Reactivity 1.2 Answers

IB CHEMISTRY SL

Reactivity 1.2.1
Understandings:
- Bond-breaking absorbs energy and bond-forming releases energy.

Learning outcomes:
- Calculate the enthalpy change of a reaction from given average bond enthalpy data.

Additional notes:
- Include explanation of why bond enthalpy data are average values and may differ from those measured experimentally.
- Average bond enthalpy values are given in the data booklet.

Linking questions:
- Structure 2.2 How would you expect bond enthalpy data to relate to bond length and polarity?
- Reactivity 3.4 How does the strength of a carbon– halogen bond affect the rate of a nucleophilic substitution reaction?

Bond enthalpy and average bond enthalpies
- Average bond enthalpy is the energy needed to break one mole of bonds in a gaseous molecule averaged over similar compounds.
- Bond breaking is endothermic (requires energy).
- Bond making is exothermic (releases energy).

$$\Delta H = \sum \text{(bonds broken)} - \sum \text{(bonds formed)}$$

Example: Calculate the enthalpy change for the following reaction.
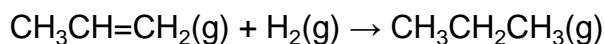
$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$$
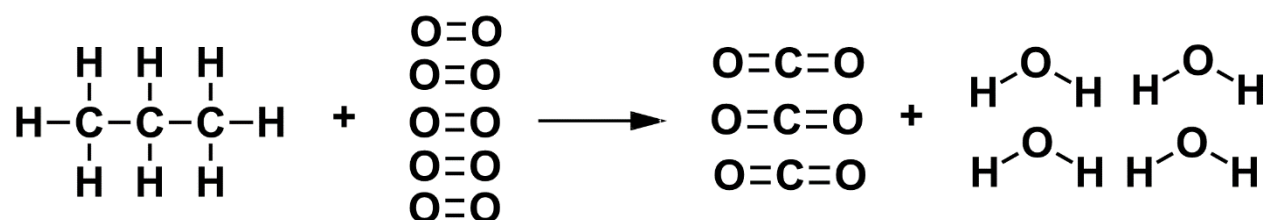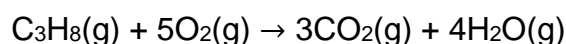
$\Delta H = \sum\text{(bonds broken)} - \sum \text{(bonds formed)}$

$\Delta H = [8(414) + 2(346) + 5(498)] - [6(804) + 8(463)]$

$\Delta H = 6494 - 8528 = -2034 \text{ kJ mol}^{-1}$

Exercises:

1. Propane can be formed by the hydrogenation of propene.

$$CH_3CH=CH_2(g) + H_2(g) \rightarrow CH_3CH_2CH_3(g)$$

Determine a value for the hydrogenation of propene using information from Table 11 of the Data Booklet. State and explain if the reaction is exothermic or endothermic.

Bonds broken: Bonds formed:

6 C-H (414 kJ)   1 H-H (436 kJ)   8 C-H (414 kJ)

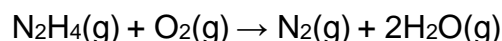1 C=C (614 kJ)   2 C-C (346 kJ)

1 C-C (346 kJ)

$\Delta H = \sum(\text{bonds broken}) - \sum(\text{bonds formed})$

$\Delta H = 3880 - 4004$

$\Delta H = -124 \text{ kJ mol}^{-1}$

Reaction is exothermic (negative ΔH)

2. Use the information from Table 11 of the Data Booklet to calculate the enthalpy change for the following reaction.

$$N_2H_4(g) + O_2(g) \rightarrow N_2(g) + 2H_2O(g)$$

Bonds broken: Bonds formed:
1 × 158 kJ   1 × 945 kJ
4 × 391 kJ   4 × 463 kJ
1 × 498 kJ

$\Delta H = \sum(\text{bonds broken}) - \sum(\text{bonds formed})$

$\Delta H = 2220 - 2797$

$\Delta H = -577 \text{ kJ mol}^{-1}$

3. State three reasons why enthalpy changes calculated using average bond enthalpies can be inaccurate.

Average bond enthalpies apply to the gaseous state only.
The actual value for the specific bonds may be different to the average value.
They do not take into account intermolecular forces.

