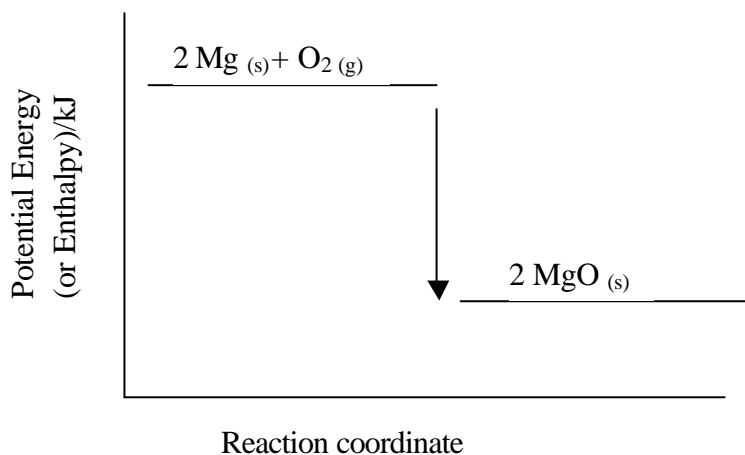


IB1 HL Chemistry

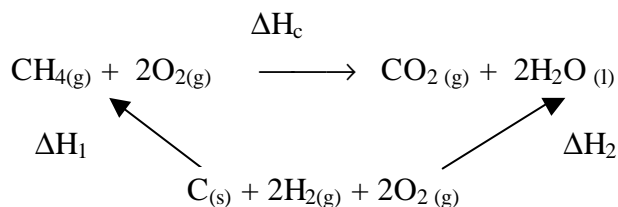
5. Energetics

Solution to further problems

1. Decrease in the strength and/or the number of bonds causes energy to be released in an exothermic reaction.
2. a)



- b) The reaction is exothermic.
 - c) Reactants.
3. Heat gained by the calorimeter + water sample in calorimeter = - heat lost by 2nd water sample
 $\Rightarrow (\text{mass of calorimeter} \times C_p \text{ of calorimeter} \times \Delta T) + (\text{mass of water} \times C_p \text{ of water} \times \Delta T) = -$
 $(\text{mass of second water sample} \times C_p \text{ of water} \times \Delta T)$
 $\Rightarrow (\text{Heat capacity of calorimeter}) (32.6 - 24.1) \text{ K} + (0.050 \text{ kg}) (4.18 \text{ kJ kg}^{-1} \text{ K}^{-1}) (32.6 - 24.1) \text{ K}$
 $= - (0.050 \text{ kg}) (4.18 \text{ kJ kg}^{-1} \text{ K}^{-1}) (32.6 - 41.7) \text{ K}$
 $\Rightarrow (\text{Heat capacity of calorimeter}) (8.5 \text{ K}) + 1.77 \text{ kJ} = 1.90 \text{ kJ}$
 $\Rightarrow \text{Heat capacity of calorimeter} = [(1.90 - 1.77) \text{ kJ}] / 8.5 \text{ K}$
 $\Rightarrow \text{Heat capacity of calorimeter} = 0.01529 \text{ kJ K}^{-1} = 0.015 \text{ kJ K}^{-1}$
 4. a) If $2 \text{ NO}_{(g)} + \text{Cl}_{2(g)} \rightarrow 2 \text{ NOCl}_{(g)} \quad \Delta H^\theta = -77.4 \text{ kJ}$
then $\text{NO}_{(g)} + \frac{1}{2} \text{ Cl}_{2(g)} \rightarrow \text{NOCl}_{(g)} \quad \Delta H^\theta = -77.4 \text{ kJ} / 2 = -38.7 \text{ kJ}$
b) $6 \text{ NOCl}_{(g)} \rightarrow 6 \text{ NO}_{(g)} + 3 \text{ Cl}_{2(g)} \quad \Delta H^\theta = -77.4 \text{ kJ} \times 3 = -232.2 \text{ kJ}$
 5. $2 \text{ Na}_{(s)} + 2 \text{ S}_{(s)} + 3/2 \text{ O}_{2(s)} \rightarrow \text{Na}_2\text{S}_2\text{O}_{3(s)} \quad \Delta H = \Delta H_f^q$
 - 6.



