15.1 Energy Cycles Key

1. Which combination of ionic charge and ionic radius give the largest lattice enthalpy for an ionic compound?

↑ charge, ↑ attraction
↓ radius, ↑ attraction

	Ionic charge	Ionic radius
A.	high	large
B.	**high**	**small**
C.	low	small
D.	low	large

2. The lattice enthalpy values for lithium fluoride and calcium fluoride are shown below.

LiF(s) $\Delta H^{\ominus} = +1022 \text{ kJ mol}^{-1}$

$CaF_2(s)$ $\Delta H^{\ominus} = +2602 \text{ kJ mol}^{-1}$ → higher charge, need to break 2 bonds

Which of the following statements help(s) to explain why the value for lithium fluoride is less than that for calcium fluoride?

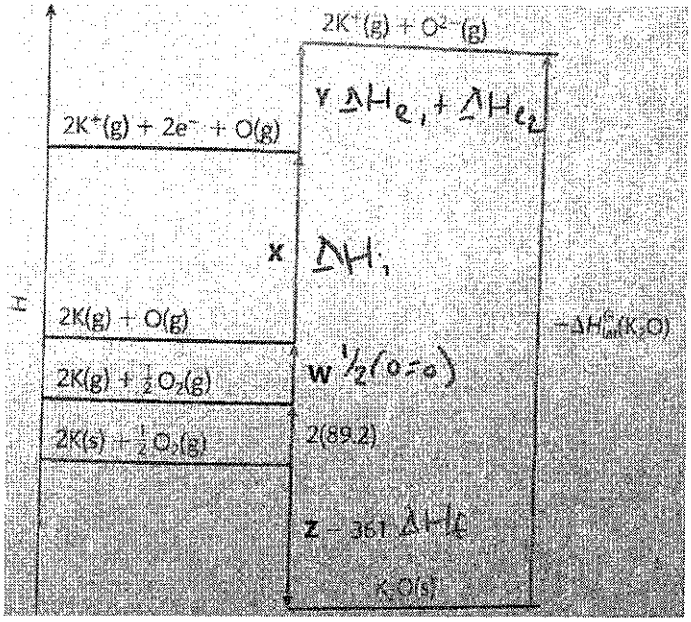
I. The ionic radius of lithium is less than that of calcium.
II. The ionic charge of lithium is less than that of calcium.

a. I only
b. II only
c. I and II
d. Neither I nor II

3. Identify the process, which has the sign of its associated enthalpy change different from the rest?

a. $Cl(g) + e^- \rightarrow Cl^-(g)$ (−) forming
b. $K(g) \rightarrow K^+(g) + e^-$ (+)
c. $KCl(g) \rightarrow K^+(g) + Cl^-(g)$ (+) breaking
d. $Cl_2(g) \rightarrow 2Cl(g)$ (+)

4. Answer the following questions using the diagram:
   a. Write an equation to represent the lattice energy of potassium oxide, K_2O. The Born-Haber cycle shown may be used to calculate the lattice energy of potassium oxide



$$K_2O(s) \rightarrow 2K^+(g) + 1O^{2-}(g)$$

   b. Identify the enthalpy changes labelled by the letters W,X,Y and Z
   c. Use the energy cycle, and further information from section 7 and 11 of the IB Data booklet to calculate an experimental value for the lattice energy of potassium oxide.

$$\Delta H_{lat} = \Delta H_{atom} + \tfrac{1}{2}E + \Delta H_i + \Delta H_e - \Delta H_f$$
$$= 2(89.2) + \tfrac{1}{2}(498) + 2(419) + (-141 + 753) - (-361)$$
$$= 2238.4 \text{ kJ mol}^{-1}$$

5. The theoretical lattice enthalpies of sodium chloride and magnesium oxide are shown below. Explain the higher lattice enthalpy of magnesium oxide compared to sodium chloride.

   Compound: NaCl; Lattice enthalpy: 769 $kJmol^{-1}$
   Compound: MgO; Lattice enthalpy: 3795 $kJmol^{-1}$

   Charge of both positive and negative ones are double, leads to a quadruple increase in lattice enthalpy

   Further enhanced by decrease in atomic raid of Mg due to increase in nuclear charge

6. The lattice enthalpies of silver bromide and sodium bromide are given below

   $\Delta H^{\ominus}$/kJ mol^{-1}
   AgBr: 905
   NaBr: 691

   Explain the relative values of the lattice enthalpies with reference to the bonding.

   Since both metals and nonmetal have the same charge, ionic charge does of affect lattice enthalpies difference. It would be expected that Na would have the larger enthalpy due to smaller radius, but Ag takes on covalent properties with its large number of electrons. The additional force caused by van deer Waals forces, the enthalpy required to separate AgBr increases.

7. Discuss the relative enthalpies of hydration of the K^+ and F^- ions in relation to their ionic radii.

Due to potassium being in the third period, it has a decreased enthalpy of hydration due to an increasing radius. There is less attraction between ions/charges due to increasing distance to nucleus. Vice versa for fluorine. Due to fluorine being the first non-metal in the halogen family, it has the smallest radius and therefore largest enthalpy.

8. Use an energy cycle to calculate the enthalpy of solution of potassium chloride from data in sections 18 and 20 of the IB data booklet.
Calculate the % inaccuracy of your value by comparing with the value in section 19 and comment on the disagreement between the two values.

$\Delta H^{\ominus}_{soln(KCl)} = \Delta H^{\ominus}_{latt(KCl)} + \Delta H^{\ominus}_{hyd(K^+)} + \Delta H^{\ominus}_{hyd(Cl^-)}$

$K^+(g) + Cl^-(g)$

$KCl(s)$ $\quad$ $K^+(aq) + Cl^-(aq)$

$\Delta H^{\ominus}_{soln(KCl)} = ?$

$= (+720) + (-340) + (-359)$

$= +21 \text{ kJmol}^{-1}$

real value: +17.22 % error 21.95 % error

* When working with large values and small outcomes, larger chance for error.