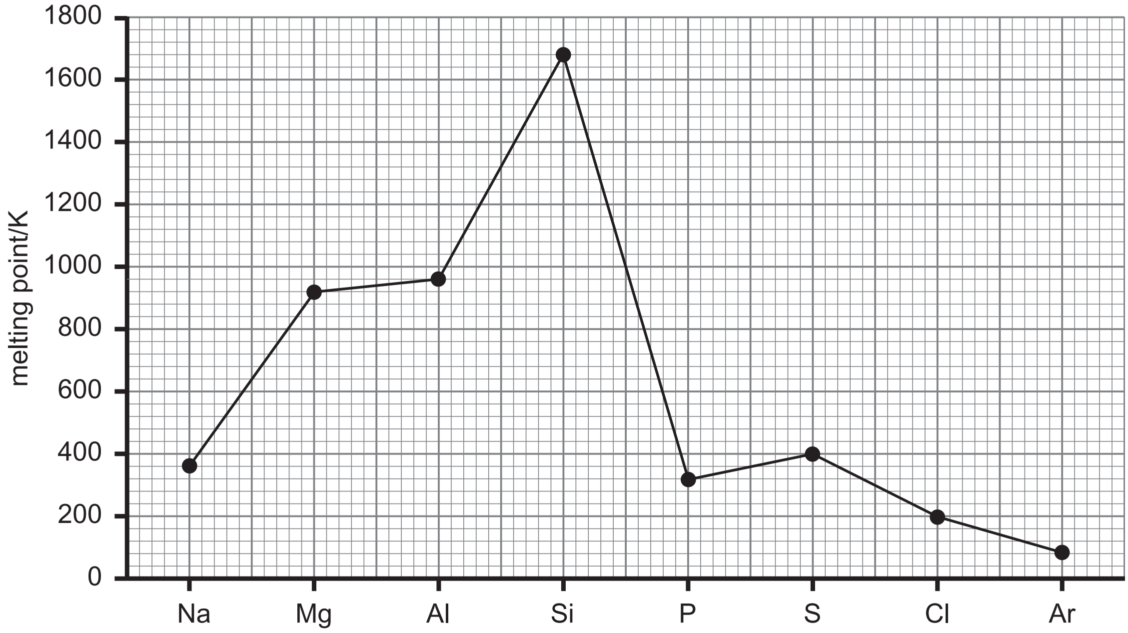
14 The Third Period from sodium to argon can be used to illustrate trends in the Periodic Table.

(a) In which block of the Periodic Table is argon found? Explain your answer.

____

____ [2]

(b) The graph below shows the melting points of the elements in the Third Period.



(i) Explain the rise in melting point from sodium to magnesium.

____

____

____ [2]

Examiner Only	
Marks	Remark

Examiner Only

Marks	Remark

(ii) Explain why silicon has the highest melting point.

[2]

(iii) Explain why the melting point of sulfur, S, is higher than phosphorus, P.

[2]

(c) State and explain the trend in atomic radius across the Third Period.

[3]

Examiner Only	
Marks	Remark

15 Aluminium chloride exists as the molecule $AlCl_3$ in the vapour state. This molecule contains covalent bonds and does not obey the octet rule.

(a) **(i)** Explain what is meant by the term **covalent bond**.

____

____ [2]

(ii) Explain what is meant by the term **octet rule**.

____

____ [2]

(b) Aluminium chloride reacts with chloride ions as follows:

$$AlCl_3 + Cl^- \rightarrow AlCl_4^-$$

(i) Draw dot and cross diagrams, using outer electrons only, to show the bonding in $AlCl_3$ and $AlCl_4^-$.

[4]

(ii) What type of bond is formed between $AlCl_3$ and the Cl^- ion?

____ [1]

(iii) Draw and name the shapes of $AlCl_3$ and $AlCl_4^-$.

[4]

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(Questions continue overleaf)

16 The halogens form Group VII of the Periodic Table.

Examiner Only

Marks | Remark

(a) The table below gives some of the physical properties of the halogens.

Element	**Atomic radius (nm)**	**Boiling point (°C)**	**Electronegativity value**	**First ionisation energy ($kJ\,mol^{-1}$)**
Fluorine	0.133	$^{-}187$	4.0	1618
Chlorine	0.181	$^{-}35$	3.0	1256
Bromine	0.196	59	2.8	1143
Iodine	0.219	183	2.0	1009

(i) Explain why the atomic radii of the halogens increase as the Group is descended.

____

____ [1]

(ii) Explain the trend in the boiling points of the halogens.

____

____ [2]

(iii) Explain what is meant by the term **electronegativity**.

____

____ [1]

(iv) Explain the trend in electronegativity values of the halogens.

____

____

____ [2]

Examiner Only

Marks	Remark

(v) Write an equation, including state symbols, for the first ionisation energy of fluorine.

____ [2]

(vi) Explain the trend in the first ionisation energy of the halogens.

____

____

____ [2]

(b) Chlorine is used to sterilise water.

(i) Write an equation for the reaction of chlorine with water.

____ [1]

(ii) Using changes in oxidation number explain why this is considered to be a disproportionation reaction.

____

____

____ [3]

(iii) Ultraviolet light does not react with water and is equally effective as chlorine at sterilising water. Suggest the advantages and disadvantages of storing and using chlorine to sterilise water.

____

____

____

____

____

____ [3]

Quality of written communication [2]

(c) Iodide ions react with a variety of reagents. For each of the following state what you would observe and write an equation for the reaction.

(i) Chlorine gas with aqueous iodide ions.

Observation ____

Equation ____ [3]

(ii) A solution containing *excess* Fe^{3+} ions with aqueous iodide ions.

Observation ____

Equation ____ [3]

(iii) Silver nitrate solution with aqueous iodide ions.

Observation ____

Equation ____ [3]

Examiner Only	
Marks	Remark

THIS IS THE END OF THE QUESTION PAPER

113078

Rewarding Learning

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
2013

Chemistry

Assessment Unit AS 1

assessing

Basic Concepts in Physical and Inorganic Chemistry

[AC112]

WEDNESDAY 12 JUNE, AFTERNOON

MARK SCHEME

General Marking Instructions

Introduction
Mark schemes are published to assist teachers and students in their preparation for examinations. Through the mark schemes teachers and students will be able to see what examiners are looking for in response to questions and exactly where the marks have been awarded. The publishing of the mark schemes may help to show that examiners are not concerned about finding out what a student does not know but rather with rewarding students for what they do know.

The Purpose of Mark Schemes
Examination papers are set and revised by teams of examiners and revisers appointed by the Council. The teams of examiners and revisers include experienced teachers who are familiar with the level and standards expected of students in schools and colleges.

The job of the examiners is to set the questions and the mark schemes; and the job of the revisers is to review the questions and mark schemes commenting on a large range of issues about which they must be satisfied before the question papers and mark schemes are finalised.

