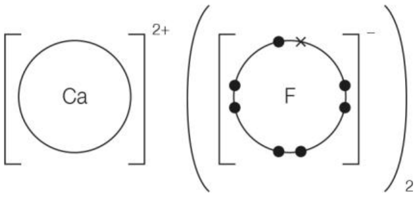
## Pages 65–66

**1 a) i)** Calcium changes from 0 to the + 2 state: oxidation. [1]

Fluorine changes from 0 to the –1 state: reduction. [1]

**ii)**




[2] (*one mark for each ion – no circles needed*)

**b) i)** Properties such as: relatively high melting points, conduct electricity when molten but not when solid. (*2 marks for any two*)

**ii)** Giant lattice of oppositely charged ions. [1] Strong attraction between oppositely charged ions accounts for the high melting points. [1] Ions can only acts as charge carriers and conduct electricity if the ions are free to move. [1]

**2 a)** The first electron affinity of an element is the energy change when one mole of gaseous atoms [1] gain electrons to form one mole of gaseous ions, each with a single negative charge. [1]

**b) i)** The second electron affinity involves the addition of a negative electron to an ion with one negative charge. [1] This requires energy to overcome the repulsion. [1] So the process is endothermic.

**ii)** $N(g) + e^- \rightarrow N^-(g)$ (1

$N^{2-}(g) + e^- \rightarrow N^{3-}(g)$ [1]

**c) i)** Polarisation of an ion is the distortion of the electron cloud [1] around the ion by a nearby positive charge. [1]

**ii)** The large charge density in the small $Mg^{2+}$ ion. [1]

The large, highly polarisable $I^-$ ion. [1]

**3 a) i)** A is Ag(g) [1]; B is $Ag^+(g)$ [1]; C is 2F(g) [1]

**ii)** D is the second ionisation energy of silver. [1]

E is the 2 × electron affinity of fluorine. [1]

F is the lattice energy of silver(II) fluoride [1]

**b)** Lattice energy is the energy change when one mole of an ionic solid [1] forms from free gaseous ions. [1]

**c)** $\Delta_f H^\ominus [AgF_2(s)] = +289\ kJ\ mol^{-1} + 732\ kJ\ mol^{-1}$

$+ 2070\ kJ\ mol^{-1}$

$+ (2 \times +79\ kJ\ mol^{-1})$

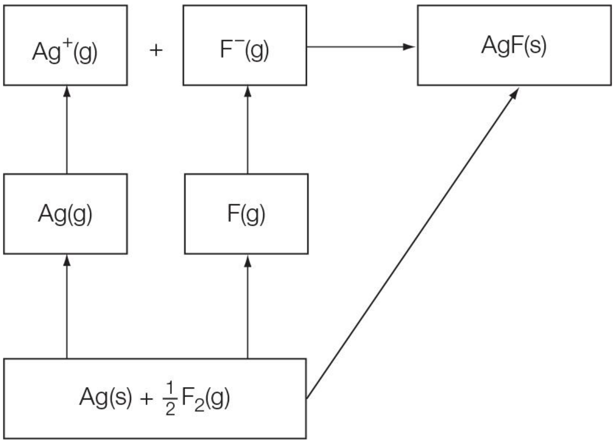
$+ (2 \times -348\ kJ\ mol^{-1})$

$+ (-2650)\ kJ\ mol^{-1}$ [1]

$= -97\ kJ\ mol^{-1}$

[1] for 97, [1] for sign and units

**d)**




$\Delta_f H^\ominus [AgF(s)] = +289\ kJ\ mol^{-1} + 732\ kJ\ mol^{-1}$

$+79\ kJ\ mol^{-1} - 348\ kJ\ mol^{-1} - 955\ kJ\ mol^{-1}$ [1]

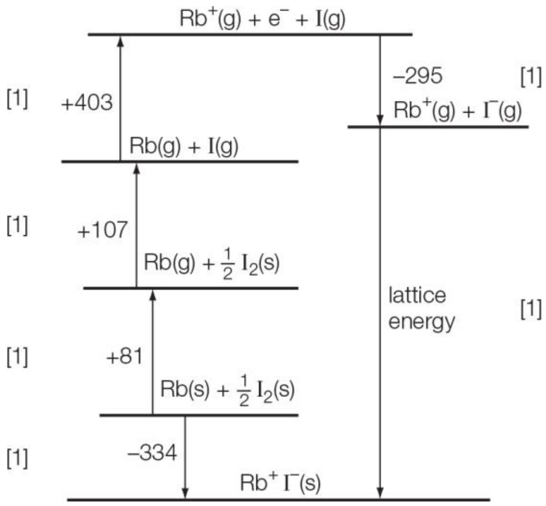
$= -203\ kJ\ mol^{-1}$ [1] for 203, [1] for sign and units

**e) i)** $\Delta_r H^\ominus = -203\ kJ\ mol^{-1} - (-97\ kJ\ mol^{-1})$

$= -106\ kJ\ mol^{-1}$ [1] for 106, [1] for sign

**ii)** $AgF_2$ has an exothermic $\Delta_f H^\ominus$ so is stable with respect to its elements [1], but is unstable with respect to AgF(s) so tends to decompose to form AgF(s) and ½$F_2$(g). [1]

**4 a)**




**b)** $\Delta H^\ominus_{lattice}[RbI(s)] = -(-295\ kJ\ mol^{-1}) - 403\ kJ\ mol^{-1} - 107\ kJ\ mol^{-1} - 81\ kJ\ mol^{-1} - 334\ kJ\ mol^{-1}$