$$\Delta H_2 = \Delta H_1 + \Delta H_c$$

$$\Delta H_c = \Delta H_2 - \Delta H_1$$

$$\text{Where, } \Delta H_1 = \Delta H_f^q [\text{CH}_{4(g)}] = (1 \text{ mol}) (-75 \text{ kJ mol}^{-1}) = -75 \text{ kJ}$$

$$\Delta H_2 = (2 \text{ mol}) \Delta H_f^q [\text{H}_2\text{O}_{(l)}] + (1 \text{ mol}) \Delta H_f^q [\text{CO}_{2(g)}]$$

$$= (2 \text{ mol}) \times (-285.83 \text{ kJ mol}^{-1}) + (1 \text{ mol})(-393.51 \text{ kJ mol}^{-1})$$

$$\Delta H_2 = -956.17 \text{ kJ}$$

$$\Delta H_c = -956.17 \text{ kJ} - (-75 \text{ kJ})$$

$$\Delta H_c = -890.17 \text{ kJ}$$

$$\Delta H_c = -8.90 \times 10^2 \text{ kJ mol}^{-1} \text{ (mol}^{-1} \text{ obtained from the equation)}$$

7. $M_R (\text{C}_5\text{H}_{10}\text{O}_5) = 155.15$

Heat gained by the calorimeter and contents = heat capacity $\times \Delta T$

$$= (15.8 \text{ kJ/}^\circ\text{C}) (210.54 - 20.00) ^\circ\text{C}$$

$$= 3010.532 \text{ kJ}$$

$\Delta H = -$ heat energy gained by calorimeter and its contents

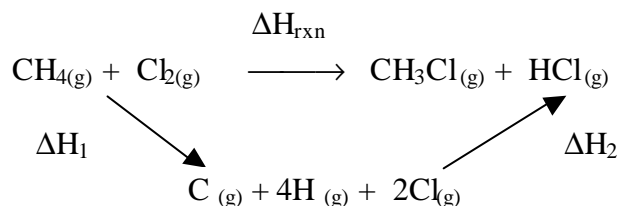
$$\Delta H = -3010 \text{ kJ}$$

$$\left(\frac{3010 \text{ kJ}}{0.548 \text{ g C}_5\text{H}_{10}\text{O}_5} \right) \left(\frac{155.15 \text{ g C}_5\text{H}_{10}\text{O}_5}{1 \text{ mol C}_5\text{H}_{10}\text{O}_5} \right) = 852343 \text{ kJ mol}^{-1}$$

$$\Delta H_c = -8.52 \times 10^5 \text{ kJ mol}^{-1}$$



9. (a)



$$\Delta H_{\text{rxn}} = \Delta H_1 + \Delta H_2$$

where, $\Delta H_1 =$ bond breaking

$$= (4 \text{ mol}) E(\text{C-H}) + (1 \text{ mol}) E(\text{Cl-Cl})$$

$$= (4 \text{ mol}) (412 \text{ kJ mol}^{-1}) + (1 \text{ mol}) (242 \text{ kJ mol}^{-1})$$

$$= 1890 \text{ kJ}$$

$\Delta H_2 =$ bond making

$$= (3 \text{ mol}) E(\text{C-H}) + (1 \text{ mol}) E(\text{C-Cl}) + (1 \text{ mol}) E(\text{H-Cl})$$

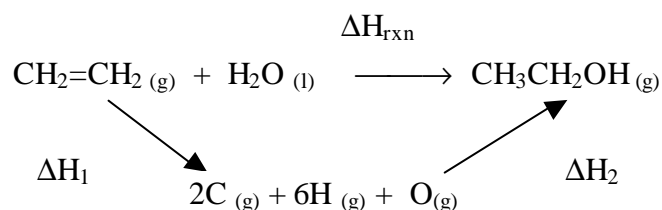
$$= (3 \text{ mol})(-412 \text{ kJ mol}^{-1}) + (1 \text{ mol})(-338 \text{ kJ mol}^{-1}) + (1 \text{ mol})(-431 \text{ kJ mol}^{-1})$$

$$= -2005 \text{ kJ}$$

$$\Delta H_{\text{rxn}} = (1890 - 2005) \text{ kJ} = -115 \text{ kJ}$$

$$\Delta H_{\text{rxn}} = -115 \text{ kJ mol}^{-1} \text{ (mol}^{-1} \text{ from the equation)}$$

