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Pearson Edexcel GCE

Centre Number | Candidate Number

Chemistry

Advanced Subsidiary

Paper 1: Core Inorganic and Physical Chemistry

Tuesday 22 May 2018 – Morning
Time: 1 hour 30 minutes

Paper Reference
8CH0/01

Candidates must have: Scientific calculator
Data Booklet

Total Marks

Instructions

- Use **black** ink or **black** ball-point pen.
- **Fill in the boxes** at the top of this page with your name, centre number and candidate number.
- Answer **all** questions.
- Answer the questions in the spaces provided
  – *there may be more space than you need.*

Information

- The total mark for this paper is 80.
- The marks for **each** question are shown in brackets
  – *use this as a guide as to how much time to spend on each question.*
- For the question marked with an **asterisk** (*), marks will be awarded for your ability to structure your answer logically showing the points that you make are related or follow on from each other where appropriate.
- A Periodic Table is printed on the back cover of this paper.

Advice

- Read each question carefully before you start to answer it.
- Check your answers if you have time at the end.
- Show all your working in calculations and include units where appropriate.

Turn over ▶

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Answer ALL questions.

Some questions must be answered with a cross in a box ☒.
If you change your mind about an answer, put a line through the box ☒ and then mark your new answer with a cross ☒.

1 This question is about covalent bonds.

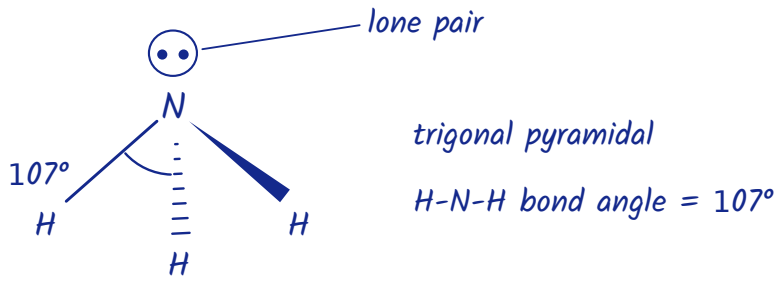
(a) State what is meant by the term covalent bond.

(2)

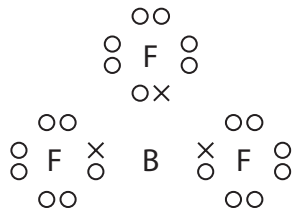
A covalent bond is the strong electrostatic attraction between two nuclei and the shared (bonding) pair of electrons between them.

(b) Draw a diagram of the ammonia molecule, clearly showing its shape.
Include any lone pairs of electrons and the value of the bond angle.

(2)



(c) The dot-and-cross diagram of BF_3 is



What is the bond angle in BF_3?

(1)

☒ **A** 90°

☒ **B** 107°

☒ **C** 109.5°

☒ **D** 120°

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(d) (i) Ammonia and boron trifluoride react to form a compound NH_3BF_3 which contains a dative covalent bond. Each of the molecules, NH_3 and BF_3, has a different feature of its electronic structure that allows this to happen.
Use these two different features to explain how a dative covalent bond is formed.

(2)

The nitrogen atom in NH_3 has a lone pair of electrons, which it donates to the boron atom.

The boron atom in BF_3 is electron deficient: it has only 6 electrons in its outer shell (an empty orbital), so it accepts the pair of electrons to complete its octet. Both electrons in the dative bond come from nitrogen.

(ii) During this reaction, the bond angles about the nitrogen atom and the boron atom change.

State the new H—N—H and F—B—F bond angles.

(2)

H-N-H bond angle becomes 109.5°

F-B-F bond angle becomes 109.5°

(Total for Question 1 = 9 marks)

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2 This question is about hydrogen, the element with atomic number $Z = 1$.

(a) (i) Hydrogen has two stable isotopes, $^{1}_{1}H$ and $^{2}_{1}H$. Complete the table to show the number of subatomic particles present in the nuclei of these two isotopes of hydrogen.

(1)

Isotope	Number of protons	Number of neutrons
$^{1}_{1}H$	1	0
$^{2}_{1}H$	1	1

(ii) Use the data in the table to explain the term isotopes.

(2)

Both are atoms of hydrogen with the same number of protons (1) but a different number of neutrons (0 in hydrogen-1 and 1 in hydrogen-2), so they are isotopes.

(b) The relative atomic mass of hydrogen in the Periodic Table is 1.0.
This is correct to two significant figures.

