forming solid, ionic compound from the respective gaseous ions → ΔH_{LE}

1. The lattice enthalpy of calcium chloride can be calculated using **three** of the enthalpy changes below.

using Born-Haber cycles.

Which enthalpy change is **not** required?

$\Delta H_{sol.}$ **A** enthalpy change of solution of calcium chloride

$\Delta H_{hyd.}$ **B** enthalpy change of hydration of Cl^- ions

ΔH_f **C** enthalpy change of formation of calcium chloride

$\Delta H_{hyd.}$ **D** enthalpy change of hydration of Ca^{2+} ions

$\Delta H_{hyd.}$ $(Ca^{2+} + 2Cl^-)$

$Ca^{2+}_{(g)} + 2Cl^-_{(g)}$

$\downarrow \Delta H_{LE}$

$CaCl_{2(s)}$

$\downarrow \Delta H_{sol.}$

$CaCl_{2(aq)}$

used for a larger Born-Haber cycle.

Your answer C

[1]

2. This question is about the chemistry of the elements in Group 2 and the halogens.

(a) A student prepares an aqueous solution of magnesium chloride by reacting magnesium with excess hydrochloric acid.

$MgCl_2$ Mg HCl

Write an equation, including state symbols, for this reaction and state the observation(s) the student should make whilst carrying out this experiment.

equation: $Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$

observation(s): effervescence and solid dissolves

[2]

(b) Lattice enthalpies give an indication of the strength of ionic bonding.

How would the lattice enthalpies of magnesium chloride and calcium chloride differ?

Explain your answer.

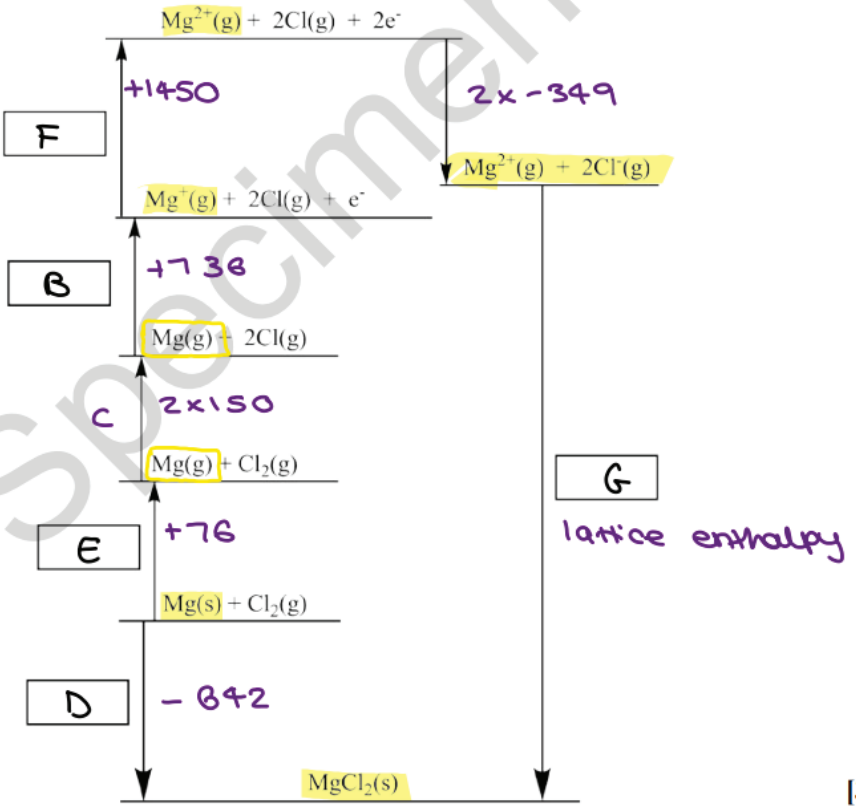
lattice enthalpy of $MgCl_2$ is more exothermic than $CaCl_2$ because Mg^{2+} is smaller than Ca^{2+} therefore, the attraction between Mg^{2+} and Cl^- is greater

[3]

(c) The table below shows the enthalpy changes that are needed to determine the lattice enthalpy of magnesium chloride, $MgCl_2$.

Letter	Enthalpy change	Energy / kJ mol^{-1}
A	1st electron affinity of chlorine	–349
B	1st ionisation energy of magnesium	+736
C	atomisation of chlorine	+150
D	formation of magnesium chloride	–642
E	atomisation of magnesium	+76
F	2nd ionisation energy of magnesium	+1450
G	lattice enthalpy of magnesium chloride	

(i) On the cycle below, write the correct letter in each box.



[3]

(ii) Use the Born–Haber cycle to calculate the lattice enthalpy of magnesium chloride.

$-(2\times-349)-1450-736-(2\times150)-76-642$

$= -2506$ kJmol^{-1}

lattice enthalpy = ...−2506..... kJ mol^{-1} **[2]**

(d)* Describe and explain the relative reactivity of the halogens, chlorine, bromine and iodine, in their redox reactions with halides, using reactions on a test-tube scale.