(b)



$$\Delta H_{\text{rxn}} = \Delta H_1 + \Delta H_2$$

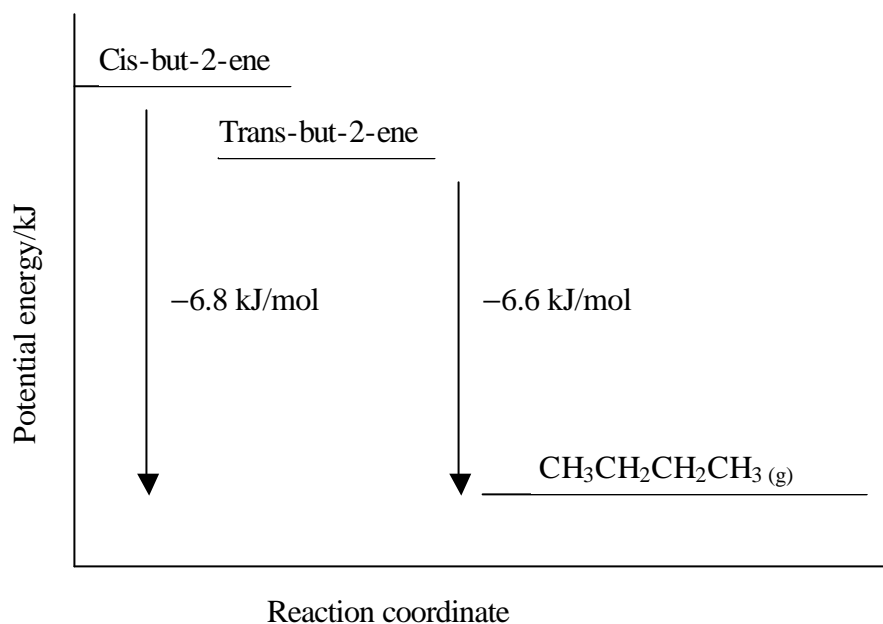
where, ΔH_1 = bond breaking

$$\begin{aligned} &= (4 \text{ mol}) E(\text{C-H}) + (1 \text{ mol}) E(\text{C}=\text{C}) + (2 \text{ mol}) E(\text{H-O}) \\ &= (4 \text{ mol}) (412 \text{ kJ mol}^{-1}) + (1 \text{ mol})(612 \text{ kJ mol}^{-1}) + (2 \text{ mol}) (463 \text{ kJ mol}^{-1}) \\ &= 3186 \text{ kJ} \end{aligned}$$

$$\begin{aligned} \Delta H_2 = \text{bond making} &= (5 \text{ mol}) E(\text{C-H}) + (1 \text{ mol}) E(\text{C-C}) + (1 \text{ mol}) E(\text{C-O}) + (1 \text{ mol}) E(\text{O-H}) \\ &= (5 \text{ mol}) X (-412 \text{ kJ mol}^{-1}) + (1 \text{ mol})(-348 \text{ kJ mol}^{-1}) + (1 \text{ mol})(-360 \text{ kJ mol}^{-1}) + (1 \text{ mol}) \\ &\quad (-463 \text{ kJ mol}^{-1}) \\ &= -3231 \text{ kJ} \end{aligned}$$

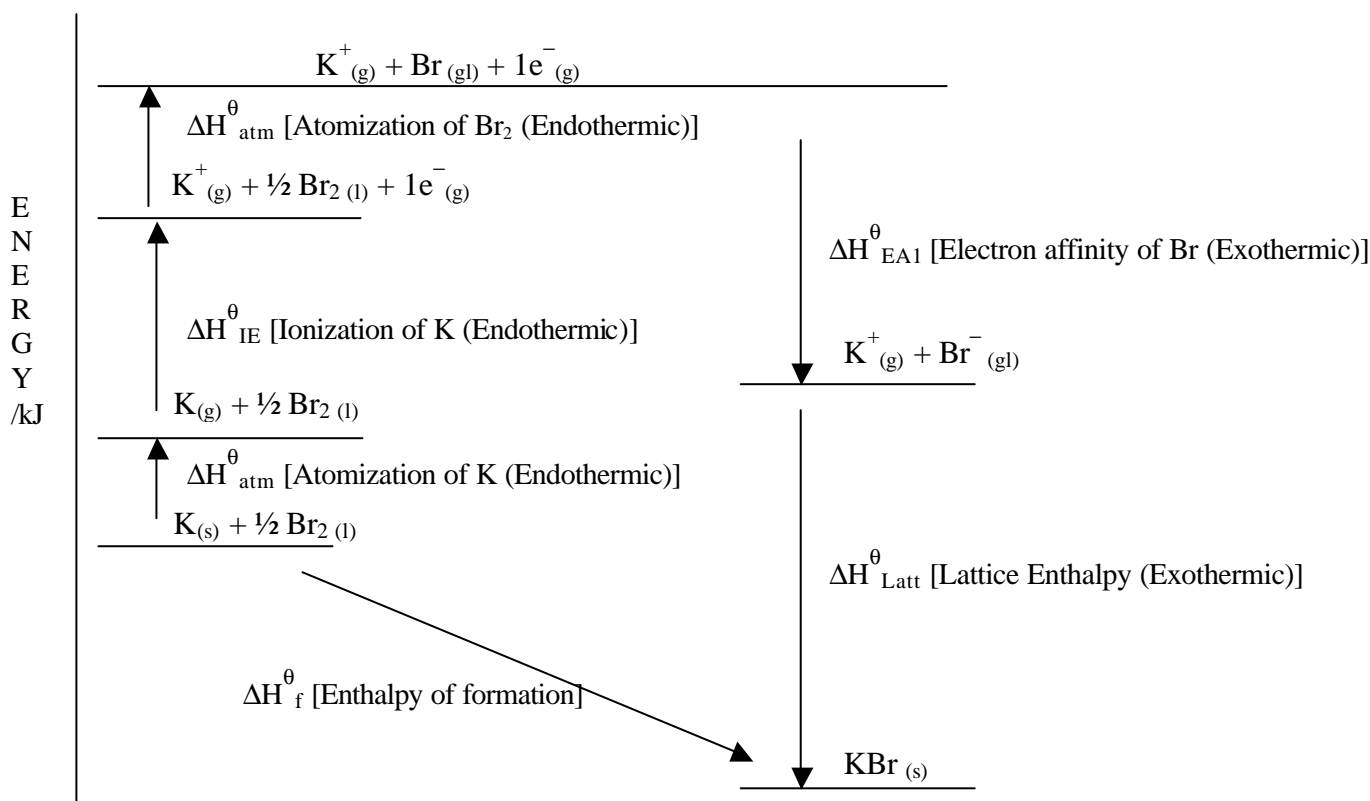
$$\Delta H_{\text{rxn}} = (3186 - 3231) \text{ kJ} = -45 \text{ kJ}$$