The table gives data for the relative isotopic mass and natural abundance of the two stable isotopes of hydrogen.

Isotope	Relative isotopic mass	Percentage abundance
$^{1}_{1}H$	1.007825	99.9885
$^{2}_{1}H$	2.014101	0.0115

(i) Using the data in the table, give a reason why it can be estimated that the relative atomic mass of hydrogen is greater than 1.0.

(1)

Both isotopes have a relative isotopic mass greater than 1 (1.007825 and 2.014101) and there is no isotope with a mass less than 1, so the average must be greater than 1.0.

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(ii) Calculate the relative atomic mass of hydrogen from these data, giving your answer to **four** decimal places.

(2)

$A_r = [(1.007825 \times 99.9885) + (2.014101 \times 0.0115)] \div 100$

$A_r = 1.0079$ (to 4 d.p.)

(c) (i) Write an equation to represent the first ionisation energy of hydrogen. Include state symbols.

(2)

$H(g) \rightarrow H^+(g) + e^-$

(ii) The sequence of the first three elements in the Periodic Table is hydrogen, helium and then lithium.

Explain why the first ionisation energy of hydrogen is less than that of helium, but greater than that of lithium.

(4)

Helium has more protons than hydrogen (2 compared with 1), so it has a greater nuclear charge.

The outer electrons in helium are in the same shell (1s) as the electron in hydrogen, so they are more strongly attracted to the nucleus and more energy is needed to remove one.

In lithium the outer electron is in a new shell (2s), further from the nucleus, and it is shielded by the inner $1s^2$ electrons, so it is less strongly attracted than the 1s electron in hydrogen, despite lithium's greater nuclear charge.

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(d) Hydrogen can be placed in several different positions in periodic tables. One is immediately above lithium in Group 1. Another is in the centre of the first row, as shown in the Periodic Table on the back cover.

Criticise the position of hydrogen immediately above lithium by giving one reason in favour and two against.

(3)

In favour: hydrogen has one electron in its outer shell ($1s^1$), like the Group 1 elements (ns^1), and can form 1+ ions.

Against: the rest of Group 1 are (alkali) metals, but hydrogen is a non-metal.

Against: hydrogen has different chemical properties from the rest of Group 1, e.g. it forms covalent bonds, and it can gain one electron to form the H^- ion.

(Total for Question 2 = 15 marks)

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3 The reaction of sulfuric acid with potassium hydroxide is a neutralisation.
The equation for this reaction is

$$H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$$

A titration was carried out using the following method.

1. Potassium hydroxide solution of unknown concentration was placed in a burette and the initial reading was recorded.
2. 25.0 cm^3 of sulfuric acid solution, concentration 0.0800 mol dm^{-3}, was transferred to a conical flask.
3. Three drops of phenolphthalein indicator were added to the sulfuric acid.
4. Potassium hydroxide was added from the burette until the solution just changed colour and then the burette reading was recorded.
5. Repeat titrations were carried out until concordant titres were obtained.

(a) Select the most appropriate piece of apparatus to measure the 25.0 cm^3 of sulfuric acid.

(1)

☒ **A** burette

☒ **B** measuring cylinder

☒ **C** pipette

☒ **D** volumetric flask

(b) What is the colour of the solution when neutralisation has just occurred?

(1)

☒ **A** colourless

☒ **B** orange

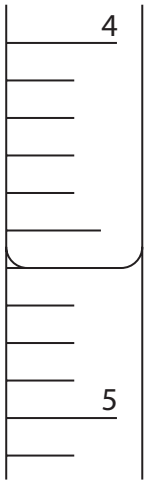
☒ **C** pale pink

☒ **D** red

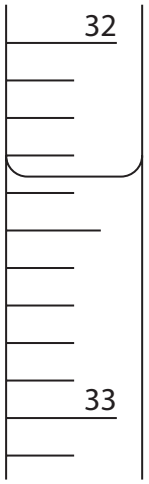
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(c) (i) Complete the table of results for titration number 1, using the diagrams to find the initial and final burette readings.

(2)



Initial reading



Final reading

Table of results

Titration number	Final reading / cm^3	Initial reading / cm^3	Titration volume / cm^3
1	32.35	4.60	27.75
2	28.05	1.10	26.95
3	37.65	10.20	27.45
4	32.05	5.00	27.05

(ii) The best value for the mean titre of this reaction is

(1)

☒ **A** 27.00 cm^3

☐ **B** 27.15 cm^3

☐ **C** 27.25 cm^3

☐ **D** 27.30 cm^3

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(iii) Calculate the concentration, in mol dm^{-3}, of the potassium hydroxide solution, giving your answer to an appropriate number of significant figures.

