Cambridge Assessment International Education

Interactive

Interactive Example Candidate Responses
Paper 22 (May/June 2016), Question 2

Cambridge International AS & A Level Chemistry 9701

Version 1.0

Cambridge Pathway

In order to help us develop the highest quality resources, we are undertaking a continuous programme of review; not only to measure the success of our resources but also to highlight areas for improvement and to identify new development needs.

We invite you to complete our survey by visiting the website below. Your comments on the quality and relevance of our resources are very important to us.

www.surveymonkey.co.uk/r/GL6ZNJB

Would you like to become a Cambridge International consultant and help us develop support materials?

Please follow the link below to register your interest.

www.cambridgeinternational.org/cambridge-for/teachers/teacherconsultants/

Select page

HIGH | MIDDLE | LOW

2 The elements in Group 17, the halogens, and their compounds, show many similarities and trends in their properties. Some data are given for the elements fluorine to iodine.

element	bond energy /kJ mol^{-1}	standard enthalpy change of atomisation, $\Delta H_{at}^{\ominus}$ /kJ mol^{-1}	boiling point of element /K	boiling point of **hydrogen halide** /K
fluorine, F–F	158	79	85	293
chlorine, Cl–Cl	242	121	238	188
bromine, Br–Br	193	112	332	206
iodine, I–I	151	107	457	238

(a) (i) Explain the meaning of the term *standard enthalpy change of atomisation*.

The enthalpy change when one mole of gaseous atoms is formed from its elements under standard conditions. The element should be in its standard state. [3]

(ii) For fluorine and chlorine, the enthalpy changes of atomisation are half the value of the bond energies.

For bromine and iodine, the enthalpy changes of atomisation are much more than half the value of the bond energies.

Suggest a reason for this difference.

Fluorine and chlorine are in gaseous form at room temperature. Bromine and iodine is a liquid and a solid gas respectively. Energy is needed to change their states. [1]

(iii) The standard enthalpy of formation of iodine monochloride, ICl, is –24.0 kJ mol^{-1}.

Use this information and the bond energies of iodine and chlorine to calculate the I–Cl bond energy.

$\frac{1}{2}I_2 + \frac{1}{2}Cl_2 \rightarrow ICl$

$\frac{1}{2}I\text{-}I + \frac{1}{2}Cl\text{-}Cl \rightarrow I\text{-}Cl$

Bonds broken

$E(\frac{1}{2}I\text{-}I) = \frac{151}{2} = 75.5$

$E(\frac{1}{2}Cl\text{-}Cl) = \frac{242}{2} = 121$

Bonds formed

$E(I\text{-}Cl) = x$

So, $+75.5 + 121 - x = -24$

$196.5 - x = -24$

$x = 196.5 + 24$

$x = 220.5$ ✱

I–Cl bond energy = 220.5 kJ mol^{-1} [2]

	Your Mark
2(a)(i)	
2(a)(ii)	
2(a)(iii)	
2(b)(i)	
2(b)(ii)	
2(c)(i)	
2(c)(ii)	
2(c)(iii)	
2(c)(iv)	
2(d)(i)	
2(d)(ii)	

Q2	Mark scheme
(a)(i)	enthalpy / energy / heat change when one mole of gaseous atoms is produced [1] from the element in its standard state [1] under standard conditions [1] **[3]**
(a)(ii)	fluorine and chlorine are gases / bromine liquid and iodine solid OR as ΔH_{at} for bromine / iodine also includes changes of state **[1]**
(a)(iii)	($\frac{1}{2}Cl_2 + \frac{1}{2}I_2 \rightarrow ICl$) $\Delta H_f = (\frac{1}{2}E(Cl_2) + \frac{1}{2}E(I_2)) - E(ICl)$ OR $E(ICl) = (151 / 2) + (242 / 2) + 24$ [1] $E(ICl) = (+)\ 220.5 / 221$ [1] **[2]**
(b)(i)	stronger / more / greater id–id / London / dispersion forces [1] due to increasing numbers of electrons [1] **[2]**
(b)(ii)	(intermolecular forces in HF are) hydrogen bonds [1] (which are) stronger (than vdW) / more energy needed to separate molecules [1] OR HF much more polar / F much more electronegative [1] Intermolecular forces in HF stronger (than in HCl, HBr, HI) [1] **[2]**
(c)(i)	**P** = iodine / I_2 /I; **Q** = chlorine / Cl_2 / Cl **[1]**
(c)(ii)	weaker H–**P** than H–**Q** bond ORA / easier /less energy to break H–**P** than H–**Q** ORA [1] due to greater distance / shielding of nucleus from bond pair ORA [1] **[2]**
(c)(iii)	2H**P** (or 2HI) → (or) H_2 + $\mathbf{P}_2$ (or I_2) **[1]**
(c)(iv)	$Ag^+(aq) + \mathbf{Q}^-(aq)$ (or Cl^-) → Ag**Q**(s) (or AgCl(s)) [1] Ag**Q**(s) / AgCl(s) + NH_3(aq) → $Ag(NH_3)_2^+(aq) + \mathbf{Q}^-(aq)$ / $Cl^-(aq)$ [1] **[2]**
(d)(i)	no of Cl increases by one each time / matches group number [1] due to increasing number of valence/outer(most/shell) electrons / oxidation number / valency (of Mg, Al, Si) [1] **[2]**
(d)(ii)	$MgCl_2$ (+aq) → $Mg^{2+} + 2Cl^-$ [1] $AlCl_3 + 6H_2O \rightarrow Al(H_2O)_6^{3+} + 3Cl^- / Al(H_2O)_5(OH)^{2+} + H^+ + 3Cl^-$ [1] $SiCl_4 + 2H_2O \rightarrow SiO_2 + 4H^+ + 4Cl^-$ [1] **[3]** **[Total:21]**

EXAMINER COMMENTS | MARK SCHEME

Select page

HIGH | MIDDLE | LOW

(b) (i) Explain the trend in the boiling points of the hydrogen halides, HCl, HBr and HI.