The questions and the mark schemes are developed in association with each other so that the issues of differentiation and positive achievement can be addressed right from the start. Mark schemes, therefore, are regarded as part of an integral process which begins with the setting of questions and ends with the marking of the examination.

The main purpose of the mark scheme is to provide a uniform basis for the marking process so that all the markers are following exactly the same instructions and making the same judgements in so far as this is possible. Before marking begins a standardising meeting is held where all the markers are briefed using the mark scheme and samples of the students' work in the form of scripts. Consideration is also given at this stage to any comments on the operational papers received from teachers and their organisations. During this meeting, and up to and including the end of the marking, there is provision for amendments to be made to the mark scheme. What is published represents this final form of the mark scheme.

It is important to recognise that in some cases there may well be other correct responses which are equally acceptable to those published: the mark scheme can only cover those responses which emerged in the examination. There may also be instances where certain judgements may have to be left to the experience of the examiner, for example, where there is no absolute correct response – all teachers will be familiar with making such judgements.

Section A

		AVAILABLE MARKS
1	C	
2	A	
3	B	
4	C	
5	D	
6	D	
7	A	
8	D	
9	C	
10	D	
[2] for each correct answer	[20]	20
	Section A	**20**

Section B

AVAILABLE MARKS

11 **(a)** **(i)**

	Relative mass	Relative charge
Proton	1	1+
Neutron	1	0
Electron	$\frac{1}{1800} - \frac{1}{2000}$ negligible	1–

error [–1] [3]

(ii)

Atomic number	116
Mass number	**293**
Number of protons	**116**
Number of neutrons	177
Number of electrons	**116**

([1] each) [3]

(b) **(i)** Atoms with the same atomic number/number of protons [1]
but with a different mass number/number of neutrons [1] [2]

(ii) $((54 \times 5.84) + (56 \times 91.76) + (57 \times 2.12) + (58 \times 0.28))/100$
$= 55.91$
([–1] for each mistake, [–1] if not correct to 2 decimal places) [2] 10

12 **(a)** Arrow from n = 3 to n = 2 [1]
Arrow from n = 4 to n = 2 [1] [2]

(b) $E = hf = (6.63 \times 10^{-34}) \times (3.28 \times 10^{15}) = 2.175 \times 10^{-18}$

$(2.175 \times 10^{-18}) \times (6.02 \times 10^{23}) = 1\,309\,350\text{ J mol}^{-1}$
1309 (kJ mol^{-1})
([–1] for each mistake, [–1] if not kJ mol^{-1}) [3]

(c)

Flame colour	Metal ion
Blue-green	$\mathbf{Cu^{2+}}$
Crimson	$\mathbf{Li^{+}}$
Green	$\mathbf{Ba^{2+}}$

(Must be formula not name, [1] each) [3] 8

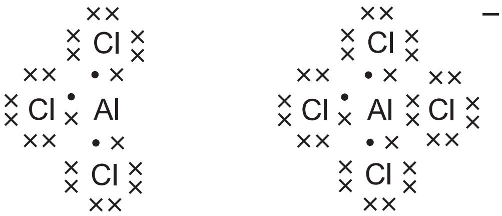
					AVAILABLE MARKS
13	**(a)**	**(i)**	A solution of (accurately) known concentration	[1]	
		(ii)	$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$	[1]	
	(b)	**(i)**	$(0.1 \times 30.3)/1000 = 3.03 \times 10^{-3}$	[1]	
		(ii)	3.03×10^{-3}	[1]	
		(iii)	3.03×10^{-2}	[1]	
		(iv)	$(3.03 \times 10^{-2}) \times 60 = 1.82\,g$ units needed	[1]	
		(v)	$\frac{1.82}{25 \times 1.02} \times 100 = 7.14(\%)$	[1]	
	(c)		Phenolphthalein [1] From colourless [1] to pink/red [1]	[3]	10
14	**(a)**		p(-block) [1] Outer electrons in the p-orbital [1]	[2]	
	(b)	**(i)**	Increasing number of valence/outer/free electrons [1] Greater attraction between these and the (fixed) cations [1]	[2]	
		(ii)	Strong covalent bonds [1] throughout a giant structure [1] (reference to ionic bonding [0])	[2]	
		(iii)	$S_8 \rightarrow P_4$ – More atoms/greater mass/more electrons [1] Greater van der Waals forces (between the molecules) [1]	[2]	
	(c)		Atomic radius decreases across the Period/from sodium to argon [1] (Outer) electrons are in the same energy level/shielding remains the same [1] Nuclear charge increases causing greater attraction between the nucleus and the (outer) electrons [1]	[3]	11

AVAILABLE MARKS

15 **(a)** **(i)** A shared pair of electrons (between two atoms) [1]
Each atom provides one electron [1] [2]

(ii) When forming a compound an atom tends to gain, lose or share electrons to achieve eight [1]
in its outer shell [1] [2]

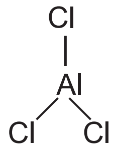
(b) **(i)**



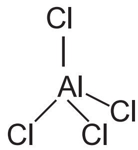
([2] each, [–1] for each mistake, [–1] if dot and cross not used) [4]

(ii) Dative/co-ordinate bond [1]

(iii) trigonal planar
(diagram [1], name [1])



tetrahedral
(diagram [1], name [1])



[4] 13

AVAILABLE MARKS

16 **(a) (i)** As the Group is descended there are more energy levels [1]

(ii) Going down the Group the molecules get heavier/there are more electrons [1]
causing van der Waals forces between the molecules [1] [2]

(iii) The extent to which an atom attracts the (bonding) electrons in a covalent bond [1]

(iv) Fluorine has the smallest radius [1]
greatest attraction between its nucleus and the bonding electrons [1] [2]

(v) $F(g) \rightarrow F^+(g) + e^-$
([1] for equation, [1] for state symbols) [2]

(vi) Going down the Group the outer electrons are further from the nucleus [1]
Increased shielding (from the inner electrons) [1] [2]

(b) (i) $Cl_2 + H_2O \rightarrow HOCl + HCl$ [1]

(ii) $\left.\begin{array}{l} Cl_2 = 0 \\ HOCl = +1 \\ HCl = -1 \end{array}\right\}$ oxidation numbers [2]
Cl is both oxidised, 0 → +1 and reduced, 0 → –1 [1] [3]

(iii) Disadvantages:
Storing large quantities of chlorine causes problems/
Chlorine poisonous/toxic/
Freedom of choice

Advantages:
Chlorine remains in the water after it leaves the treatment plant/
Chlorine gas can be compressed/chlorine is relatively cheap
To a maximum of [3] [3]

Quality of written communication [2]

(c) (i) Colourless solution [1] turns yellow/brown [1]
$Cl_2 + 2I^- \rightarrow 2Cl^- + I_2$ [1] [3]

(ii) Yellow/orange solution [1] turns brown/yellow [1]
or colourless → yellow/brown
not yellow → yellow
$2Fe^{3+} + 2I^- \rightarrow 2Fe^{2+} + I_2$ [1] [3]

(iii) Yellow [1] precipitate [1]
$Ag^+ + I^- \rightarrow AgI$ [1] [3] 28

Section B	**80**
Total	**100**

Rewarding Learning

Centre Number	
71	

Candidate Number

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
January 2014

Chemistry

Assessment Unit AS 1

assessing

Basic Concepts in Physical and Inorganic Chemistry

[AC112]

AC112

THURSDAY 9 JANUARY, MORNING

TIME

1 hour 30 minutes.