(3)

$n(H_2SO_4) = 0.0800 \times 25.0 \div 1000 = 0.00200$ mol

$n(KOH) = 2 \times 0.00200 = 0.00400$ mol

$[KOH] = 0.00400 \div (27.00 \div 1000)$

$[KOH] = 0.148$ mol dm^{-3} (3 s.f.)

(Total for Question 3 = 8 marks)

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4 An ionic compound contains a metal cation and a non-metal anion in a 1 : 1 ratio, and water of crystallisation. The compound can be represented as $MN.xH_2O$, where x is the number of moles of water of crystallisation per mole of MN.

A sample of $MN.xH_2O$ was dissolved in distilled water to produce a colourless solution, with a concentration of about 0.5 mol dm^{-3}. 2 cm^3 of the resulting solution was transferred to each of two test tubes.

The following tests were carried out to identify the ions present.

(a) **Test 1**

(i) Addition of a few drops of a solution of barium chloride to one of the test tubes gave a white precipitate.

Identify, by name or formula, **two** possible anions that would give this result.

(1)

sulfate, SO_4^{2-} and carbonate, CO_3^{2-}

(ii) Addition of 1 cm^3 of dilute hydrochloric acid to the test tube in (a)(i) resulted in no further change.

Give the **formula** of the anion.

(1)

SO_4^{2-}

(iii) What is the charge on the cation?

(1)

☒ **A** +1

☒ **B** −1

☒ **C** +2

☒ **D** −2

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(b) **Test 2**

A flame test on a sample of solid $MN.xH_2O$ gave no change in the flame colour.

Give a possible identity of the cation, M.

(1)

magnesium ion, Mg^{2+}

(c) Heating the hydrated compound results in the formation of the anhydrous ionic solid MN by the following reaction:

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

Heating a sample of the hydrated compound reduced the mass to 48.9% of its original value.

Use this information and your answer to (a)(ii) and (b) to calculate the value of x.

[*Note: If you have been unable to identify* MN, *you may use this hydrated compound,* $CoCl_2.yH_2O$ *in which the sample reduced in mass to* 54.6% *of its original value. Use this information to calculate the value of* y.]

(4)

MN is $MgSO_4$: $M_r = 24.3 + 32.1 + (4 \times 16.0) = 120.4$

In 100 g: mass $MgSO_4$ = 48.9 g, mass $H_2O = 100 - 48.9 = 51.1$ g

$n(MgSO_4) = 48.9 \div 120.4 = 0.4061$ mol

$n(H_2O) = 51.1 \div 18.0 = 2.839$ mol

ratio $H_2O : MgSO_4 = 2.839 \div 0.4061 = 6.99$

so the formula is $MgSO_4.7H_2O$

$x = 7$

(Total for Question 4 = 8 marks)

5 A student made crystals of a metal chloride, $JCl_2.6H_2O$, by reacting the metal carbonate, JCO_3, with hydrochloric acid, HCl(aq). The product was purified.

Procedure

Step **1** 150 cm^3 of hydrochloric acid, concentration 0.80 mol dm^{-3}, was transferred to a 400 cm^3 conical flask. The flask was warmed gently using a Bunsen burner. A spatula measure (about 1.0 g) of metal carbonate was added to the acid.

Step **2** When the reaction in Step **1** was finished, more metal carbonate was added until the metal carbonate was in excess.

Step **3** The resulting mixture was filtered into an evaporating basin.

Step **4** The evaporating basin was heated using a Bunsen burner to concentrate the solution. The concentrated solution was allowed to cool and crystallise.

Step **5** Once crystal formation was complete, the resulting mixture was filtered for a second time.

Step **6** The resulting white crystals were rinsed with a small volume of ice-cold water.

The equation for the reaction between the metal carbonate and hydrochloric acid is

$$JCO_3(s) + 2HCl(aq) \rightarrow JCl_2(aq) + H_2O(l) + CO_2(g)$$

(a) (i) Describe **two** observations that the student might make which show that the reaction in Step **1** has finished.

(2)

The fizzing (effervescence) stops, as no more bubbles of CO_2 are given off.