Number of electrons increases from HCl & HBr & HI. So, strength of van der waals increases from HCl to HI. Greater energy is needed to overcome the forces. [2]

(ii) Suggest why the hydrogen halide HF does not follow the trend in boiling points shown by HCl, HBr and HI.

Fluorine is more electronegative than hydrogen. So, HF has hydrogen bonds between the molecules. Hydrogen bonds are stronger than van der waals and greater energy is needed to overcome. [2]

(c) In an experiment, two of the halogens are represented as P_2 and Q_2.

P_2 combines with hydrogen on heating to form HP, which can be easily broken down into its elements. A solution of HP in water reacts with aqueous silver ions to form a yellow precipitate that is insoluble in dilute aqueous ammonia.

Q_2 combines explosively with hydrogen in sunlight to form HQ, which is stable to heat. A solution of HQ in water reacts with aqueous silver ions to form a white precipitate that is soluble in dilute aqueous ammonia.

(i) Identify the halogens P_2 and Q_2.

P_2 = Iodine / I_2 Q_2 = Chlorine / Cl_2 [1]

(ii) HP readily decomposes into its elements when heated but HQ is stable to heat. Explain this with reference to bond energies.

H-P bond length is ~~greater~~ greater than bond length of H-Q. So, H-P has bond energy of 299 $kJmol^{-1}$ which is less than bond energy of H-Q (431 $kJmol^{-1}$) [2]

(iii) Write an equation for the thermal decomposition of HP.

$2HI \rightleftharpoons H_2 + I_2$ [1]

Your Mark

2(a)(i) ☐

2(a)(ii) ☐

2(a)(iii) ☐

2(b)(i) ☐

2(b)(ii) ☐

2(c)(i) ☐

2(c)(ii) ☐

2(c)(iii) ☐

2(c)(iv) ☐

2(d)(i) ☐

2(d)(ii) ☐

Q2	Mark scheme
(a)(i)	enthalpy / energy / heat change when one mole of gaseous atoms is produced [1] from the element in its standard state [1] under standard conditions [1] **[3]**
(a)(ii)	fluorine and chlorine are gases / bromine liquid and iodine solid OR as ΔH_{at} for bromine / iodine also includes changes of state **[1]**
(a)(iii)	$(\frac{1}{2}Cl_2 + \frac{1}{2}I_2 \rightarrow ICl)$ $\Delta H_f = (\frac{1}{2}E(Cl_2) + \frac{1}{2}E(I_2)) - E(ICl)$ OR $E(ICl) = (151 / 2) + (242 / 2) + 24$ [1] $E(ICl) = (+) 220.5 / 221$ [1] **[2]**
(b)(i)	stronger / more / greater id–id / London / dispersion forces [1] due to increasing numbers of electrons [1] **[2]**
(b)(ii)	(intermolecular forces in HF are) hydrogen bonds [1] (which are) stronger (than vdW) / more energy needed to separate molecules [1] OR HF much more polar / F much more electronegative [1] Intermolecular forces in HF stronger (than in HCl, HBr, HI) [1] **[2]**
(c)(i)	**P** = iodine / I_2 /I; **Q** = chlorine / Cl_2 / Cl **[1]**
(c)(ii)	weaker H–**P** than H–**Q** bond ORA / easier /less energy to break H–**P** than H–**Q** ORA [1] due to greater distance / shielding of nucleus from bond pair ORA [1] **[2]**
(c)(iii)	$2H\mathbf{P}$ (or 2HI) → (or) H_2 + $\mathbf{P}_2$ (or I_2) **[1]**
(c)(iv)	$Ag^+(aq) + \mathbf{Q}^-(aq)$ (or Cl^-) → $Ag\mathbf{Q}(s)$ (or AgCl(s)) [1] $Ag\mathbf{Q}(s)$ / AgCl(s) + $NH_3(aq) \rightarrow Ag(NH_3)_2^+(aq) + \mathbf{Q}^-(aq)$ / $Cl^-(aq)$ [1] **[2]**
(d)(i)	no of Cl increases by one each time / matches group number [1] due to increasing number of valence/outer(most/shell) electrons / oxidation number / valency (of Mg, Al, Si) [1] **[2]**
(d)(ii)	$MgCl_2 (+aq) \rightarrow Mg^{2+} + 2Cl^-$ [1] $AlCl_3 + 6H_2O \rightarrow Al(H_2O)_6^{3+} + 3Cl^- / Al(H_2O)_5(OH)^{2+} + H^+ + 3Cl^-$ [1] $SiCl_4 + 2H_2O \rightarrow SiO_2 + 4H^+ + 4Cl^-$ [1] **[3]** **[Total:21]**

EXAMINER COMMENTS | MARK SCHEME

Select page

HIGH | MIDDLE | LOW

(iv) Write ionic equations, including state symbols, for

1. the formation of the white precipitate on addition of aqueous silver ions to aqueous HQ,

$Ag^+_{(aq)} + Cl^-_{(aq)} \rightarrow AgCl_{(s)}$

2. the subsequent dissolving of this precipitate in dilute aqueous ammonia.

