

# Energetics SL (answers)

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IB CHEMISTRY SL

25 <b>Mn</b> Manganese 54.938045	16 <b>S</b> Sulfur 32.065	<b>J</b>	6 <b>C</b> Carbon 12.0107	2 <b>He</b> Helium 4.002602	25 <b>Mn</b> Manganese 54.938045
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## 5.1 Measuring energy changes

### Understandings:

- Heat is a form of energy.
- Temperature is a measure of the average kinetic energy of the particles.
- Total energy is conserved in chemical reactions.
- Chemical reactions that involve transfer of heat between the system and the surroundings are described as endothermic or exothermic.
- The enthalpy change ( $\Delta H$ ) for chemical reactions is indicated in  $\text{kJ mol}^{-1}$
- $\Delta H$  values are usually expressed under standard conditions, given by  $\Delta H^\ominus$ , including standard states.

### Applications and skills:

- Calculation of the heat change when the temperature of a pure substance is changed using  $q = mc\Delta T$
- A calorimetry experiment for an enthalpy of reaction should be covered and the results evaluated.

### Guidance:

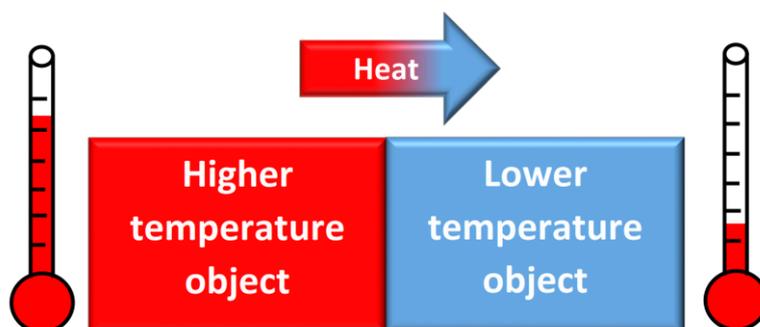
- Enthalpy changes of combustion ( $\Delta H_c^\ominus$ ) and formation ( $\Delta H_f^\ominus$ ) should be covered.
- Consider reactions in aqueous solution and combustion reactions.
- Standard state refers to the normal, most pure stable state of a substance measured at 100 kPa. Temperature is not a part of the definition of standard state, but 298 K is commonly given as the temperature of interest.
- The specific heat capacity of water is provided in the data booklet in section 2.
- Students can assume the density and specific heat capacities of aqueous solutions are equal to those of water but should be aware of this limitation.
- Heat losses to the environment and the heat capacity of the calorimeter in experiments should be considered, but the use of a bomb calorimeter is not required.

## Syllabus checklist

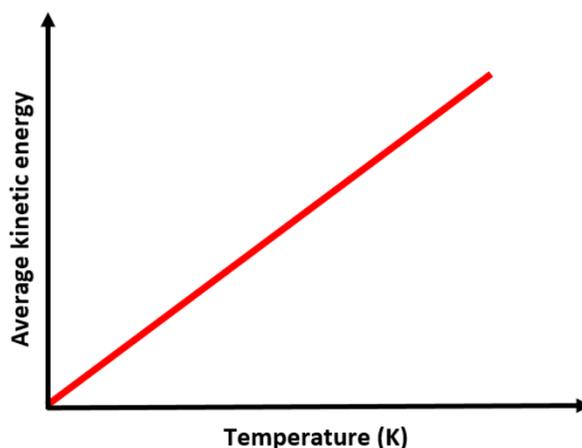
Objective	I am confident with this	I need to review this	I need help with this
Outline the difference between heat and temperature			
State the conditions for the standard enthalpy change ( $\Delta H^\ominus$ )			
Describe the characteristics of exothermic and endothermic reactions			
Sketch enthalpy level diagrams for exothermic and endothermic reactions			
Calculate enthalpy changes using the equation $q=mc\Delta T$			
Describe possible errors and limitations in experiments to calculate enthalpy change			

## Heat and temperature

- Heat is the transfer of energy from a higher temperature object to a lower temperature object (measured in joules, J).
- The direction of heat flow is from hot (higher temperature) to cold (lower temperature).



- Absolute temperature is the measure of the average kinetic energy of the particles in a substance.
- Absolute temperature in kelvin (K) is directly proportional to the average kinetic energy of the particles in a substance.



## Converting between kelvin (K) and °C

- To convert between °C and K, add or subtract 273.
- The kelvin scale does not have a negative value.