All of the solid metal carbonate disappears (dissolves), leaving a clear solution.

(ii) State the purpose of the filtration in Step **3**.

(1)

To remove the excess (unreacted) metal carbonate, which is an insoluble solid.

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(iii) Explain the use of a small volume of ice-cold water in Step **6**.

(2)

A small volume of ice-cold water is used so that as little of the product (crystals) as possible dissolves and is lost.

It rinses away any soluble impurities left on the surface of the crystals.

(b) The student obtained a mass of 14.26 g of hydrated crystals.
Assuming that the percentage yield is 100%, use the information in the procedure to give a possible identity of J.

(5)

$n(HCl) = 150 \div 1000 \times 0.80 = 0.120$ mol

$n(JCl_2.6H_2O) = 0.120 \div 2 = 0.0600$ mol

$M_r(JCl_2.6H_2O) = 14.26 \div 0.0600 = 237.7$ g mol^{-1}

$A_r(J) = 237.7 - (2 \times 35.5) - (6 \times 18.0)$

$A_r(J) = 237.7 - 71.0 - 108.0 = 58.7$

So J is nickel, Ni ($A_r = 58.7$)

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(c) The student was surprised by the white colour of the crystals of $JCl_2.6H_2O$ in Step **6**. This did not agree with the possible identity for J from the calculation in (b). The student decided to perform a flame test on the crystals.

(i) Explain why the student was surprised and decided to carry out a flame test.

(2)

Nickel is a transition metal and its compounds are coloured (nickel chloride is green), not white.

White crystals suggest a Group 2 (s-block) metal, so a flame test was used to identify the metal ion (cation).

(ii) The flame test colour was crimson red. Identify J.

(1)

- ☒ **A** barium
- ☒ **B** calcium
- ☒ **C** lithium
- ☒ **D** strontium

(iii) Calculate the actual percentage yield of the reaction, which produced 14.26 g of crystals.

Give your answer to **two** significant figures.

(2)

$M_r(SrCl_2.6H_2O) = 87.6 + 71.0 + 108.0 = 266.6$

maximum mass = $0.0600 \times 266.6 = 15.996$ g

% yield = $14.26 \div 15.996 \times 100 = 89.1$ %

percentage yield = 89 %

(Total for Question 5 = 15 marks)

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6 Chlorine and iodine are in the same group in the Periodic Table.

(a) (i) Complete the electronic configuration of chlorine using the s, p, d notation.

(1)

$1s^2$ $2s^2\ 2p^6\ 3s^2\ 3p^5$

(ii) Explain why iodine and chlorine have many similar chemical reactions.

(2)

Iodine also has 7 electrons in its outer shell ($5s^2 5p^5$, like $3s^2 3p^5$ in chlorine).

Chemical reactions are governed by the electronic configuration (the number of outer shell electrons), so they have many similar reactions.

(b) Members of the same group sometimes react in different ways.

Iodine and chlorine react differently with thiosulfate ions, $S_2O_3^{2-}$.
Iodine gives $S_4O_6^{2-}$, whilst chlorine gives SO_4^{2-}.

(i) Complete the table by identifying the oxidation numbers of sulfur in the three sulfur-containing ions.

(2)

Ion	Oxidation number of sulfur
$S_2O_3^{2-}$	+2
SO_4^{2-}	+6
$S_4O_6^{2-}$	+2.5

(ii) The equation for the reaction of iodine with thiosulfate ions is

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

State, in terms of electrons, why iodine is classified as an oxidising agent in this reaction.

(1)

Iodine gains electrons ($I_2 + 2e^- \rightarrow 2I^-$), taking them from the thiosulfate ions.

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(iii) Use your answer to b(i) to show that chlorine is a stronger oxidising agent than iodine.

(1)

Chlorine oxidises sulfur from +2 to +6, but iodine only oxidises it from +2 to +2.5, so chlorine causes a greater increase in oxidation number.

(iv) Chlorine reacts in aqueous solution with $S_2O_3^{2-}$ to give SO_4^{2-}.
The ionic half-equation for the reaction of chlorine is

$$Cl_2 + 2e^- \rightarrow 2Cl^-$$

Write the ionic half-equation for the reaction of aqueous $S_2O_3^{2-}$ to give SO_4^{2-}.
State symbols are not required.