10.



Trans-but-2-ene is thermodynamically more stable. Because hydrogenation of trans-but-2-ene give off less heat than cis-but-2-ene, it can be said that the former is at a lower potential energy than the latter and therefore more stable as depicted above.

11.



Reaction coordinate

12. Enthalpy of formation of CaCl_2 = Enthalpy of atomization of Ca (enthalpy of vaporization of Ca) + 1st and 2nd ionization of Ca + Bond enthalpy of Cl_2 (or 2 X atomization of Cl_2) + 2 mol X Electron affinity of chlorine + Lattice enthalpy
 $(-795 \text{ kJ/mol}) = (192 \text{ kJ/mol}) + (589 \text{ kJ/mol}) + (1146 \text{ kJ/mol}) + (238 \text{ kJ/mol}) + (2)(-348 \text{ kJ/mol})$
 + Lattice enthalpy
 $-795 \text{ kJ/mol} = 192 \text{ kJ/mol} + 589 \text{ kJ/mol} + 1146 \text{ kJ/mol} + 238 \text{ kJ/mol} - 696 \text{ kJ/mol}$
 + Lattice enthalpy
 $\text{Lattice enthalpy} = -795 \text{ kJ/mol} - 192 \text{ kJ/mol} - 589 \text{ kJ/mol} - 1146 \text{ kJ/mol} - 238 \text{ kJ/mol}$
 + 696 kJ/mol
 $\text{Lattice enthalpy} = -2960 \text{ kJ/mol} + 696 \text{ kJ/mol}$
 $\text{Lattice enthalpy} = -2264 \text{ kJ/mol}$
13. (a) NaF has a higher lattice enthalpy than NaCl because of the smaller size of the F^- relative to Cl^- . The smaller the ions, the closer they can get and the stronger the electrostatic force of attraction between the ions leading to a higher lattice enthalpy.
 (b) LiBr has a higher lattice enthalpy than CsBr because of the smaller size of Li^+ relative to Cs^+ . The smaller the ions, the smaller the internuclear distance between the oppositely charged ions and therefore stronger the electrostatic force of attraction leading to a higher lattice enthalpy.
 (c) SrS has a higher lattice enthalpy than SrCl_2 because of the fact that the lattice of SrCl_2 would contain a cation to anion ratio of 1:2 thus increasing the distance between the oppositely charged ions in the lattice. Though the Cl^- is smaller than S^{2-} , the inter-nuclear distance between the ions in SrS is most likely to be smaller than in SrCl_2 on account of the stoichiometry and therefore a higher lattice enthalpy.
14. The type of compounds that show the greatest discrepancies between experimental and theoretical values of lattice enthalpy are those that have a small difference in the electronegativities of the elements involved.

A possible explanation is the nature of the actual bonds. The actual bonds in these compounds are not completely ionic as assumed in the theoretical determination of the lattice enthalpies of those compounds. The bonds in those compounds have significant covalent characters resulting from the polarization of the anion by the cation in the compound, and therefore the greatest discrepancies between their theoretical and the experimental lattice enthalpies.