$= -630\ kJ\ mol^{-1}$ [1] for 630, [1] for sign and units

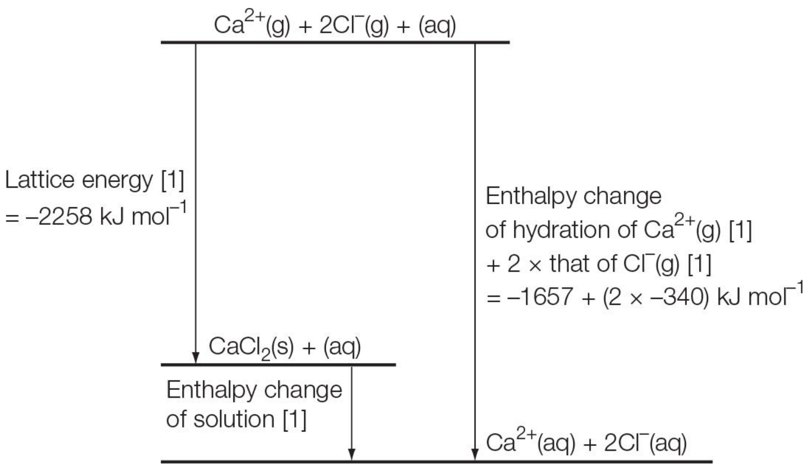
**c)** The $Li^+$ ion is much smaller than the $Rb^+$ ion [1] and can get closer to $I^-$ ions in the solid ionic crystal. [1] This results in a more exothermic process when LiI forms. [1]

**d)** This question assesses a student's ability to show a coherent and logically structured answer with linkages and fully sustained line of reasoning. Assess the quality of the answer taking into account both the key points made (*up to 4 marks*) and the logic and coherence of the discussion (*up to 2 marks*).

Points to make in the answer:

- $Li^+$ ions are smaller that $Rb^+$ ions because they have fewer inner full shells.
- This means that $Li^+$ ions have a higher charge density than $Rb^+$ ions.
- The $Li^+$ ions therefore cause greater polarisation of the $I^-$ ions.
- This distorts the negative charge cloud around the $I^-$ ions, pulling the electrons towards the $Li^+$ ions.
- This gives rise to a degree of electron sharing (covalent bonding) …
- … which is not accounted for by the ionic model.

**5 a)**




**b)** $\Delta_{sol}H[CaCl_2(s)]$ $= -(-2258 \text{ kJ mol}^{-1}) + (-1657 \text{ kJ mol}^{-1}) + (2 \times -340 \text{ kJ mol}^{-1})$ [1]

$= +2258 \text{ kJ mol}^{-1} - 1657 \text{ kJ mol}^{-1} - 680 \text{ kJ mol}^{-1}$

$= -79 \text{ kJ mol}^{-1}$ [1] for 79, [1] for sign and units

**c)** The $Ca^{2+}$ ion has a higher charge and a larger radius than $Li^+$. The higher charge tends to make its electron density greater than $Li^+$ [1] but its larger radius tends to make its electron density smaller than $Li^+$. [1]

**d)** Water molecules are polar with a δ+ charge between the H atoms and a δ– charge on the O atom. [1] The δ+ charge attracts anions and the δ– attracts cations. [1] So the hydration enthalpies of both anions and cations are exothermic/negative.

**6** **a)** **i)** B: $1s^2 2s^2 2p^1$ [1]

**ii)** The first electron is shielded from the 5+ nuclear charge by the two electrons in the first shell and partially by the two 2s electrons. It is easiest to remove. [1] The next two electrons are shielded by the two 2s electrons so that they are removed against an effective nuclear charge of 3+. [1] There are no electrons shielding the last two electrons, which are closes to the nucleus and have to be removed from the full 5+ nuclear charge, and so there is a big jump in the ionisation energies once electrons are removed from the inner shell. [1]

**b)** The following presented as a calculation, energy cycle or energy level diagram [1]:

- energy needed to atomise two B atoms = 2 × +590 kJ $mol^{-1}$ = +1180 kJ $mol^{-1}$ [1]
- energy needed to ionise two B(g) atoms = 2 × (800 + 2400 + 3700) kJ $mol^{-1}$ = 13 800 kJ $mol^{-1}$ [1]
- energy needed to atomise three O atoms = 3 × +250 kJ $mol^{-1}$ = +750 kJ $mol^{-1}$ [1]
- energy needed to turn three O(g) atoms into ions = 3 × (–140 + 790) kJ $mol^{-1}$ = 1950 kJ $mol^{-1}$ [1]
- lattice energy = –1270 kJ $mol^{-1}$ – (1180 + 13 800 + 750 + 1950) kJ $mol^{-1}$
  = –18 950 kJ $mol^{-1}$ [1]

**c)** Boron is a first row element and so its atom and ion are relatively small. The ion has a 3+ charge. So the polarising power of the ion is relatively large. [1] This suggests significant deviation from pure ionic bonding in its compounds. [1] So probably poor agreement between the calculated value and a theoretical value based on the ionic model.

**d)** **i)** $B_2O_3(s) + 3H_2O(l) \rightarrow H_3BO_3(aq)$ [1]

**ii)** Formation of an acidic oxide is characteristic of the covalent, molecular oxides of non-metal elements. [1] This property of  boron oxide is consistent with the idea that the $B^{3+}$ ion means that there is extensive covalent bonding in the oxide. [1]

**e)** The data shows that the chloride is likely to be a gas a room temperature. [1] This is characteristic of covalent, molecular chlorides of non-metals. [1] Such chlorides are hydrolysed by water to give hydrogen chloride and an oxoacid. [1]