(2)

$S_2O_3^{2-} + 5H_2O \rightarrow 2SO_4^{2-} + 10H^+ + 8e^-$

(v) Use your answer to (b)(iv) and the half-equation for chlorine, to write the overall ionic equation for the reaction between chlorine and thiosulfate ions.
State symbols are not required.

(1)

$4Cl_2 + S_2O_3^{2-} + 5H_2O \rightarrow 8Cl^- + 2SO_4^{2-} + 10H^+$

(Total for Question 6 = 10 marks)

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*7 The compounds hydrogen fluoride, water and methane, all have simple molecular structures, but they have significantly different boiling temperatures.

Discuss the reasons for the differences in the boiling temperatures of the three compounds, using the data in the table and the Pauling electronegativity values in the Data Booklet.

Compound	Boiling temperature /°C	Number of electrons
CH_4	−161.5	10
H_2O	100.0	10
HF	19.5	10

(6)

All three compounds have the same number of electrons (10), so the London (dispersion) forces between their molecules are similar in strength. The large differences in boiling temperature must therefore be caused by other intermolecular forces.

Pauling electronegativity values: F = 4.0, O = 3.5, C = 2.5 and H = 2.1. So the electronegativity difference is 1.9 in H-F, 1.4 in O-H, but only 0.4 in C-H.

CH_4 has only a small electronegativity difference and is a non-polar, symmetrical molecule, so there are only weak London forces between its molecules (no hydrogen bonds). Very little energy is needed to overcome them, so methane has the lowest boiling temperature (−161.5 °C).

In HF and H_2O, hydrogen is bonded to a very electronegative atom which has lone pairs, so the bonds are highly polar and there is hydrogen bonding between the molecules, as well as London forces. Hydrogen bonds are much stronger than London forces, so much more energy is needed to separate the molecules, and both HF and H_2O have much higher boiling temperatures than CH_4.

Although the H-F bond is more polar than the O-H bond, each H_2O molecule has two hydrogen atoms and two lone pairs, so water forms on average two hydrogen bonds per molecule, whereas HF (only one hydrogen atom per molecule) forms on average only one. So there are twice as many hydrogen bonds in water as in HF. More energy is needed

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to overcome the stronger overall intermolecular forces in water, so water (100.0 °C) has a higher boiling temperature than HF (19.5 °C), which in turn boils at a much higher temperature than CH_4.

(Total for Question 7 = 6 marks)

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8 This question is about ionic bonding.

(a) The elements sodium and fluorine react together to form an ionic compound.

(i) Select the correct equation for this reaction.

(1)

☒ **A** $Na(s) + F(g) \rightarrow NaF(s)$

☒ **B** $2Na(s) + F_2(g) \rightarrow 2NaF(s)$

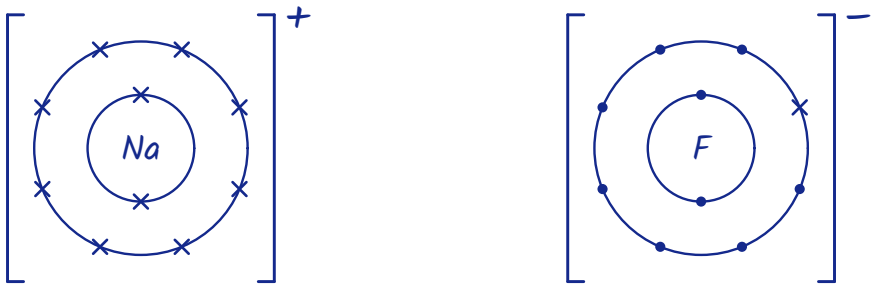
☒ **C** $Na(s) + F_2(g) \rightarrow NaF_2(s)$

☒ **D** $2Na(s) + F(g) \rightarrow Na_2F(s)$

(ii) Draw dot-and-cross diagrams of the ions in sodium fluoride, showing all the electrons.

Use your diagram to explain why the ions are described as isoelectronic.