$Ag^+_{(aq)} + 2NH_{3(aq)} \rightarrow [Ag(NH_3)_2]^+_{(aq)}$ [2]

(d) Chlorine reacts directly with many elements to form chlorides. Three such compounds are $MgCl_2$, $AlCl_3$ and $SiCl_4$.

(i) State and explain the pattern shown by the formulae of these three chlorides.

Number of Chlorine atoms in an ionic compound increases from $MgCl_2$ to $SiCl_4$. The oxidation state increases from Mg to Si. So, more chlorine atoms are needed to gain the electrons. [2]

(ii) Write equations to show the behaviour of each of these chlorides when added to water.

$MgCl_2$ ~~$MgCl_2 + H_2O$~~ $\rightarrow MgCl_2 + aq \rightarrow Mg^{2+} + 2Cl^-$

$AlCl_3$ $[Al(H_2O)_6]^{+3} \rightarrow [Al(H_2O_5)OH]^{+2} + H^+$

$SiCl_4$ $SiCl_4 + 2H_2O \rightarrow SiO_2 + 4HCl$ [3]

[Total: 21]

	Your Mark
2(a)(i)	
2(a)(ii)	
2(a)(iii)	
2(b)(i)	
2(b)(ii)	
2(c)(i)	
2(c)(ii)	
2(c)(iii)	
2(c)(iv)	
2(d)(i)	
2(d)(ii)	

Q2	Mark scheme
(a)(i)	enthalpy / energy / heat change when one mole of gaseous atoms is produced [1] from the element in its standard state [1] under standard conditions [1] **[3]**
(a)(ii)	fluorine and chlorine are gases / bromine liquid and iodine solid OR as ΔH_{at} for bromine / iodine also includes changes of state **[1]**
(a)(iii)	$(\frac{1}{2}Cl_2 + \frac{1}{2}I_2 \rightarrow ICl)$ $\Delta H_f = (\frac{1}{2}E(Cl_2) + \frac{1}{2}E(I_2)) - E(ICl)$ OR $E(ICl) = (151 / 2) + (242 / 2) + 24$ [1] $E(ICl) = (+) 220.5 / 221$ [1] **[2]**
(b)(i)	stronger / more / greater id–id / London / dispersion forces [1] due to increasing numbers of electrons [1] **[2]**
(b)(ii)	(intermolecular forces in HF are) hydrogen bonds [1] (which are) stronger (than vdW) / more energy needed to separate molecules [1] OR HF much more polar / F much more electronegative [1] Intermolecular forces in HF stronger (than in HCl, HBr, HI) [1] **[2]**
(c)(i)	**P** = iodine / I_2 /I; **Q** = chlorine / Cl_2 / Cl **[1]**
(c)(ii)	weaker H–**P** than H–**Q** bond ORA / easier /less energy to break H–**P** than H–**Q** ORA [1] due to greater distance / shielding of nucleus from bond pair ORA [1] **[2]**
(c)(iii)	2H**P** (or 2HI) → (or) H_2 + $\mathbf{P}_2$ (or I_2) **[1]**
(c)(iv)	$Ag^+(aq) + \mathbf{Q}^-(aq)$ (or Cl^-) → Ag**Q**(s) (or AgCl(s)) [1] Ag**Q**(s) / AgCl(s) + NH_3(aq) → $Ag(NH_3)_2^+(aq) + \mathbf{Q}^-(aq)$ / $Cl^-(aq)$ [1] **[2]**
(d)(i)	no of Cl increases by one each time / matches group number [1] due to increasing number of valence/outer(most/shell) electrons / oxidation number / valency (of Mg, Al, Si) [1] **[2]**
(d)(ii)	$MgCl_2(+aq) \rightarrow Mg^{2+} + 2Cl^-$ [1] $AlCl_3 + 6H_2O \rightarrow Al(H_2O)_6^{3+} + 3Cl^- / Al(H_2O)_5(OH)^{2+} + H^+ + 3Cl^-$ [1] $SiCl_4 + 2H_2O \rightarrow SiO_2 + 4H^+ + 4Cl^-$ [1] **[3]** **[Total:21]**

EXAMINER COMMENTS | MARK SCHEME

Select page

HIGH | MIDDLE | LOW

2 The elements in Group 17, the halogens, and their compounds, show many similarities and trends in their properties. Some data are given for the elements fluorine to iodine.

element	bond energy /kJ mol^{-1}	standard enthalpy change of atomisation, $\Delta H^{\ominus}_{at}$/kJ mol^{-1}	boiling point of element /K	boiling point of **hydrogen halide** /K
fluorine, F–F	158	79	85	293
chlorine, Cl–Cl	242	121	238	188
bromine, Br–Br	193	112	332	206
iodine, I–I	151	107	457	238

(a) (i) Explain the meaning of the term *standard enthalpy change of atomisation*.

The enthalpy change when one mole of gaseous atom is formed from its element under standard state conditions. [3]

(ii) For fluorine and chlorine, the enthalpy changes of atomisation are half the value of the bond energies.

For bromine and iodine, the enthalpy changes of atomisation are much more than half the value of the bond energies.

Suggest a reason for this difference.