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Answer **all sixteen** questions.
Answer **all ten** questions in **Section A**. Record your answers by marking the appropriate letter on the answer sheet provided. Use only the spaces numbered 1 to 10. Keep in sequence when answering.
Answer **all six** questions in **Section B**. Write your answers in the spaces provided in this question paper.

INFORMATION FOR CANDIDATES

The total mark for this paper is 100.
Quality of written communication will be assessed in Question **12(d)(iv)**.
In Section A all questions carry equal marks, i.e. **two** marks for each question.
In Section B the figures in brackets printed down the right-hand side of pages indicate the marks awarded to each question or part question.
A Periodic Table of Elements, containing some data, is included in this question paper.

For Examiner's use only	
Question Number	**Marks**
Section A	
1–10	
Section B	
11	
12	
13	
14	
15	
16	
Total Marks	

Section A

For each of the following questions only **one** of the lettered responses (A–D) is correct.

Select the correct response in each case and mark its code letter by connecting the dots as illustrated on the answer sheet.

1 An element in the Periodic Table has the following successive ionisation energies ($kJ\ mol^{-1}$).

590 1145 4912 6474 8144 10496 12320

In which one of the following groups is this element found?

A Group I

B Group II

C Group III

D Group IV

2 Which one of the following is the oxidation number of hafnium in HfF_7^{3-}?

A -3

B $+3$

C -4

D $+4$

3 Boron consists of the isotopes $^{10}_{5}B$ and $^{11}_{5}B$. The relative atomic mass of the element is 10.80. Which one of the following is the approximate ratio of the number of lighter atoms to heavier atoms?

A 1:3

B 1:4

C 1:9

D 4:1

4 Which one of the following equations shows hydrogen peroxide, H_2O_2, behaving as a reducing agent?

A $2Fe^{2+} + H_2O_2 + 2H^+ \rightarrow 2Fe^{3+} + 2H_2O$

B $2I^- + H_2O_2 + 2H^+ \rightarrow I_2 + 2H_2O$

C $MnO_2 + 2H^+ + H_2O_2 \rightarrow Mn^{2+} + 2H_2O + O_2$

D $PbS + 4H_2O_2 \rightarrow PbSO_4 + 4H_2O$

5 The electronegativity values, not in order, for caesium, cobalt, fluorine and nitrogen are listed below. Which one of the following is the value for the cobalt atom?

A 0.70

B 1.80

C 3.00

D 4.00

6 Which one of the following molecules is linear?

A CH_3CH_3

B CO_2

C H_2O_2

D H_2Te

7 Which one of the following is the reason why water boils at 100 °C while the hydrides of the other Group VI elements boil below 0 °C?

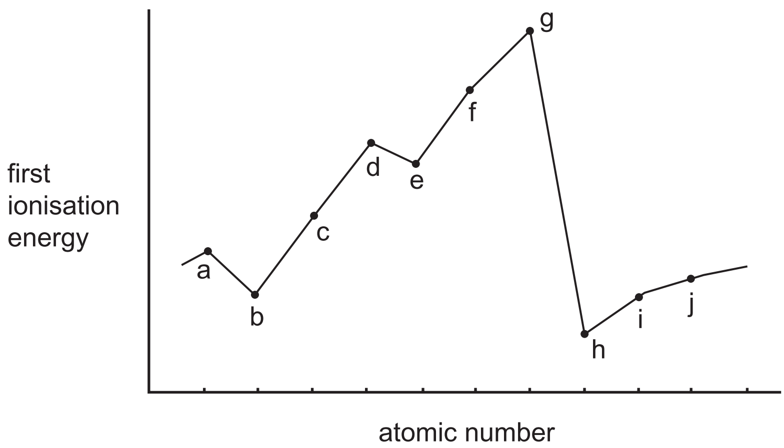
A Hydrogen bonding between water molecules

B Ionic bonding in water molecules

C The lower molar mass of water molecules

D The stability of the bonding in water molecules

8 The first ionisation energy is shown against increasing atomic number.



Which one of the following shows a Group I element together with a Group VII element?

	Group I	Group VII
A	b	f
B	b	g
C	h	f
D	h	g

9 Which one of the following properties is a characteristic of astatine?

A It has an electronegativity value greater than that of iodine.

B It is a solid at room temperature and pressure.

C It oxidises bromide ions to bromine.

D Its hydride exhibits more hydrogen bonding than hydrogen iodide.

10 3.12 g of MCl_2 were dissolved in water and made up to one litre of solution. 25.0 cm^3 of this solution reacts with 7.5 cm^3 of 0.100 M silver nitrate solution.

$$MCl_2(aq) \rightarrow M^{2+}(aq) + 2Cl^-(aq)$$
$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$

Which one of the following Group II elements is M?

A barium

B calcium

C magnesium

D strontium

Section B

Answer **all six** questions in this section.

11 Complete the following table about the silver halides.

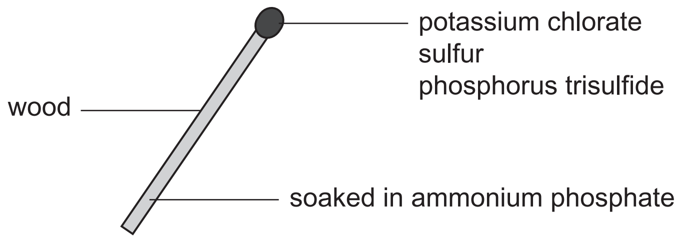
silver halide	**formula**	**colour**	**ionic/ covalent**	**soluble in dilute ammonia solution**	**soluble in concentrated ammonia solution**
silver fluoride	AgF	white	ionic	yes	yes
silver chloride					
silver bromide					
silver iodide					

[4]

Examiner Only	
Marks	Remark

12 The creation of the friction match took many years and involved a variety of chemicals based on phosphorus.

The modern match is shown below. The head is a mixture of potassium chlorate, sulfur and phosphorus trisulfide held together by glue. The wood is soaked in ammonium phosphate which acts as a fire retardant.



(a) Potassium chlorate reacts with the sulfur to form potassium chloride and sulfur dioxide as shown by the following equation.

$$2KClO_3 + 3S \rightarrow 2KCl + 3SO_2$$

(i) Deduce the oxidation number for each element in the reactants.