(3)



Na^+ 2.8 F^- 2.8

x = electron from Na, · = electron from F

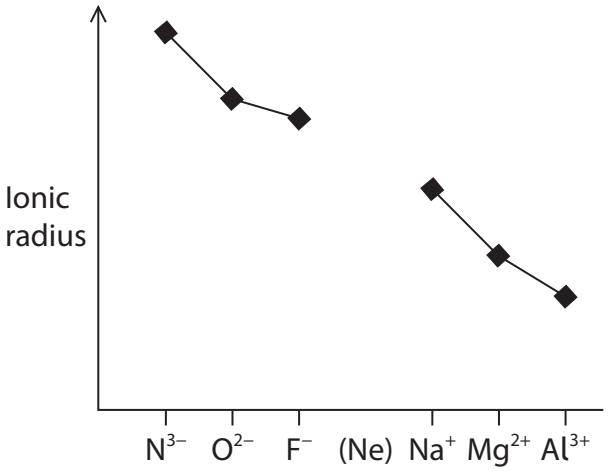
Both ions have 10 electrons arranged 2.8 ($1s^2 2s^2 2p^6$), the same electronic configuration, so they are isoelectronic.

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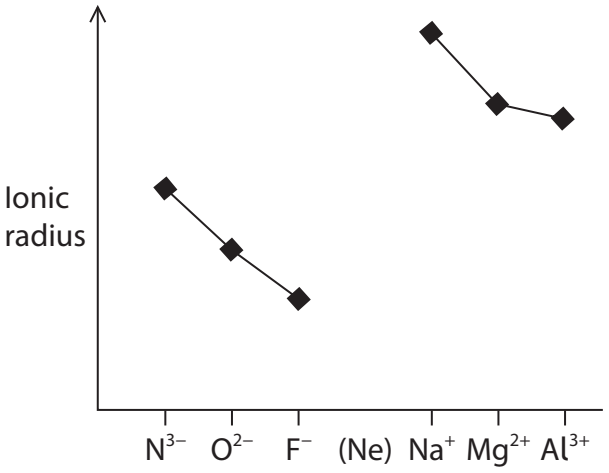
(iii) Which diagram shows the trend in ionic radius for the isoelectronic ions N^{3-} to Al^{3+}?

(1)

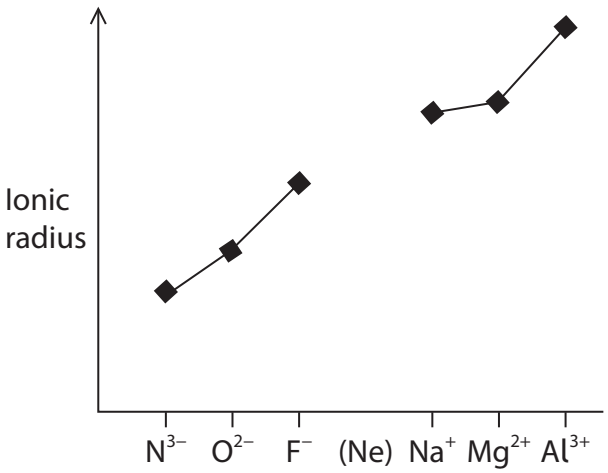
☒ **A**



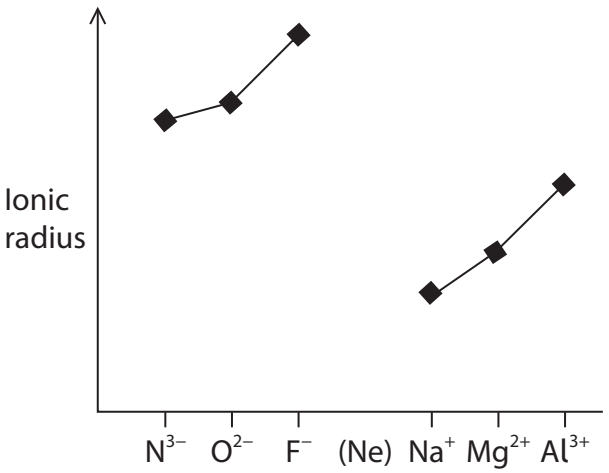
☒ **B**



☒ **C**



☒ **D**



(iv) Explain your answer to (a)(iii) in terms of the structure of the ions.

(2)

From N^{3-} to Al^{3+} the number of protons in the nucleus increases (from 7 to 13), while every ion has the same number of electrons (10) in the same shells.

So the nuclear charge increases, which increases the attraction for the electrons, pulling them closer to the nucleus, so the ionic radius decreases.

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(b) The strength of ionic bonding in different compounds can be compared by using the amount of energy required to separate the ions. Some values for this energy are given in the table.

Compound	Amount of energy required to separate the ions / kJ mol^{-1}
LiF	1031
KF	817
CaF_2	2957

Using the data provided, explain how changes in the cation affect the bond strength in an ionic compound.