Down the group, number of electrons increases, V.W.F's increases. So more energy is required for atomisation. [1]

(iii) The standard enthalpy of formation of iodine monochloride, ICl, is –24.0 kJ mol^{-1}.

Use this information and the bond energies of iodine and chlorine to calculate the I–Cl bond energy.

$I_2 + Cl_2 \xrightarrow{-24} 2ICl$

$151 + 242 \rightarrow I\text{-}Cl + (-24)$.

$393 + 24 \rightarrow I\text{-}Cl$

$I\text{-}Cl \rightarrow 417$.

$I_2 + Cl_2 \rightarrow 2ICl$

$151 + 242 \rightarrow I\text{-}Cl(-24)$

$I\text{-}Cl \rightarrow 417$

I–Cl bond energy = 417 kJ mol^{-1} [2]

Your Mark	
2(a)(i)	
2(a)(ii)	
2(a)(iii)	
2(b)(i)	
2(b)(ii)	
2(c)(i)	
2(c)(ii)	
2(c)(iii)	
2(c)(iv)	
2(d)(i)	
2(d)(ii)	

Q2	Mark scheme
(a)(i)	enthalpy / energy / heat change when one mole of gaseous atoms is produced [1] from the element in its standard state [1] under standard conditions [1] **[3]**
(a)(ii)	fluorine and chlorine are gases / bromine liquid and iodine solid OR as ΔH_{at} for bromine / iodine also includes changes of state **[1]**
(a)(iii)	$(\frac{1}{2}Cl_2 + \frac{1}{2}I_2 \rightarrow ICl)$ $\Delta H_f = (\frac{1}{2}E(Cl_2) + \frac{1}{2}E(I_2)) - E(ICl)$ OR $E(ICl) = (151 / 2) + (242 / 2) + 24$ [1] $E(ICl) = (+)\ 220.5 / 221$ [1] **[2]**
(b)(i)	stronger / more / greater id–id / London / dispersion forces [1] due to increasing numbers of electrons [1] **[2]**
(b)(ii)	(intermolecular forces in HF are) hydrogen bonds [1] (which are) stronger (than vdW) / more energy needed to separate molecules [1] OR HF much more polar / F much more electronegative [1] Intermolecular forces in HF stronger (than in HCl, HBr, HI) [1] **[2]**
(c)(i)	**P** = iodine / I_2 /I; **Q** = chlorine / Cl_2 / Cl **[1]**
(c)(ii)	weaker H–**P** than H–**Q** bond ORA / easier /less energy to break H–**P** than H–**Q** ORA [1] due to greater distance / shielding of nucleus from bond pair ORA [1] **[2]**
(c)(iii)	$2H\mathbf{P}$ (or 2HI) → (or) $H_2 + \mathbf{P}_2$ (or I_2) **[1]**
(c)(iv)	$Ag^+(aq) + \mathbf{Q}^-(aq)$ (or Cl^-) → $Ag\mathbf{Q}(s)$ (or AgCl(s)) [1] $Ag\mathbf{Q}(s) / AgCl(s) + NH_3(aq) \rightarrow Ag(NH_3)_2^+(aq) + \mathbf{Q}^-(aq) / Cl^-(aq)$ [1] **[2]**
(d)(i)	no of Cl increases by one each time / matches group number [1] due to increasing number of valence/outer(most/shell) electrons / oxidation number / valency (of Mg, Al, Si) [1] **[2]**
(d)(ii)	$MgCl_2(+aq) \rightarrow Mg^{2+} + 2Cl^-$ [1] $AlCl_3 + 6H_2O \rightarrow Al(H_2O)_6^{3+} + 3Cl^- / Al(H_2O)_5(OH)^{2+} + H^+ + 3Cl^-$ [1] $SiCl_4 + 2H_2O \rightarrow SiO_2 + 4H^+ + 4Cl^-$ [1] **[3]** **[Total:21]**

EXAMINER COMMENTS | MARK SCHEME

Select page

HIGH | MIDDLE | LOW

(b) (i) Explain the trend in the boiling points of the hydrogen halides, HCl, HBr and HI.

Boiling point increases from HCl to HI, as the number of electrons increases. Hence Vander Waal's forces increases and more energy is required to overcome these forces. [2]

(ii) Suggest why the hydrogen halide HF does not follow the trend in boiling points shown by HCl, HBr and HI.

H-F has a greater electronegativity difference between H and F, as compared to other hydrogen halides. [2]

(c) In an experiment, two of the halogens are represented as P_2 and Q_2.

P_2 combines with hydrogen on heating to form HP, which can be easily broken down into its elements. A solution of HP in water reacts with aqueous silver ions to form a yellow precipitate that is insoluble in dilute aqueous ammonia.

Q_2 combines explosively with hydrogen in sunlight to form HQ, which is stable to heat. A solution of HQ in water reacts with aqueous silver ions to form a white precipitate that is soluble in dilute aqueous ammonia.

(i) Identify the halogens P_2 and Q_2.

P_2 = Iodine Q_2 = chlorine [1]

(ii) HP readily decomposes into its elements when heated but HQ is stable to heat. Explain this with reference to bond energies.

HP is H-I which has the bond energy 299 $kJ\,mol^{-1}$ which is lower than bond energy of HCl (HCl) that is 431 $kJ\,mol^{-1}$. So, HP decomposes easily as less heat required. H-I has longer bond lengths than H-Cl. [2]

(iii) Write an equation for the thermal decomposition of HP.