____ [1]

(ii) Deduce the oxidation number for each element in the products.

____ [1]

(iii) Explain, using these oxidation numbers, why this is a redox reaction.

____

____ [1]

(b) Potassium chlorate, $KClO_3$, is manufactured using the reaction between chlorine and potassium hydroxide.

(i) Write the equation for the reaction.

____ [2]

(ii) State the conditions under which the reaction is carried out.

____ [1]

Examiner Only	
Marks	Remark

(c) Phosphorus trisulfide is easily ignited. It provides the heat to initiate the reaction between potassium chlorate and sulfur.

(i) Phosphorus has an oxidation number of +3 in phosphorus trisulfide. State the formula of phosphorus trisulfide.

____ [1]

(ii) Suggest whether phosphorus trisulfide is ionic or covalent. Explain your reasons.

____

____ [1]

(iii) Name the **two** products formed when phosphorus trisulfide is completely burnt. No oxidation numbers are needed.

____

____ [2]

(d) Ammonium phosphate is an ionic compound consisting of ammonium and phosphate ions, PO_4^{3-}.

(i) Write the formula of the ammonium ion.

____ [1]

(ii) Name and draw the shape of the ammonium ion stating the interbond angle.

____ [3]

(iii) Write the formula of ammonium phosphate.

____ [1]

Examiner Only	
Marks	Remark

(iv) State and explain **three** physical properties you would expect ammonium phosphate to have.

____

____

____

____

____

____

____

____ [3]

Quality of written communication [2]

Examiner Only	
Marks	Remark

13 Francium is found in Group I of the Periodic Table and was discovered by Marguerite Perey in 1939 in the Curie Laboratory in France. It was isolated from uranium ore. Since then it has been synthesised by the nuclear reaction of oxygen atoms with gold atoms. It exists as 34 isotopes.

In the Periodic Table it has an atomic number of 87 and is given a relative atomic mass of 223.

(a) Francium is found in period 7 of the Periodic Table and is regarded as an s-block element.

Suggest the subshell occupied by the outermost electron in a francium atom.

____ [1]

(b) Francium was first synthesised according to the following equation.

$$^{197}Au + {}^{18}O \rightarrow {}^{210}Fr + 5n$$

The symbol n represents a neutron.

(i) What is the relative mass of a neutron?

____ [1]

(ii) Using the relative mass of the neutron from part **(i)** show, by calculation, that the equation is balanced according to mass.

____

____ [2]

(iii) Why are electrons not used when balancing the equation according to mass?

____

____ [1]

Examiner Only	
Marks	Remark

(c) Francium is one of the least electronegative elements in the Periodic Table.

(i) Explain the meaning of the term **electronegativity**.

____

____ [2]

(ii) State how electronegativity values change on going across a period.

____ [1]

(d) Francium has a melting point of 27 °C and would melt in the hand just as caesium does. It has the highest electrical conductivity of the alkali metals.

(i) Explain, in terms of metallic bonding, why francium has a low melting point.

____

____ [2]

(ii) Explain, in terms of metallic bonding, why francium has the highest electrical conductivity.

____

____ [2]

Examiner Only	
Marks	Remark

Examiner Only	
Marks	Remark

(e) Francium loses electrons when it reacts with chlorine and the chlorine gains these electrons.

(i) Write the equation for the loss of an electron from a francium atom.

____ [1]

(ii) Write the equation for the formation of chloride ions from a chlorine molecule.

____ [1]

(iii) Write the equation for the reaction of francium with chlorine.

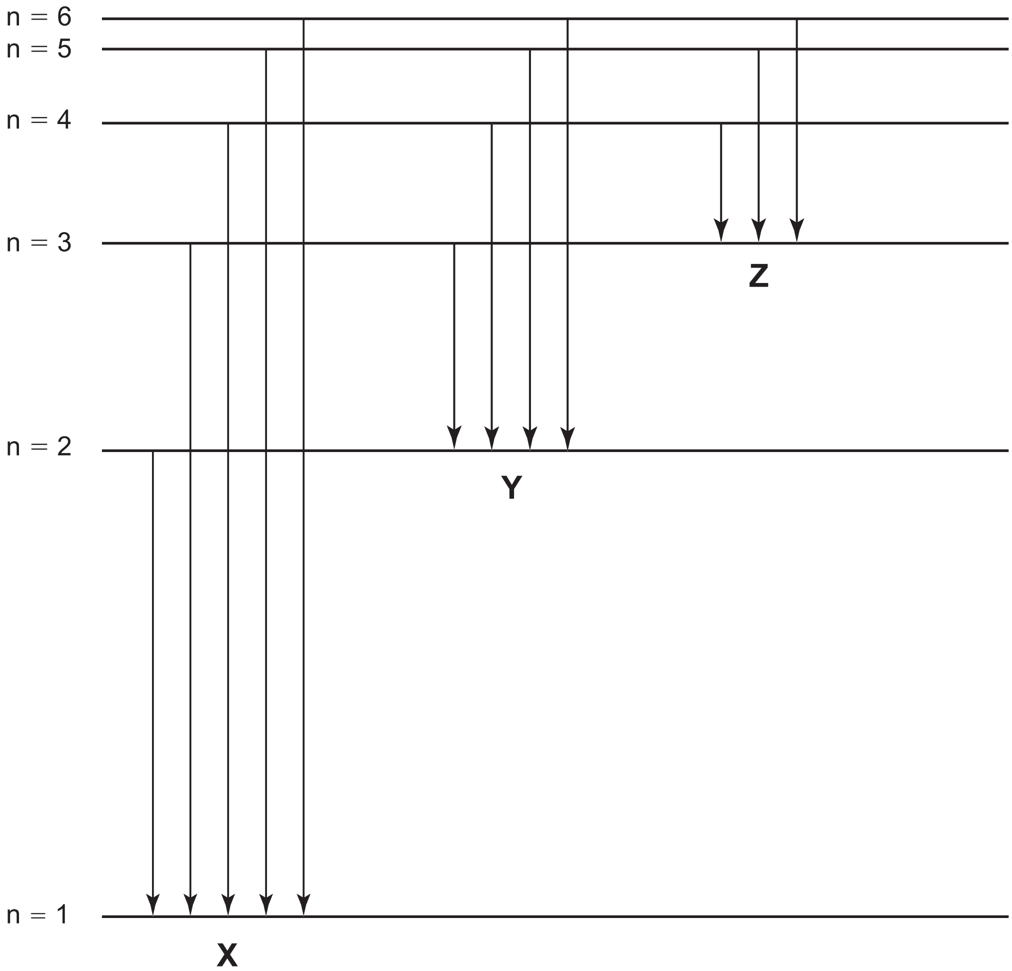
____ [1]

(iv) Francium chloride exists as a lattice structure similar to that of NaCl. Explain the term **lattice structure**.