### Example:

- $100\text{ }^{\circ}\text{C} = 373\text{ K}$
- $0\text{ }^{\circ}\text{C} = 273\text{ K}$
- $0\text{ K} = -273\text{ }^{\circ}\text{C}$  (absolute zero)

## Exercises:

1. Define the terms heat and absolute temperature.

Absolute temperature (in K) is a measure of the average kinetic energy of the particles in a substance.

Heat is the transfer of energy from an area of higher temperature to lower temperature.

2. Convert the following temperatures from °C to K ( $0\text{ °C} = 273\text{ K}$ )

a) 298 K      b) 351 K      c) 373 K      d) 0 K      e) 260 K

3. Two samples of chlorine gas (both containing the same number of molecules) are at 350 K and 550 K respectively. Which sample has the higher average kinetic energy?

The sample of gas at 550 K has the highest average kinetic energy.

4. Two beakers of water, one containing  $250\text{cm}^3$  and the other  $500\text{cm}^3$  of water, are at the same temperature (298 K). Compare the average kinetic energy and the heat content of the two beakers of water.

Both beakers of water are at the same temperature, therefore they have the same average kinetic energy. The beaker with  $500\text{ cm}^3$  of water has more heat than the other beaker.

5. A hot piece of iron with a temperature of 500 K is put into a beaker of water at 298 K. Describe which way the heat will flow and what will happen to the temperature of the iron and the water after one hour.

The heat will flow from the block of iron to the water (from higher temperature to lower temperature). After one hour, the temperature of the iron and the water will be the same – they will have reached thermal equilibrium.

## Enthalpy change ( $\Delta H$ )

- Enthalpy,  $H$ , is the heat content of a substance.
- Enthalpy is the sum of the internal energy ( $U$ ) and the product of pressure and volume ( $PV$ ):

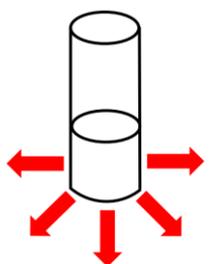
$$H = U + PV$$

- Enthalpy change ( $\Delta H$ ) is the amount of heat absorbed or released in a chemical reaction.
- If a process occurs at constant pressure, the heat released or absorbed is equal to the change in enthalpy ( $q = \Delta H$ ).
- The standard enthalpy change of reaction ( $\Delta H^\ominus$ ) is the enthalpy change measured under standard conditions (100 kPa and usually 298 K).

## Exothermic and endothermic reactions

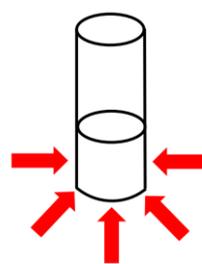
- Chemical reactions can be classified as exothermic or endothermic depending on whether heat is absorbed or released by the system or surroundings.

### Exothermic reaction



Heat flows from the system to the surroundings

### Endothermic reaction



Heat flows from surroundings into the system

- The system is the reactants and products (the chemical reaction).
- The surroundings consist of everything else outside of the system (including the reaction mixture).

## Exothermic reactions

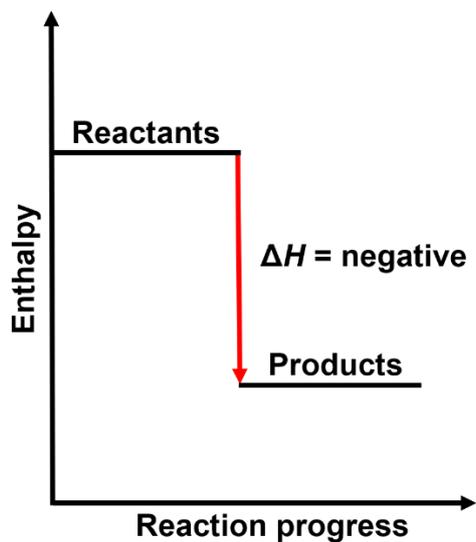
- In an exothermic reaction, heat flows from the system to the surroundings.
- The temperature of the reaction mixture and the surroundings increases as heat is released by the system.
- Examples of exothermic reactions include combustion and neutralisation.

## Endothermic reactions

- In an endothermic reaction, heat flows from the surroundings to the system.
- The temperature of the reaction mixture and the surroundings decreases as heat is absorbed by the system.
- Examples of endothermic reactions include photosynthesis and thermal decomposition reactions.