Reactivity 1.2.2

Understandings:

- Hess's law states that the enthalpy change for a reaction is independent of the pathway between the initial and final states.
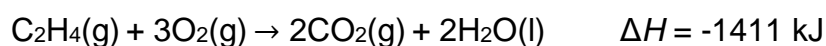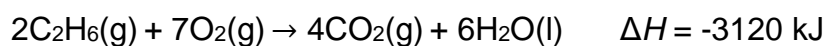
Learning outcomes:

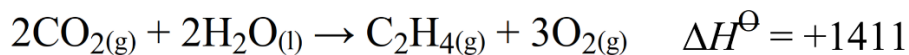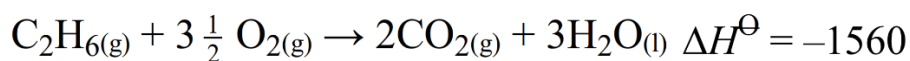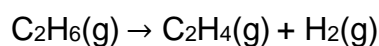- Apply Hess's law to calculate enthalpy changes in multistep reactions.

Calculating the enthalpy change of a reaction that is the sum of multiple reactions with known enthalpy changes

- Applying Hess's law, the enthalpy change for a reaction can be calculated using multiple reactions with known enthalpy changes.

Example:

$$2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l) \quad \Delta H = \text{-3120 kJ}$$
$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l) \quad \Delta H = \text{-572 kJ}$$
$$C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l) \quad \Delta H = \text{-1411 kJ}$$

Calculate the enthalpy change for the following reaction:

$$C_2H_6(g) \rightarrow C_2H_4(g) + H_2(g)$$

$$C_2H_{6(g)} + 3\tfrac{1}{2}\,O_{2(g)} \rightarrow 2CO_{2(g)} + 3H_2O_{(l)} \quad \Delta H^\ominus = -1560$$
$$H_2O_{(l)} \rightarrow H_{2(g)} + \tfrac{1}{2}\,O_{2(g)} \quad \Delta H^\ominus = +286$$
$$2CO_{2(g)} + 2H_2O_{(l)} \rightarrow C_2H_{4(g)} + 3O_{2(g)} \quad \Delta H^\ominus = +1411$$

ΔH = -1560 + 286 + 1411 = 137 kJ

Exercises:

1. Using the equations below:

$$C(s) + O_2(g) \rightarrow CO_2(g) \qquad \Delta H = \text{-390 kJ}$$
$$Mn(s) + O_2(g) \rightarrow MnO_2(s) \quad \Delta H = \text{-520 kJ}$$

Determine the ΔH (in kJ) for the following reaction.

$$MnO_2(s) + C(s) \rightarrow Mn(s) + CO_2(g)$$
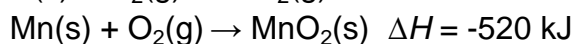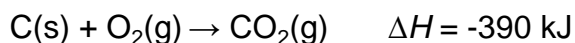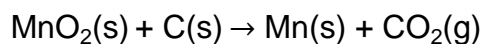
ΔH = +130 kJ

2. Using the equations below:

$$Cu(s) + \tfrac{1}{2}O_2(g) \rightarrow CuO(s) \qquad \Delta H = \text{-156 kJ}$$
$$2Cu(s) + \tfrac{1}{2}O_2(g) \rightarrow Cu_2O(s) \quad \Delta H = \text{-170 kJ}$$

Determine the ΔH (in kJ) for the following reaction.

$$2CuO(s) \rightarrow Cu_2O(s) + \tfrac{1}{2}O_2(g)$$
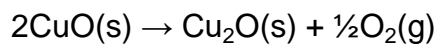
ΔH = +142 kJ

3. Calculate the enthalpy change, ΔH, for the reaction:

$$C(s) + 2H_2(g) + \tfrac{1}{2}O_2(g) \rightarrow CH_3OH(l)$$
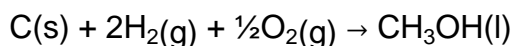
using Hess's Law and the following information.

$$CH_3OH(l) + 1\tfrac{1}{2}O_2(g) \rightarrow CO_2(g) + 2H_2O(l) \qquad \Delta H_1 = \text{-676 kJ}$$
$$C(s) + O_2(g) \rightarrow CO_2(g) \qquad\qquad\qquad\quad \Delta H_2 = \text{-394 kJ}$$
$$H_2(g) + \tfrac{1}{2}O_2(g) \rightarrow H_2O(l) \qquad\qquad\qquad\quad \Delta H_3 = \text{-242 kJ}$$

$CO_2 + 2H_2O \rightarrow CH_3OH + 1\tfrac{1}{2}O_2$ ΔH_1 = +676 kJ mol^{-1} (reversed)
$C + O_2 \rightarrow CO_2$ ΔH_2 = -394 kJ mol^{-1} (don't change)
$2H_2 + O_2 \rightarrow 2H_2O$ ΔH_3 = -484 kJ mol^{-1} (multiply by two)

ΔH = -202 kJ