____

____ [1]

14 The energy levels of a hydrogen atom are shown below and the arrows indicate the transition of electrons between successive energy levels.



The electromagnetic spectrum is shown below.

Radio waves	Microwaves	Infrared	Visible	Ultraviolet	X-rays	Gamma rays

Energy increases →

(a) Write the equation that relates energy to frequency, explain the meanings of the symbols used and state the units in which they are measured.

____

____

____

____ [3]

Examiner Only

Marks	Remark

Examiner Only	
Marks	Remark

(b) There are three series of lines. The first series **X** occurs in the ultraviolet region of the electromagnetic spectrum.

(i) In which part of the electromagnetic spectrum does the second series, **Y**, occur?

____ [1]

(ii) Suggest in which part of the electromagnetic spectrum the third series, **Z**, occurs.

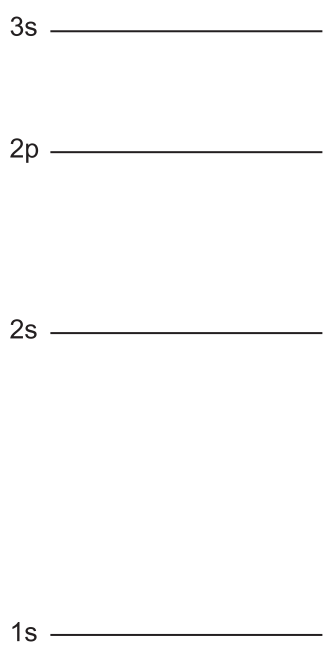
____ [1]

(iii) What happens to the atom when its electron passes from energy level n = 1 to an infinite energy level?

____

____ [1]

(c) **(i)** Use the energy level diagram below to show the ground state of a sodium atom. Use arrows to represent the sodium electrons.



[2]

(ii) What is the predominant colour in the emission spectrum of sodium?

____ [1]

Examiner Only	
Marks	Remark

15 Iron reacts with dilute hydrochloric acid to form iron(II) chloride, $FeCl_2$, and hydrogen. The solution deposits crystals of hydrated iron(II) chloride.

(a) Write the ionic equation, with state symbols, for the reaction of iron with hydrochloric acid.

____ [2]

(b) A solution of iron(II) ions is oxidised by chlorine water to form iron(III) ions.

(i) Write the ionic equation for the reaction.

____ [1]

(ii) Describe the colour of the solution after the reaction has taken place.

____ [1]

(iii) Explain whether iron(II) ions would react with bromine water.

____

____ [1]

(c) Iron(II) chloride is extremely soluble in water. 69 g of the anhydrous solid dissolve in 100 cm^3 of water at 20 °C. Assuming there is no volume change calculate the molarity of the resulting solution.

____

____

____ [3]

(d) 14.1 g of the hydrated iron(II) chloride crystals contain 6.5 g of water.
Use these figures to calculate the formula of the crystals.

Examiner Only	
Marks	Remark

mass of iron(II) chloride

____

moles of iron(II) chloride

____

moles of water

____

ratio of moles of water to moles of iron(II) chloride

____

formula of iron(II) chloride crystals

____ [5]

16 Sulfur forms the following fluorides:

sulfur difluoride	SF_2
sulfur tetrafluoride	SF_4
sulfur hexafluoride	SF_6

Sulfur hexafluoride is the best known and can be used as a safe electrical insulator. The other fluorides are toxic.

(a) Draw the dot and cross diagrams showing the outer electrons only for each of the fluorides.

[3]

(b) **(i)** State the **octet rule**.

____

____ [2]

(ii) Explain whether sulfur is obeying the octet rule in each fluoride.

____

____

____ [2]

(c) Sulfur difluoride has the same shape as a water molecule but the bond angle is 6° smaller. Draw and name the shape of sulfur difluoride, stating its bond angle.

____ [3]

Examiner Only	
Marks	Remark

Examiner Only	
Marks	Remark

(d) The sulfur hexafluoride molecule has an octahedral shape.

(i) State the bond angle(s) in the sulfur hexafluoride molecule.

____ [1]

(ii) Explain why sulfur hexafluoride has an octahedral shape.

____

____ [2]

(iii) Explain why sulfur hexafluoride is a non-polar molecule.

____

____ [1]

(e) Sulfur tetrafluoride has a boiling point of –38 °C whereas sulfur hexafluoride has a boiling point of –64 °C.

(i) Which compound has the higher boiling point?

____ [1]

(ii) Explain, in terms of mass, which compound has the greater van der Waals forces.

____

____ [1]

(iii) Explain, in terms of intermolecular forces, the difference in boiling points.

____

____ [2]

THIS IS THE END OF THE QUESTION PAPER

Rewarding Learning

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
January 2014

Chemistry

Assessment Unit AS 1

assessing

Basic Concepts in Physical and Inorganic Chemistry

[AC112]

THURSDAY 9 JANUARY, MORNING

MARK SCHEME

General Marking Instructions

Introduction
Mark schemes are published to assist teachers and students in their preparation for examinations. Through the mark schemes teachers and students will be able to see what the examiners are looking for in response to questions and exactly where the marks have been awarded. The publishing of the mark schemes may help to show that examiners are not concerned about finding out what a student does not know but rather with rewarding students for what they do know.

The purpose of mark schemes
Examination papers are set and revised by teams of examiners and revisers appointed by the Council. The teams of examiners and revisers include experienced teachers who are familiar with the level and standards expected of students in schools and colleges.

The job of the examiners is to set the questions and the mark schemes; and the job of the revisers is to review the questions and mark schemes commenting on a large range of issues about which they must be satisfied before the question papers and mark schemes are finalised.

The questions and the mark schemes are developed in association with each other so that the issues of differentiation and positive achievement can be addressed right from the start. Mark schemes, therefore, are regarded as part of an integral process which begins with the setting of questions and ends with the marking of the examination.

The main purpose of the mark scheme is to provide a uniform basis for the marking process so that all the markers are following exactly the same instructions and making the same judgements in so far as this is possible. Before marking begins a standardising meeting is held where all the markers are briefed using the mark scheme and samples of the students' work in the form of scripts. Consideration is also given at this stage to any comments on the operational papers received from teachers and their organisations. During this meeting, and up to and including the end of the marking, there is provision for amendments to be made to the mark scheme. What is published represents the final form of the mark scheme.

It is important to recognise that in some cases there may well be other correct responses which are equally acceptable to those published: the mark scheme can only cover those responses which emerged in the examination. There may also be instances where certain judgements may have to be left to the experience of the examiner, for example where there is no absolute correct response – all teachers will be familiar with making such judgements.