$2HI \longrightarrow H_2 + I_2$ [1]

	Your Mark
2(a)(i)	
2(a)(ii)	
2(a)(iii)	
2(b)(i)	
2(b)(ii)	
2(c)(i)	
2(c)(ii)	
2(c)(iii)	
2(c)(iv)	
2(d)(i)	
2(d)(ii)	

Q2	Mark scheme
(a)(i)	enthalpy / energy / heat change when one mole of gaseous atoms is produced [1] from the element in its standard state [1] under standard conditions [1] **[3]**
(a)(ii)	fluorine and chlorine are gases / bromine liquid and iodine solid OR as ΔH_{at} for bromine / iodine also includes changes of state **[1]**
(a)(iii)	$(\frac{1}{2}Cl_2 + \frac{1}{2}I_2 \rightarrow ICl)$ $\Delta H_f = (\frac{1}{2}E(Cl_2) + \frac{1}{2}E(I_2)) - E(ICl)$ OR $E(ICl) = (151 / 2) + (242 / 2) + 24$ [1] $E(ICl) = (+)\ 220.5 / 221$ [1] **[2]**
(b)(i)	stronger / more / greater id–id / London / dispersion forces [1] due to increasing numbers of electrons [1] **[2]**
(b)(ii)	(intermolecular forces in HF are) hydrogen bonds [1] (which are) stronger (than vdW) / more energy needed to separate molecules [1] OR HF much more polar / F much more electronegative [1] Intermolecular forces in HF stronger (than in HCl, HBr, HI) [1] **[2]**
(c)(i)	**P** = iodine / I_2 /I; **Q** = chlorine / Cl_2 / Cl **[1]**
(c)(ii)	weaker H–**P** than H–**Q** bond ORA / easier /less energy to break H–**P** than H–**Q** ORA [1] due to greater distance / shielding of nucleus from bond pair ORA [1] **[2]**
(c)(iii)	2H**P** (or 2HI) → (or) H_2 + $\mathbf{P}_2$ (or I_2) **[1]**
(c)(iv)	$Ag^+(aq) + \mathbf{Q}^-(aq)$ (or Cl^-) → Ag**Q**(s) (or AgCl(s)) [1] Ag**Q**(s) / AgCl(s) + NH_3(aq) → $Ag(NH_3)_2^+(aq) + \mathbf{Q}^-(aq)$ / $Cl^-(aq)$ [1] **[2]**
(d)(i)	no of Cl increases by one each time / matches group number [1] due to increasing number of valence/outer(most/shell) electrons / oxidation number / valency (of Mg, Al, Si) [1] **[2]**
(d)(ii)	$MgCl_2$ (+aq) → $Mg^{2+} + 2Cl^-$ [1] $AlCl_3 + 6H_2O \rightarrow Al(H_2O)_6^{3+} + 3Cl^- / Al(H_2O)_5(OH)^{2+} + H^+ + 3Cl^-$ [1] $SiCl_4 + 2H_2O \rightarrow SiO_2 + 4H^+ + 4Cl^-$ [1] **[3]** **[Total:21]**

EXAMINER COMMENTS | MARK SCHEME

(iv) Write ionic equations, including state symbols, for

1. the formation of the white precipitate on addition of aqueous silver ions to aqueous HQ,

$Cl^-_{(aq)} + Ag^+_{(aq)} \rightarrow AgCl_{(s)}$

2. the subsequent dissolving of this precipitate in dilute aqueous ammonia.

$AgCl_{(s)} \rightarrow Ag^+_{(aq)} + Cl^-_{(aq)}$ [2]

(d) Chlorine reacts directly with many elements to form chlorides. Three such compounds are $MgCl_2$, $AlCl_3$ and $SiCl_4$.

(i) State and explain the pattern shown by the formulae of these three chlorides.

The number of chlorine atoms attached to the elements increases from Mg to Si because the charge ~~of~~ number increases from Mg(+2) to Al(+3) and Si(+4). [2]

(ii) Write equations to show the behaviour of each of these chlorides when added to water.

$MgCl_2$ $MgCl_2 + 2H_2O \rightarrow Mg(OH)_2 + 2HCl$

$AlCl_3$ $2AlCl_3 + 3H_2O \rightarrow Al_2O_3 + 6HCl$

$SiCl_4$ $SiCl_4 + 2H_2O \rightarrow SiO_2 + 4HCl$

[3]

[Total: 21]

Select page

HIGH MIDDLE LOW

	Your Mark
2(a)(i)	
2(a)(ii)	
2(a)(iii)	
2(b)(i)	
2(b)(ii)	
2(c)(i)	
2(c)(ii)	
2(c)(iii)	
2(c)(iv)	
2(d)(i)	
2(d)(ii)	