(2)

The higher the charge on the cation, the stronger the attraction between the ions: CaF_2 contains Ca^{2+} (2+) whereas LiF and KF contain 1+ ions (Li^+, K^+), so CaF_2 needs far more energy (2957 kJ mol^{-1}).

The smaller the radius of the cation, the stronger the attraction between the ions: Li^+ is smaller than K^+, so LiF (1031 kJ mol^{-1}) has stronger ionic bonding than KF (817 kJ mol^{-1}).

(Total for Question 8 = 9 marks)

TOTAL FOR PAPER = 80 MARKS

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The Periodic Table of Elements

1	2											3	4	5	6	7	0 (8) (18)
							1.0 H hydrogen 1										4.0 He helium 2
(1)	(2)		Key									(13)	(14)	(15)	(16)	(17)	
6.9 Li lithium 3	9.0 Be beryllium 4		relative atomic mass atomic symbol name atomic (proton) number									10.8 B boron 5	12.0 C carbon 6	14.0 N nitrogen 7	16.0 O oxygen 8	19.0 F fluorine 9	20.2 Ne neon 10
23.0 Na sodium 11	24.3 Mg magnesium 12	(3)	(4)	(5)	(6)	(7)	(8)	(9)	(10)	(11)	(12)	27.0 Al aluminium 13	28.1 Si silicon 14	31.0 P phosphorus 15	32.1 S sulfur 16	35.5 Cl chlorine 17	39.9 Ar argon 18
39.1 K potassium 19	40.1 Ca calcium 20	45.0 Sc scandium 21	47.9 Ti titanium 22	50.9 V vanadium 23	52.0 Cr chromium 24	54.9 Mn manganese 25	55.8 Fe iron 26	58.9 Co cobalt 27	58.7 Ni nickel 28	63.5 Cu copper 29	65.4 Zn zinc 30	69.7 Ga gallium 31	72.6 Ge germanium 32	74.9 As arsenic 33	79.0 Se selenium 34	79.9 Br bromine 35	83.8 Kr krypton 36
85.5 Rb rubidium 37	87.6 Sr strontium 38	88.9 Y yttrium 39	91.2 Zr zirconium 40	92.9 Nb niobium 41	95.9 Mo molybdenum 42	[98] Tc technetium 43	101.1 Ru ruthenium 44	102.9 Rh rhodium 45	106.4 Pd palladium 46	107.9 Ag silver 47	112.4 Cd cadmium 48	114.8 In indium 49	118.7 Sn tin 50	121.8 Sb antimony 51	127.6 Te tellurium 52	126.9 I iodine 53	131.3 Xe xenon 54
132.9 Cs caesium 55	137.3 Ba barium 56	138.9 La* lanthanum 57	178.5 Hf hafnium 72	180.9 Ta tantalum 73	183.8 W tungsten 74	186.2 Re rhenium 75	190.2 Os osmium 76	192.2 Ir iridium 77	195.1 Pt platinum 78	197.0 Au gold 79	200.6 Hg mercury 80	204.4 Tl thallium 81	207.2 Pb lead 82	209.0 Bi bismuth 83	[209] Po polonium 84	[210] At astatine 85	[222] Rn radon 86
[223] Fr francium 87	[226] Ra radium 88	[227] Ac* actinium 89	[261] Rf rutherfordium 104	[262] Db dubnium 105	[266] Sg seaborgium 106	[264] Bh bohrium 107	[277] Hs hassium 108	[268] Mt meitnerium 109	[271] Ds darmstadtium 110	[272] Rg roentgenium 111							

Elements with atomic numbers 112-116 have been reported but not fully authenticated

* Lanthanide series	140 Ce cerium 58	141 Pr praseodymium 59	144 Nd neodymium 60	[147] Pm promethium 61	150 Sm samarium 62	152 Eu europium 63	157 Gd gadolinium 64	159 Tb terbium 65	163 Dy dysprosium 66	165 Ho holmium 67	167 Er erbium 68	169 Tm thulium 69	173 Yb ytterbium 70	175 Lu lutetium 71
* Actinide series	232 Th thorium 90	[231] Pa protactinium 91	238 U uranium 92	[237] Np neptunium 93	[242] Pu plutonium 94	[243] Am americium 95	[247] Cm curium 96	[245] Bk berkelium 97	[251] Cf californium 98	[254] Es einsteinium 99	[253] Fm fermium 100	[256] Md mendelevium 101	[254] No nobelium 102	[257] Lr lawrencium 103

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