Section A

			AVAILABLE MARKS
1	B		
2	D		
3	B		
4	C		
5	B		
6	B		
7	A		
8	C		
9	B		
10	A		
[2] for each correct answer		[20]	20
		Section A	**20**

Section B

						AVAILABLE MARKS
11 Each mistake is [–1]						
AgCl	white	ionic	yes	yes		
AgBr	cream	ionic	no	yes		
AgI	yellow	ionic	no	no	[4]	4

12 **(a)** **(i)** K = +1; Cl = +5; O = –2; S = 0 [1]

(ii) K = +1; Cl = –1; O = –2; S = +4 [1]

(iii) Cl oxidation number goes down and S oxidation number goes up [1]

(b) **(i)** $6KOH + 3Cl_2 \rightarrow KClO_3 + 5KCl + 3H_2O$ [2]

(ii) hot and concentrated (potassium hydroxide solution) [1]

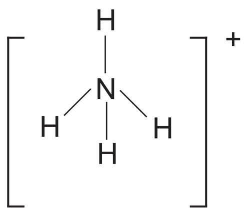
(c) **(i)** P_2S_3 [1]

(ii) covalent as both elements are non-metals/covalent as small difference in electronegativity between P and S [1]

(iii) phosphorus trioxide or pentoxide/phosphorus oxide [1]
sulfur oxide or sulfur dioxide [1] [2]

(d) **(i)** NH_4^+ [1]

(ii) tetrahedral [1]
109º/109.5° [1]



[1] [3]

(iii) $(NH_4)_3PO_4/PO_4(NH_4)_3$ [1]

(iv) melting point which is high because of attraction between ions [1]
boiling point which is high because of attraction between ions [1]
electrical conductivity which is high in solution or molten because ions can move [1]
solubility, ions are surrounded by H_2O molecules [1]
3 from 4 [3]

Quality of written communication [2] — 20

						AVAILABLE MARKS
13	**(a)**	$7s^1$			[1]	
	(b)	**(i)**	1		[1]	
		(ii)	LHS = 197 + 18 = 215 RHS = 210 + 5 × 1 = 215		[2]	
		(iii)	electrons have negligible mass/one two thousandth relative mass		[1]	
	(c)	**(i)**	the extent to which an atom attracts the bonding electrons in a covalent bond		[2]	
		(ii)	electronegativity increases across a period		[1]	
	(d)	**(i)**	sea of positive ions surrounded by electrons attraction between positive ions and electrons is weak	[1] [1]	[2]	
		(ii)	electrons move and conduct/carry electricity large radius of Fr means less attraction for electrons	[1] [1]	[2]	
	(e)	**(i)**	$Fr \rightarrow Fr^+ + e$		[1]	
		(ii)	$Cl_2 + 2e \rightarrow 2Cl^-$		[1]	
		(iii)	$2Fr + Cl_2 \rightarrow 2FrCl$		[1]	
		(iv)	regular arrangement of (francium and chloride) ions		[1]	16

AVAILABLE MARKS

14 **(a)** $E = hf$
E is energy, h is Planck's constant, f is frequency
E in joules, h in J s, f in Hz or s^{-1} [3]

(b) **(i)** visible [1]

(ii) infrared [1]

(iii) the atom has ionised/lost the electron [1]

(c) **(i)** ↑
⇅ ⇅ ⇅
⇅
⇅ [2]

(ii) yellow/orange [1] 9

15 **(a)** $Fe(s) + 2H^+(aq) \rightarrow Fe^{2+}(aq) + H_2(g)$ [2]

(b) **(i)** $2Fe^{2+} + Cl_2 \rightarrow 2Fe^{3+} + 2Cl^-$ [1]

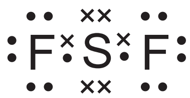
(ii) yellow/orange [1]

(iii) yes, bromine is a sufficiently powerful oxidising agent [1]

(c) $FeCl_2 = 56 + 2 \times 35.5 = 127$
69 g = 69/127 = 0.54 mol
= 5.4 M [3]

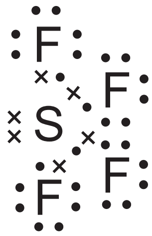
(d) mass of iron(II) chloride = 14.1 – 6.5 = 7.6
moles of iron(II) chloride = 7.6/127 = 0.0598
moles of water = 6.5/18 = 0.36
ratio of moles = 0.36:0.0598 = 6.02
$FeCl_2.6H_2O$ [5] 13

16 **(a)** SF_2



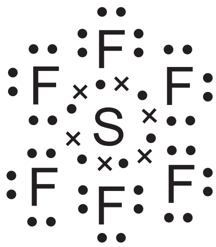
[1]

SF_4



[1]

SF_6

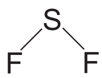


[1] [3]

(b) **(i)** when forming a compound an atom tends to gain, lose or share electrons in its outer shell to achieve 8 [2]

(ii) SF_2 yes; SF_4 no; SF_6 no, there are 10 and 12 electrons around S [2]

(c) shape



[1]

bent [1]
104.5° – 6° = 98.5° [1] [3]

(d) **(i)** 90° [1]

(ii) the six bonds repel each other [1]
as far apart from each other as possible [1] [2]

(iii) the molecule is symmetrical
or dipoles cancel [1]

(e) **(i)** sulfur tetrafluoride [1]

(ii) sulfur hexafluoride has the greatest mass hence the greatest van der Waals [1]

(iii) SF_4 is polar (SF_6 is non-polar) [1]
polar forces greater than van der Waals [1] [2]

	AVAILABLE MARKS
	18
Section B	**80**
Total	**100**

Rewarding Learning

ADVANCED SUBSIDIARY (AS)
General Certificate of Education
2014

Chemistry

Assessment Unit AS 1

assessing

Basic Concepts in Physical and Inorganic Chemistry

[AC112]

MONDAY 9 JUNE, AFTERNOON

Centre Number

71	

Candidate Number

AC112

TIME

1 hour 30 minutes.

INSTRUCTIONS TO CANDIDATES

Write your Centre Number and Candidate Number in the spaces provided at the top of this page.
Answer **all fifteen** questions.
Answer **all ten** questions in **Section A**. Record your answers by marking the appropriate letter on the answer sheet provided. Use only the spaces numbered 1 to 10. Keep in sequence when answering.
Answer **all five** questions in **Section B**. Write your answers in the spaces provided in this question paper.

INFORMATION FOR CANDIDATES

The total mark for this paper is 100.
Quality of written communication will be assessed in Question **12(b)**.
In Section A all questions carry equal marks, i.e. **two** marks for each question.
In Section B the figures in brackets printed down the right-hand side of pages indicate the marks awarded to each question or part question.
A Periodic Table of the Elements, containing some data, is included in this question paper.

For Examiner's use only	
Question Number	**Marks**
Section A	
1–10	
Section B	
11	
12	
13	
14	
15	
Total Marks	

Section A

For each of the following questions only **one** of the lettered responses (A–D) is correct.

Select the correct response in each case and mark its code letter by connecting the dots as illustrated on the answer sheet.