Q2	Mark scheme
(a)(i)	enthalpy / energy / heat change when one mole of gaseous atoms is produced [1] from the element in its standard state [1] under standard conditions [1] **[3]**
(a)(ii)	fluorine and chlorine are gases / bromine liquid and iodine solid OR as ΔH_{at} for bromine / iodine also includes changes of state **[1]**
(a)(iii)	$(\frac{1}{2}Cl_2 + \frac{1}{2}I_2 \rightarrow ICl)$ $\Delta H_f = (\frac{1}{2}E(Cl_2) + \frac{1}{2}E(I_2)) - E(ICl)$ OR $E(ICl) = (151 / 2) + (242 / 2) + 24$ [1] $E(ICl) = (+) 220.5 / 221$ [1] **[2]**
(b)(i)	stronger / more / greater id–id / London / dispersion forces [1] due to increasing numbers of electrons [1] **[2]**
(b)(ii)	(intermolecular forces in HF are) hydrogen bonds [1] (which are) stronger (than vdW) / more energy needed to separate molecules [1] OR HF much more polar / F much more electronegative [1] Intermolecular forces in HF stronger (than in HCl, HBr, HI) [1] **[2]**
(c)(i)	**P** = iodine / I_2 /I; **Q** = chlorine / Cl_2 / Cl **[1]**
(c)(ii)	weaker H–**P** than H–**Q** bond ORA / easier /less energy to break H–**P** than H–**Q** ORA [1] due to greater distance / shielding of nucleus from bond pair ORA [1] **[2]**
(c)(iii)	2H**P** (or 2HI) → (or) H_2 + $\mathbf{P}_2$ (or I_2) **[1]**
(c)(iv)	$Ag^+(aq) + \mathbf{Q}^-(aq)$ (or Cl^-) → Ag**Q**(s) (or AgCl(s)) [1] Ag**Q**(s) / AgCl(s) + NH_3(aq) → $Ag(NH_3)_2^+(aq) + \mathbf{Q}^-(aq)$ / $Cl^-(aq)$ [1] **[2]**
(d)(i)	no of Cl increases by one each time / matches group number [1] due to increasing number of valence/outer(most/shell) electrons / oxidation number / valency (of Mg, Al, Si) [1] **[2]**
(d)(ii)	$MgCl_2(+aq) \rightarrow Mg^{2+} + 2Cl^-$ [1] $AlCl_3 + 6H_2O \rightarrow Al(H_2O)_6^{3+} + 3Cl^- / Al(H_2O)_5(OH)^{2+} + H^+ + 3Cl^-$ [1] $SiCl_4 + 2H_2O \rightarrow SiO_2 + 4H^+ + 4Cl^-$ [1] **[3]** **[Total:21]**

EXAMINER COMMENTS MARK SCHEME

Select page

HIGH | MIDDLE | LOW

2 The elements in Group 17, the halogens, and their compounds, show many similarities and trends in their properties. Some data are given for the elements fluorine to iodine.

element	bond energy /kJ mol^{-1}	standard enthalpy change of atomisation, $\Delta H_{at}^{\ominus}$ /kJ mol^{-1}	boiling point of element /K	boiling point of **hydrogen halide** /K
fluorine, F–F	158	79	85	293
chlorine, Cl–Cl	242	121	238	188
bromine, Br–Br	193	112	332	206
iodine, I–I	151	107	457	238

(a) (i) Explain the meaning of the term *standard enthalpy change of atomisation*.

The enthalpy change needed when 1 mole of an atom is converted to its gaseous state under standard conditions [3]

(ii) For fluorine and chlorine, the enthalpy changes of atomisation are half the value of the bond energies.

For bromine and iodine, the enthalpy changes of atomisation are much more than half the value of the bond energies.

Suggest a reason for this difference.

Fluorine and Chlorine have low boiling points so similar energy is needed for atomization. They are reactive and have weak vanderwall forces between them [1]

(iii) The standard enthalpy of formation of iodine monochloride, ICl, is –24.0 kJ mol^{-1}.

Use this information and the bond energies of iodine and chlorine to calculate the I–Cl bond energy.

$Cl_2 + I_2 \longrightarrow 2ICl$

$(242 + 151)$

$\Delta H_1 - \Delta H_2 = -24$

$393 + 24 = 2x$

$x = 208.5$

I–Cl bond energy = 208.5 kJ mol^{-1} [2]

Your Mark

2(a)(i) 2(a)(ii) 2(a)(iii) 2(b)(i) 2(b)(ii) 2(c)(i) 2(c)(ii) 2(c)(iii) 2(c)(iv) 2(d)(i) 2(d)(ii)

Q2	Mark scheme	
(a)(i)	enthalpy / energy / heat change when one mole of gaseous atoms is produced from the element in its standard state under standard conditions	[1] [1] [1] **[3]**
(a)(ii)	fluorine and chlorine are gases / bromine liquid and iodine solid OR as ΔH_{at} for bromine / iodine also includes changes of state	**[1]**
(a)(iii)	$(\frac{1}{2}Cl_2 + \frac{1}{2}I_2 \rightarrow ICl)$ $\Delta H_f = (\frac{1}{2}E(Cl_2) + \frac{1}{2}E(I_2)) - E(ICl)$ OR $E(ICl) = (151 / 2) + (242 / 2) + 24$ [1] $E(ICl) = (+)\ 220.5 / 221$	[1] **[2]**
(b)(i)	stronger / more / greater id–id / London / dispersion forces due to increasing numbers of electrons	[1] [1] **[2]**
(b)(ii)	(intermolecular forces in HF are) hydrogen bonds (which are) stronger (than vdW) / more energy needed to separate molecules OR HF much more polar / F much more electronegative Intermolecular forces in HF stronger (than in HCl, HBr, HI)	[1] [1] [1] [1] **[2]**
(c)(i)	**P** = iodine / I_2 /I; **Q** = chlorine / Cl_2 / Cl	**[1]**
(c)(ii)	weaker H–**P** than H–**Q** bond ORA / easier /less energy to break H–**P** than H–**Q** ORA due to greater distance / shielding of nucleus from bond pair ORA	[1] [1] **[2]**
(c)(iii)	2H**P** (or 2HI) → (or) H_2 + $\mathbf{P}_2$ (or I_2)	**[1]**
(c)(iv)	$Ag^+(aq) + \mathbf{Q}^-(aq)$ (or Cl^-) → Ag**Q**(s) (or AgCl(s)) Ag**Q**(s) / AgCl(s) + NH_3(aq) → $Ag(NH_3)_2^+(aq) + \mathbf{Q}^-(aq) / Cl^-(aq)$	[1] [1] **[2]**
(d)(i)	no of Cl increases by one each time / matches group number due to increasing number of valence/outer(most/shell) electrons / oxidation number / valency (of Mg, Al, Si)	[1] [1] **[2]**
(d)(ii)	$MgCl_2$ (+aq) → $Mg^{2+} + 2Cl^-$ $AlCl_3 + 6H_2O \rightarrow Al(H_2O)_6^{3+} + 3Cl^- / Al(H_2O)_5(OH)^{2+} + H^+ + 3Cl^-$ $SiCl_4 + 2H_2O \rightarrow SiO_2 + 4H^+ + 4Cl^-$	[1] [1] [1] **[3]**
		[Total:21]