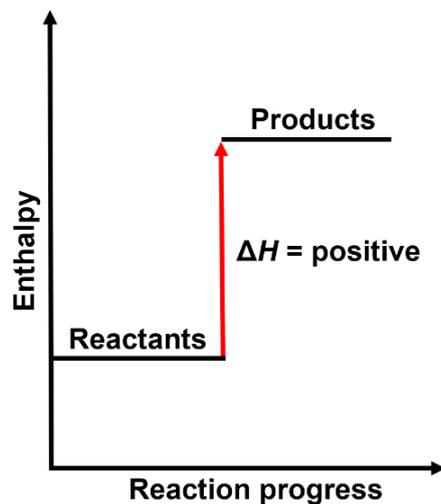
## Enthalpy level diagrams

Enthalpy level diagram an exothermic reaction



- $\Delta H = \text{negative} (< 0)$
- Products have lower enthalpy than reactants.
- Products are more energetically stable than reactants.

Enthalpy level diagram an endothermic reaction



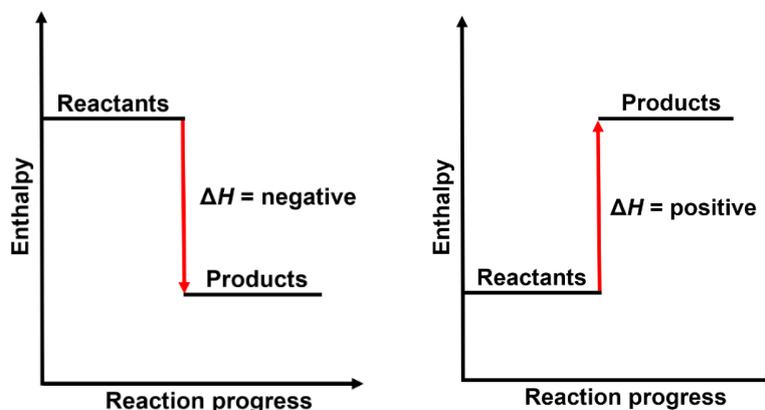
- $\Delta H = \text{positive} (> 0)$
- Products have higher enthalpy than reactants.
- Products are less energetically stable than reactants.

## Exercises

1. Define the term *standard enthalpy change of reaction* ( $\Delta H^\ominus$ ).

Standard enthalpy change of reaction is the enthalpy change of a reaction carried out under standard conditions (100 kPa, 298 K).

2. Sketch enthalpy level diagrams for an exothermic reaction and an endothermic reaction.

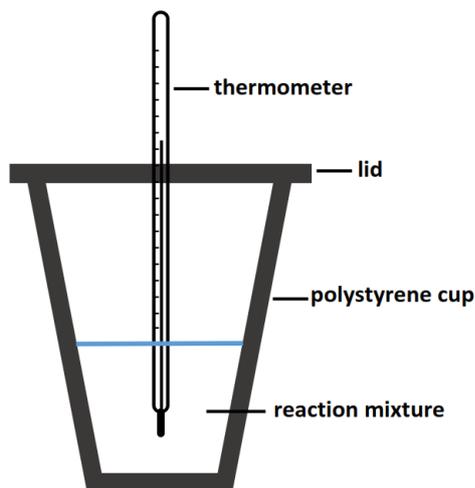


3. Complete the table summarising exothermic and endothermic reactions.

	<b>Exothermic reaction</b>	<b>Endothermic reaction</b>
Sign of enthalpy change ( $\Delta H$ )	negative	positive
Which has lower enthalpy, reactants or products?	products	reactants
Which is more energetically stable?	products	reactants
Is heat absorbed or released by the system?	released	absorbed

4. Classify the following reactions as endothermic or exothermic:
  - a. Propane burns in excess oxygen releasing heat. **exothermic**
  - b. The reading on the thermometer increases when sodium hydroxide and hydrochloric acid are mixed together in a beaker. **exothermic**
  - c. When ammonium nitrate and water are mixed together in a cold pack, the temperature decreases. **endothermic**

## Calculating enthalpy changes from experimental data



Enthalpy changes can be measured using a simple calorimeter (shown opposite). The heat absorbed or released in the reaction increases or decreases the temperature of a known mass of water (or solution).