1 Which one of the following is **not** a redox reaction?

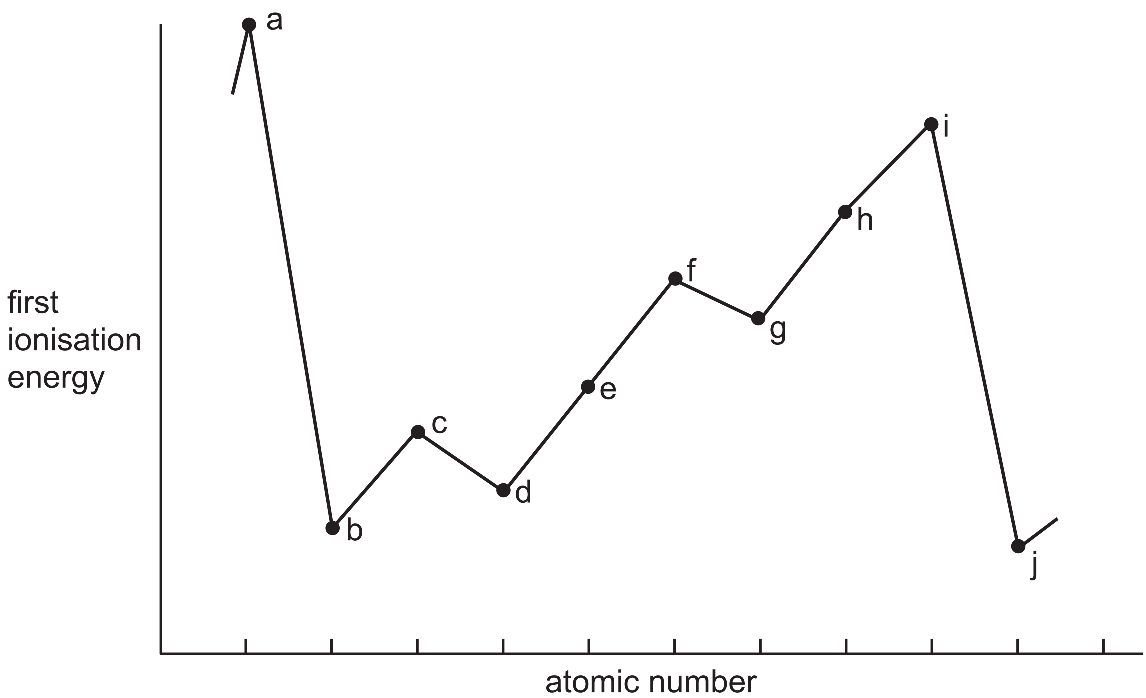
A $2Ca(NO_3)_2 \rightarrow 2CaO + 4NO_2 + O_2$

B $Cl_2 + 2I^- \rightarrow I_2 + 2Cl^-$

C $Fe + Cu^{2+} \rightarrow Fe^{2+} + Cu$

D $H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$

2 The graph of first ionisation energy against atomic number for a series of ten consecutive elements in the Periodic Table is shown below. Which one of the following indicates a Group II metal and a halogen?



	Group II metal	Halogen
A	a	h
B	b	g
C	c	h
D	c	i

3 Which one of the following is the strongest reducing agent?

A F^-

B F_2

C I^-

D I_2

4 4.35 g of potassium sulfate is dissolved in water and made up to 50.0 cm^3. Which one of the following is the concentration of potassium ions in this solution?

A 0.025 mol dm^{-3}

B 0.500 mol dm^{-3}

C 0.644 mol dm^{-3}

D 1.000 mol dm^{-3}

5 Which one of the following describes the trend in bond energies of the halogen molecules down Group VII?

A Decreases

B Decreases to bromine then increases

C Increases

D Increases to chlorine then decreases

6 When 0.28 g of a basic oxide, MO, is reacted with 250 cm^3 of 0.05 mol dm^{-3} hydrochloric acid the excess acid required 50 cm^3 of 0.05 mol dm^{-3} sodium hydroxide solution for neutralisation. Which one of the following is the relative atomic mass of M?

A 12

B 28

C 40

D 56

7 Which one of the following diagrams represents the distribution of electrons in the 3d and 4s subshells in the ground state of an iron(III) ion?

	3d					4s
A	1⇂	1⇂	1			
B	1	1	1	1	1	
C	1⇂	1				1⇂
D	1	1	1			1⇂

8 Which one of the following describes the reaction between solid sodium chloride and concentrated sulfuric acid?

A Disproportionation

B Exothermic

C Neutralisation

D Redox

9 Chlorine was bubbled through a pale green solution causing the solution to turn yellow/orange. Which one of the following ions was in the original solution?

A Br^-

B Fe^{2+}

C Fe^{3+}

D I^-

10 Which one of the following molecules does **not** contain a polar bond?

A Fluorine

B Hydrogen fluoride

C Oxygen difluoride (OF_2)

D Tetrafluoromethane (CF_4)

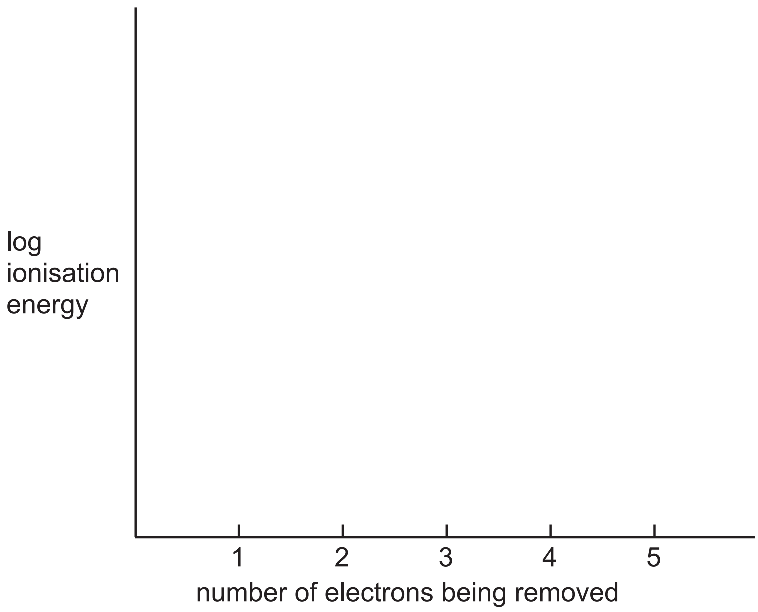
Section B

Answer **all five** questions in this section.

Examiner Only	
Marks	Remark

11 Boron is the only element in Group III of the Periodic Table which is not a metal.

(a) On the axes below sketch a graph to show the successive ionisation energies of boron.



[3]

(b) Boron trifluoride can react with a fluoride ion as shown in the equation below:

$$BF_3 + F^- \rightarrow BF_4^-$$

(i) Draw a dot and cross diagram for the BF_4^- ion and use it to suggest the shape of the ion and its bond angle.

