InkChemistry

ANSWER BOOK

Lattice enthalpy and Born–Haber cycles

Thermodynamics · Unit C08

AQA 3.1.8.1 · OCR A 5.2.1

17 questions · 53 marks

NAVY EDITION

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SECTION A · Short, 1 to 2 marks

A1 LE-Q1

B1 the enthalpy change when one mole of a solid ionic compound is formed. **B1** from its gaseous ions, under standard conditions; reject 'from its elements', which describes the enthalpy of formation. [2]

A2 LE-Q3

B1 $\frac{1}{2}Cl_2(g) \rightarrow Cl(g)$, one mole of chlorine atoms formed; reject $Cl_2(g) \rightarrow 2Cl(g)$, which forms two moles. **B1** $Ca^+(g) \rightarrow Ca^{2+}(g) + e^-$, starting from the gaseous 1+ ion; reject $Ca(g) \rightarrow Ca^{2+}(g) + 2e^-$. [2]

A3 LE-Q22

B1 $2Ca(NO_3)_2(s) \rightarrow 2CaO(s) + 4NO_2(g) + O_2(g)$, balanced; allow multiples. **B1** $2KNO_3(s) \rightarrow 2KNO_2(s) + O_2(g)$, the nitrite and oxygen only; reject any equation giving NO_2. [2]

SECTION B · Standard, 3 to 5 marks

B1 LE-Q4

B1 the second electron is removed from an Mg^+ ion, which is already positively charged. **B1** the same nuclear charge acts on one fewer electron, so the remaining electrons are held more strongly and more energy is needed to remove one. **B1** the first electron affinity is exothermic because the incoming electron is attracted by the oxygen nucleus. **B1** the second electron is added to an O^- ion, which is already negative, so energy is needed to overcome the repulsion; reject 'oxygen has a full outer shell'. [4]

B2 LE-Q12

B1 the experimental value is more exothermic than the theoretical value because the bonding has some covalent character. **B1** the Ag^+ ion polarises the halide ion, distorting its electron cloud towards the cation. **B1** the halide ions get larger from F^- to I^-, with outer electrons further from the nucleus and more shielded. **B1** so the larger ions are more easily polarised, the covalent character increases and the difference grows; reject 'the bonding becomes covalent'. [4]

B3 LE-Q15

B1 the lattice enthalpy of CaO is almost five times as exothermic as that of KCl. **B1** Ca^{2+} and O^{2-} carry twice the charges of K^+ and Cl^-, so the attraction between a pair of ions is four times as strong at the same distance. **B1** Ca^{2+} is smaller than K^+ and O^{2-} smaller than Cl^-, so the ions in CaO sit closer together and attract more strongly. **B1** both effects make the attraction stronger, and together they account for a factor larger than four; ignore any reference to the enthalpies of formation. [4]

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B4 LE-Q16 — [4]

B1 the same cycle method gives sodium chloride within 3% of the theoretical value, so the method itself is sound. **B1** the silver halide differences are systematic: always in the same direction and growing from fluoride to iodide, which random error would not produce. **B1** the difference arises because the perfect ionic model ignores polarisation of the anion by Ag^+, so it underestimates the attraction in a lattice with covalent character. **B1** the statement is therefore wrong: the model, not the measurement, is the source of the difference.

B5 LE-Q20 — [4]

B1 both decompose to the metal oxide and carbon dioxide, $MCO_3(s) \rightarrow MO(s) + CO_2(g)$. **B1** Mg^{2+} is smaller than Ba^{2+} with the same charge, so it has a higher charge density and a greater polarising power. **B1** Mg^{2+} distorts the electron cloud of the large carbonate ion more, drawing electron density towards itself. **B1** this weakens a C–O bond in the carbonate ion, so less energy is needed to break it and release CO_2; reject 'the Mg–O bond is weaker'.

B6 LE-Q2 — [3]

B1 the enthalpy change when one mole of gaseous atoms each gain one electron. **B1** to form one mole of gaseous 1- ions; accept 'negative ions' for '1- ions'. **B1** $O(g) + e^- \rightarrow O^-(g)$, with every state symbol shown; allow e for e^-.

B7 LE-Q11 — [3]

B1 each ion is treated as a perfect sphere. **B1** with its charge evenly distributed, so there is no polarisation and no covalent character. **B1** the lattice enthalpy is calculated from the electrostatic attractions and repulsions between the ions, using their charges and the distances between their centres.

B8 LE-Q13 — [3]

B1 Cs^+ is a large ion with a single charge, so it has a low charge density and is only weakly polarising. **B1** F^- is a small anion whose electron cloud is held tightly and is hard to distort. **B1** there is very little polarisation, so the bonding is almost purely ionic and the perfect ionic model fits.

B9 LE-Q14 — [3]

B1 forming Mg^{2+} costs an extra 1451 kJ mol^{-1}, and a second chlorine atom must also be atomised and ionised. **B1** the lattice of Mg^{2+} and Cl^- ions has a far more exothermic lattice enthalpy, because the cation has twice the charge and is smaller, and there are more ions per formula unit. **B1** the extra lattice enthalpy more than repays the second ionisation energy, so the enthalpy of formation of $MgCl_2$ is much more negative than that of MgCl would be.

B10 LE-Q17 — [3]

B1 a gas of separate ions cannot be prepared and reacted in a calorimeter, so the lattice-forming step cannot be carried out and measured. **B1** by Hess's law the enthalpy change from the elements to the solid is the same by any route. **B1** the other steps in the cycle, atomisation, ionisation energy, electron affinity and formation, are measured, so the lattice enthalpy is the only unknown and is found by subtraction.

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B11 LE-Q18

B1 a bromine atom is larger, so the incoming electron enters a shell further from the nucleus, 4p rather than 3p. **B1** the incoming electron is more shielded from the nucleus by more inner shells of electrons. **B1** the attraction between the nucleus and the incoming electron is weaker, so less energy is released; ignore any reference to nuclear charge alone. [3]

B12 LE-Q19

B1 the incoming electron enters the 2p subshell of fluorine, which already holds five electrons in a small volume. **B1** the repulsion between the incoming electron and those electrons is large. **B1** the repulsion cancels part of the stronger attraction to the nucleus, so less energy is released than for chlorine. [3]

B13 LE-Q21

B1 Li^+ is much smaller than Na^+ with the same charge, so its charge density is higher. **B1** Li^+ polarises the carbonate ion more, distorting its electron cloud and weakening a C–O bond. **B1** lithium carbonate therefore decomposes to Li_2O and CO_2 at a temperature a Bunsen burner reaches, while sodium carbonate needs a far higher one. [3]

B14 LE-Q23

B1 the enthalpy of atomisation of chlorine, $\frac{1}{2}Cl_2(g) \rightarrow Cl(g)$, forms one mole of atoms, and two moles of gaseous chlorine atoms are needed for one mole of $MgCl_2$, so the value is doubled; allow one Cl–Cl bond enthalpy instead. **B1** the lattice holds Mg^{2+} ions, so two electrons are removed from each gaseous magnesium atom, $Mg(g) \rightarrow Mg^+(g) \rightarrow Mg^{2+}(g)$, which needs the first and then the second ionisation energy. **B1** the first electron affinity of chlorine is doubled for the same reason, one for each Cl^- ion; every step must match the one mole of $MgCl_2(s)$ that the lattice enthalpy and the enthalpy of formation refer to. [3]

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