EXAMINER COMMENTS | MARK SCHEME

Select page

HIGH | MIDDLE | LOW

(b) (i) Explain the trend in the boiling points of the hydrogen halides, HCl, HBr and HI.

Boiling points generally ~~decrease~~ increase . HI > HBr > HCl

Boiling points decrease . HCl > HBr > HI . This is due to more energy to break bonds [2]

(ii) Suggest why the hydrogen halide HF does not follow the trend in boiling points shown by HCl, HBr and HI.

HF is polar and has strong electronegitivity so more energy is needed to break the bond. [2]

(c) In an experiment, two of the halogens are represented as P_2 and Q_2.

P_2 combines with hydrogen on heating to form HP, which can be easily broken down into its elements. A solution of HP in water reacts with aqueous silver ions to form a yellow precipitate that is insoluble in dilute aqueous ammonia.

Q_2 combines explosively with hydrogen in sunlight to form HQ, which is stable to heat. A solution of HQ in water reacts with aqueous silver ions to form a white precipitate that is soluble in dilute aqueous ammonia.

(i) Identify the halogens P_2 and Q_2.

P_2 = Iodine (I_2) Q_2 = ~~Cl~~ Chlorine (Cl_2) [1]

(ii) HP readily decomposes into its elements when heated but HQ is stable to heat. Explain this with reference to bond energies.

More energy is needed to break the H–Q bond . Less is needed for H–P so it easily breaks into its elements [2]

(iii) Write an equation for the thermal decomposition of HP.

~~2HP heat P2 + H2~~ $2HI \xrightarrow{heat} H_2 + I_2$ [1]

	Your Mark
2(a)(i)	
2(a)(ii)	
2(a)(iii)	
2(b)(i)	
2(b)(ii)	
2(c)(i)	
2(c)(ii)	
2(c)(iii)	
2(c)(iv)	
2(d)(i)	
2(d)(ii)	

Q2	Mark scheme
(a)(i)	enthalpy / energy / heat change when one mole of gaseous atoms is produced [1] from the element in its standard state [1] under standard conditions [1] **[3]**
(a)(ii)	fluorine and chlorine are gases / bromine liquid and iodine solid OR as ΔH_{at} for bromine / iodine also includes changes of state **[1]**
(a)(iii)	$(\frac{1}{2}Cl_2 + \frac{1}{2}I_2 \rightarrow ICl)$ $\Delta H_f = (\frac{1}{2}E(Cl_2) + \frac{1}{2}E(I_2)) - E(ICl)$ OR $E(ICl) = (151 / 2) + (242 / 2) + 24$ [1] $E(ICl) = (+) 220.5 / 221$ [1] **[2]**
(b)(i)	stronger / more / greater id–id / London / dispersion forces [1] due to increasing numbers of electrons [1] **[2]**
(b)(ii)	(intermolecular forces in HF are) hydrogen bonds [1] (which are) stronger (than vdW) / more energy needed to separate molecules [1] OR HF much more polar / F much more electronegative [1] Intermolecular forces in HF stronger (than in HCl, HBr, HI) [1] **[2]**
(c)(i)	**P** = iodine / I_2 /I; **Q** = chlorine / Cl_2 / Cl **[1]**
(c)(ii)	weaker H–**P** than H–**Q** bond ORA / easier /less energy to break H–**P** than H–**Q** ORA [1] due to greater distance / shielding of nucleus from bond pair ORA [1] **[2]**
(c)(iii)	2H**P** (or 2HI) → (or) H_2 + $\mathbf{P}_2$ (or I_2) **[1]**
(c)(iv)	$Ag^+(aq) + \mathbf{Q}^-(aq)$ (or Cl^-) → Ag**Q**(s) (or AgCl(s)) [1] Ag**Q**(s) / AgCl(s) + NH_3(aq) → $Ag(NH_3)_2^+(aq) + \mathbf{Q}^-(aq)$ / $Cl^-(aq)$ [1] **[2]**
(d)(i)	no of Cl increases by one each time / matches group number [1] due to increasing number of valence/outer(most/shell) electrons / oxidation number / valency (of Mg, Al , Si) [1] **[2]**
(d)(ii)	$MgCl_2$ (+aq) → $Mg^{2+} + 2Cl^-$ [1] $AlCl_3 + 6H_2O \rightarrow Al(H_2O)_6^{3+} + 3Cl^- / Al(H_2O)_5(OH)^{2+} + H^+ + 3Cl^-$ [1] $SiCl_4 + 2H_2O \rightarrow SiO_2 + 4H^+ + 4Cl^-$ [1] **[3]** **[Total:21]**

EXAMINER COMMENTS | MARK SCHEME

Select page

HIGH | MIDDLE | LOW

(iv) Write ionic equations, including state symbols, for $Ag^+_{(aq)} + Cl^-_{(aq)} \rightarrow AgCl_{(s)}$ $NH_3 + AgCl + H^+ \rightarrow$

1. the formation of the white precipitate on addition of aqueous silver ions to aqueous HQ.