Shape ______________________________

Bond angle ______________________________ [4]

(ii) Name the type of bond formed between the fluoride ion and boron.

____ [1]

12 Phosphorus is a non-metal with a low melting point. It reacts explosively with liquid bromine and more gently with bromine vapour. In each case phosphorus tribromide is formed.

(a) **(i)** Write an equation for the reaction of phosphorus, P_4, with bromine.

____ [2]

(ii) State the octet rule and explain whether or not phosphorus obeys the octet rule in phosphorus tribromide.

____

____

____

____ [3]

(b) The melting points of silicon, phosphorus and sulfur are given in the table below.

element	Si	P_4	S_8
melting point/°C	1410	44	113

With reference to the structures of silicon and sulfur explain why each has a higher melting point than phosphorus.

____

____

____

____

____

____

____

____ [4]

Quality of written communication [2]

Examiner Only	
Marks	Remark

13 Sodium is a reactive, soft, silvery metal. Chlorine is a poisonous gas. The two react together to form sodium chloride that is essential to our diet.

Examiner Only	
Marks	Remark

(a) **(i)** Using a labelled diagram explain the bonding in sodium metal.

____

____ [3]

(ii) Metals are good conductors of electricity. Explain why the electrical conductivity of aluminium is greater than that of sodium.

____

____

____ [2]

(b) What type of structure is present in the element chlorine?

____ [1]

(c) **(i)** Draw dot and cross diagrams to show how sodium bonds with chlorine gas. Only outer shell electrons should be shown.

[3]

(ii) Name the type of bonding in sodium chloride.

____ [1]

(iii) The structure of sodium chloride is described as a lattice. Explain what is meant by the term **lattice**.

____

____ [2]

(iv) Apart from its appearance give **three** physical properties of sodium chloride.

____

____

____ [3]

(d) Sodium chloride can be made by reacting sodium carbonate with hydrochloric acid.

(i) Write the equation for this reaction.

____ [2]

(ii) Using the following headings calculate the mass of sodium chloride formed when 5.3 g of sodium carbonate is reacted with 0.06 dm^3 of 1.5 mol dm^{-3} hydrochloric acid.

Number of moles of sodium carbonate used

____

Number of moles of hydrochloric acid used

____

State which reagent is in excess

____

Number of moles of sodium chloride formed

____

Mass of sodium chloride formed in grams

____ [5]

Examiner Only	
Marks	Remark

14 *Bromine tablets* are used as a disinfectant in hot tubs and some swimming pools because of bromine's ability to act as an oxidising agent.

Examiner Only	
Marks	Remark

(a) Bromine reacts with water in a similar way to chlorine.

(i) Suggest the equation for the reaction of bromine with water.

____ [1]

(ii) Using oxidation numbers explain why this reaction is an example of disproportionation.

____

____

____ [3]

(b) Manufacturers recommend maintaining the bromine concentration in swimming pools at 4 mg per litre. Calculate the molarity of bromine, Br_2, in the water at this concentration.

____

____

____ [2]

(c) Occasionally a 'shock treatment' with chlorine is required to further disinfect the water.

(i) Suggest, in chemical terms, why chlorine is used for this purpose.

____

____ [1]

(ii) The compound used to provide the chlorine for the shock treatment is "sodium dichlor", $NaCl_2C_3N_3O_3$. Calculate the percentage of chlorine in "sodium dichlor" to **one** decimal place.

____

____

____ [2]

Examiner Only	
Marks	Remark

(d) Bromine is produced from the reaction of sodium bromide with concentrated sulfuric acid. Name **four** other products formed when sodium bromide reacts with concentrated sulfuric acid.

1. ____

2. ____

3. ____

4. ____ [4]

(e) Describe how you could show that a solution contains bromide ions.

____

____

____ [3]

Examiner Only	
Marks	Remark

15 (a) The first three ionisation energies of calcium are given in the table below.

1st ionisation energy	2nd ionisation energy	3rd ionisation energy
590 kJ mol^{-1}	1145 kJ mol^{-1}	4912 kJ mol^{-1}

(i) Write the equation for the second ionisation of calcium including state symbols.

______________________________ [2]

(ii) Using the following headings calculate the amount of energy, in kJ, required to form 8.0 g of Ca^{2+}(g) ions from Ca(g).

Energy required to form one mole of Ca^{2+}(g) from one mole of Ca(g)

______________________________ [1]

Number of moles of Ca^{2+}(g) in 8.0 g

______________________________ [1]

Energy required to form 8.0 g of Ca^{2+}(g)

______________________________ [1]

(b) The Ca^{2+} ion has the same electron arrangement as an argon atom.

(i) Write the electron arrangement for the Ca^{2+} ion.

______________________________ [1]

(ii) The first ionisation energy of argon is 1520 kJ mol^{-1}. Explain why the third ionisation energy of calcium is much higher than the first ionisation energy of argon.

____

______________________________ [2]

Examiner Only

Marks	Remark

(c) The table below shows the relative abundance of the four main isotopes of calcium.

isotope	^{40}Ca	^{42}Ca	^{43}Ca	^{44}Ca
relative abundance	96.9%	0.6%	0.2%	2.3%

(i) What is meant by the term **isotopes**?

____

____ [2]

(ii) Calculate the relative atomic mass of calcium to **two** decimal places.

____

____

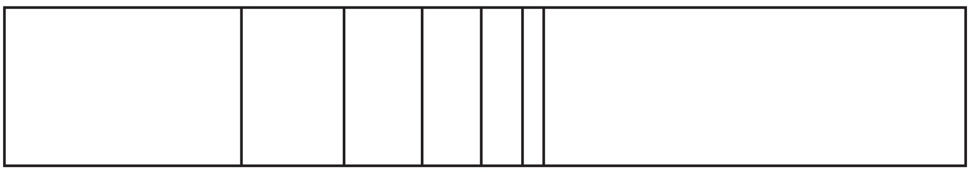
____ [2]

(iii) Complete the following table to show the number of subatomic particles in a ^{43}Ca atom.

	neutrons	**electrons**	**protons**
^{43}Ca			

[2]

(d) A line emission spectrum of calcium, shown below, can be observed through a spectroscope.



frequency

(i) Draw an arrow in the box under 'frequency' pointing in the direction in which frequency increases. [1]

Examiner Only	
Marks	Remark

(ii) Describe how the movement of an electron within an atom gives rise to a line in an emission spectrum.

____

____

____ [3]

(iii) What flame colour is observed when calcium burns?

____ [1]

(iv) Using the following headings and the first ionisation energy of calcium, 590 kJ mol^{-1}, calculate the frequency of the convergence limit of a calcium atom and state its units.

Energy, in joules, required to ionise one calcium atom

____

____

____ [2]

Frequency of the convergence limit of a calcium atom

____

____

____ [2]

THIS IS THE END OF THE QUESTION PAPER

113078