Include reaction equations and observations in your answer.

$Cl_2 + 2Br^- \longrightarrow Br_2 + 2Cl^-$ yellow solution

$Cl_2 + 2I^- \longrightarrow I_2 + 2Cl^-$ orange / brown solution

Down the group:

- more shells
- increased shielding
- more difficult to gain an electron

[6]

3. Which enthalpy change(s) is/are endothermic? +ve

1 The bond enthalpy of the C–H bond ✓ always +ve

2 The second electron affinity of oxygen ✓ 2nd = +ve, 1st = -ve

$O^-(g) + e^- \rightarrow O^{2-}(g)$

~~3~~ The standard enthalpy change of formation of magnesium = 0

A 1, 2 and 3

B Only 1 and 2

C Only 2 and 3

D Only 1

Your answer B

[1]

4. This question is about enthalpy changes.

(a) **Table 16.1** shows enthalpy changes that can be used to determine the enthalpy change of hydration of fluoride ions, F^-.

aqueous/hydrated ions

Enthalpy change	Energy / kJ mol^{-1}
Hydration of Ca^{2+}	–1609
Solution of CaF_2	+13
Lattice enthalpy of CaF_2	–2630

Table 16.1

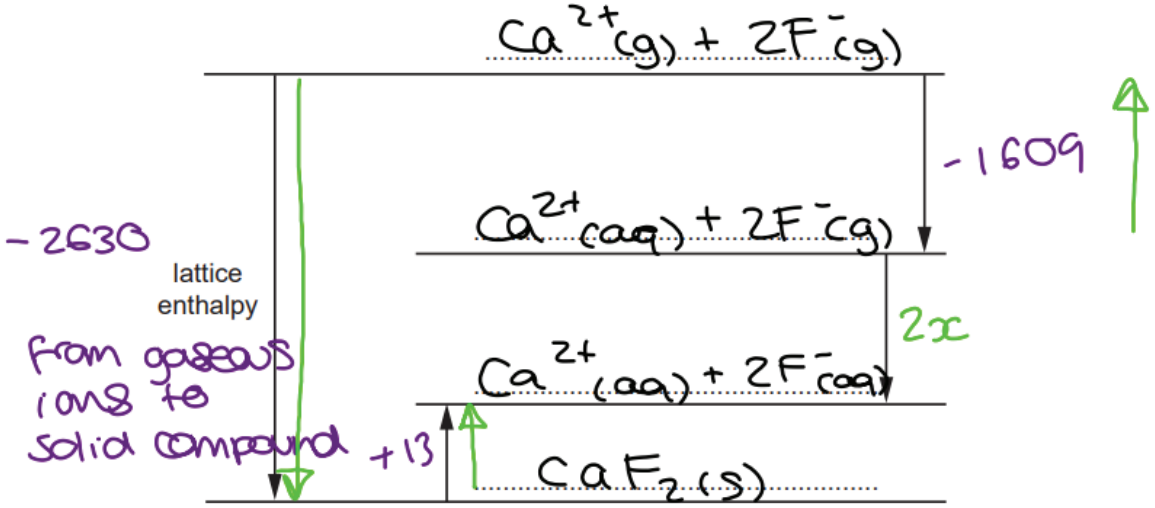
(i) Explain what is meant by the term *enthalpy change of hydration*.

1 mole of gaseous ions react to form 1 mole of aqueous / hydrated ions

[2]

(ii) The enthalpy change of hydration of F^- can be determined using the enthalpy changes in **Table 16.1** and the incomplete energy cycle below.

On the dotted lines, add the species present, including state symbols.



[4]

(iii) Calculate the enthalpy change of hydration of fluoride ions, F^-.

$2x = +1609 - 2630 + 13$

$2x = -1008$

$x = -504$

enthalpy change of hydration = −504 kJ mol^{-1} **[2]**

(iv) Predict how the enthalpy changes of hydration of F^- and Cl^- would differ.

Explain your answer.

$\Delta_{hyd}H$ of F^- is more exothermic because of F^-'s smaller size meaning greater attraction to H_2O. **[2]**

(b) Fluorine reacts with steam as shown in the equation below.

$$2F_2(g) + 2H_2O(g) \rightarrow O_2(g) + 4HF(g) \qquad \Delta H = -598\,kJ\,mol^{-1}$$

Average bond enthalpies are shown in the table. reactants - products = ΔH, bond enthalpies of

Bond	Average bond enthalpy / kJ mol^{-1}
O–H	+464
O=O	+498
H–F	+568

recall - learn this definition

(i) Explain what is meant by the term *average bond enthalpy*.

The breaking of 1 mole of bonds in gaseous molecules

[2]

(ii) Calculate the bond enthalpy of the F–F bond.

$(2x + 2(2 \times 464)) - (498 + (4 \times 568)) = -598$

$(2x + 1856) - 2770 = -598$

$2x + 1856 = 2172$

$2x = 316$

$x = +158$

bond enthalpy = +158 kJ mol^{-1} **[3]**