$Ag^+_{(aq)} + Cl^-_{(aq)} \longrightarrow AgCl_{(s)}$

2. the subsequent dissolving of this precipitate in dilute aqueous ammonia.

$NH_3(aq) + Cl(s) \longrightarrow NH_4Cl(aq)$

[2]

(d) Chlorine reacts directly with many elements to form chlorides. Three such compounds are $MgCl_2$, $AlCl_3$ and $SiCl_4$.

(i) State and explain the pattern shown by the formulae of these three chlorides.

$MgCl_2$ is ionic bond, Mg transfered 1 electron to each Cl atom.

$AlCl_3$ is bonded by sharing of electron to each electron.

$SiCl_4$ is giant covalent structure. Each Cl is covalently bonded [2]

(ii) Write equations to show the behaviour of each of these chlorides when added to water.

$MgCl_2$ $MgCl_2 + 2H_2O \longrightarrow Mg(OH)_2 + 2HCl$

$AlCl_3$ $AlCl_3 + 3H_2O \longrightarrow Al(OH)_3 + 3HCl$

$SiCl_4$ $SiCl_4 + 4H_2O \longrightarrow Si(OH)_4 + 4HCl$

[3]

$AlCl_3 + 3H_2O \rightarrow Al(OH)_3 + 3HCl$

$SiCl_4 + 4H_2O \rightarrow Si(OH)_4 + 4HCl$

[Total: 21]

	Your Mark
2(a)(i)	
2(a)(ii)	
2(a)(iii)	
2(b)(i)	
2(b)(ii)	
2(c)(i)	
2(c)(ii)	
2(c)(iii)	
2(c)(iv)	
2(d)(i)	
2(d)(ii)	

Q2	Mark scheme
(a)(i)	enthalpy / energy / heat change when one mole of gaseous atoms is produced [1] from the element in its standard state [1] under standard conditions [1] **[3]**
(a)(ii)	fluorine and chlorine are gases / bromine liquid and iodine solid OR as ΔH_{at} for bromine / iodine also includes changes of state **[1]**
(a)(iii)	$(\frac{1}{2}Cl_2 + \frac{1}{2}I_2 \rightarrow ICl)$ $\Delta H_f = (\frac{1}{2}E(Cl_2) + \frac{1}{2}E(I_2)) - E(ICl)$ OR $E(ICl) = (151 / 2) + (242 / 2) + 24$ [1] $E(ICl) = (+) 220.5 / 221$ [1] **[2]**
(b)(i)	stronger / more / greater id–id / London / dispersion forces [1] due to increasing numbers of electrons [1] **[2]**
(b)(ii)	(intermolecular forces in HF are) hydrogen bonds [1] (which are) stronger (than vdW) / more energy needed to separate molecules [1] OR HF much more polar / F much more electronegative [1] Intermolecular forces in HF stronger (than in HCl, HBr, HI) [1] **[2]**
(c)(i)	**P** = iodine / I_2 /I; **Q** = chlorine / Cl_2 / Cl **[1]**
(c)(ii)	weaker H–**P** than H–**Q** bond ORA / easier /less energy to break H–**P** than H–**Q** ORA [1] due to greater distance / shielding of nucleus from bond pair ORA [1] **[2]**
(c)(iii)	2H**P** (or 2HI) → (or) H_2 + $\mathbf{P}_2$ (or I_2) **[1]**
(c)(iv)	$Ag^+(aq) + \mathbf{Q}^-(aq)$ (or Cl^-) → Ag**Q**(s) (or AgCl(s)) [1] Ag**Q**(s) / AgCl(s) + $NH_3(aq) \rightarrow Ag(NH_3)_2^+(aq) + \mathbf{Q}^-(aq)$ / $Cl^-(aq)$ [1] **[2]**
(d)(i)	no of Cl increases by one each time / matches group number [1] due to increasing number of valence/outer(most/shell) electrons / oxidation number / valency (of Mg, Al, Si) [1] **[2]**
(d)(ii)	$MgCl_2(+aq) \rightarrow Mg^{2+} + 2Cl^-$ [1] $AlCl_3 + 6H_2O \rightarrow Al(H_2O)_6^{3+} + 3Cl^- / Al(H_2O)_5(OH)^{2+} + H^+ + 3Cl^-$ [1] $SiCl_4 + 2H_2O \rightarrow SiO_2 + 4H^+ + 4Cl^-$ [1] **[3]** **[Total:21]**

EXAMINER COMMENTS | MARK SCHEME

Cambridge Assessment International Education
The Triangle Building, Shaftesbury Road, Cambridge, CB2 8EA, United Kingdom
t: +44 1223 553554 f: +44 1223 553558
e: info@cambridgeinternational.org www.cambridgeinternational.org