The amount of heat absorbed or released can be calculated using the equation  $q = mc\Delta T$

$$q = mc\Delta T$$

$q$  - heat (J)

$m$  - mass of water (or solution)

$c$  - specific heat capacity ( $\text{J g}^{-1} \text{ } ^\circ\text{C}^{-1}$ )

$\Delta T$  - change in temperature

### Specific heat capacity

- Specific heat capacity,  $c$ , is defined as the amount of heat required to raise the temperature of 1 kg of a substance by 1 K (or  $^\circ\text{C}$ ).
- The units are  $\text{J kg}^{-1} \text{ K}^{-1}$  or  $\text{J g}^{-1} \text{ K}^{-1}$  (or  $^\circ\text{C}$ ).

$$c = \frac{q}{m\Delta T}$$

- Substances with higher specific heat capacity values require more heat to change their temperature than substances with lower specific heat capacity values.
- The specific heat capacity of water is  $4.18 \text{ J g}^{-1} \text{ } ^\circ\text{C}^{-1}$

**Exercise:** The same amount of heat is added to equal masses of each of the substances in the table below. List the substances in order of the smallest temperature change to the largest temperature change when the heat is added.

Substance	Specific heat capacity $\text{J g}^{-1} \text{ }^\circ\text{C}^{-1}$
Aluminium	0.903
Ethanol	2.42
Copper	0.385
Silver	0.235
Iron	0.449

Smallest temperature change				Largest temperature change
Ethanol	Aluminium	Iron	Copper	Silver

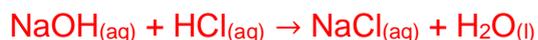
The higher the specific heat capacity, the smaller the temperature change.

### Exercises:

1. Calculate the enthalpy change for the following reactions.

For examples (a) and (b) the enthalpy change is known as the standard enthalpy change of neutralisation ( $\Delta H^\ominus_n$ ), which is the enthalpy change when an acid and base are reacted together to form one mole of water under standard conditions.

- a.  $200.0 \text{ cm}^3$  of  $0.400 \text{ mol dm}^{-3}$  HCl is added to the same volume and concentration of NaOH solution inside a coffee-cup calorimeter. The initial temperature was  $25.1 \text{ }^\circ\text{C}$ , and the maximum temperature was  $27.7 \text{ }^\circ\text{C}$ . Calculate the enthalpy change of the reaction.



$$n(\text{NaOH}) = 0.400 \times (200.0 \div 1000) = 0.0800 \text{ mol}$$

$$n(\text{HCl}) = 0.400 \times (200.0 \div 1000) = 0.0800 \text{ mol}$$

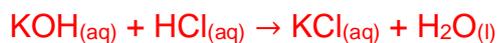
$$q = mc\Delta T$$

$$q = 400.0 \times 4.18 \times (27.7 - 25.1)$$

$$q = 4347.2 \text{ J}$$

For one mole of water produced:  $4347.2 \div 0.0800 = -54340 \text{ J mol}^{-1} = -54.3 \text{ kJ mol}^{-1}$   
(note the negative sign because reaction is exothermic)

- b. 400.0 cm<sup>3</sup> of 0.500 mol dm<sup>-3</sup> hydrochloric acid was reacted with 400.0 cm<sup>3</sup> of 1.00 mol dm<sup>-3</sup> potassium hydroxide and the temperature increased by 3.20 °C. Calculate the enthalpy change of the reaction.



$$n(\text{KOH}) = 1.00 \times (400.0 \div 1000) = 0.400 \text{ mol}$$

$$n(\text{HCl}) = 0.500 \times (400.0 \div 1000) = 0.200 \text{ mol}$$

HCl is the limiting reactant – ratio of HCl to H<sub>2</sub>O is 1:1

$$q = mc\Delta T$$

$$q = 800.0 \times 4.18 \times 3.20$$

$$q = 10700.8 \text{ J}$$

For one mole of water produced:  $-10700.8 \div 0.200 = -53504 \text{ J mol}^{-1} = -53.5 \text{ kJ mol}^{-1}$   
(note the negative sign because reaction is exothermic)

- c. The literature value for the standard enthalpy change of neutralisation is  $-57.0 \text{ kJ mol}^{-1}$ . For each of the above examples, calculate the percentage error and suggest possible sources of error.

$$\% \text{ error} = \frac{\text{experimental value} - \text{theoretical value}}{\text{theoretical value}} \times 100$$

Percentage error for reaction a:

$$\% \text{ error} = \frac{(-54.3) - (-57.0)}{(-57.0)} \times 100 = 5\%$$

Percentage error for reaction b:

$$\% \text{ error} = \frac{(-53.5) - (-57.0)}{(-57.0)} \times 100 = 6\%$$

Sources of error include: heat loss to the surroundings/some heat absorbed by the calorimeter

Assumptions: the solution has the same density and specific heat capacity as water (1 g cm<sup>-3</sup> / 4.18 J g<sup>-1</sup> °C<sup>-1</sup>).

- d. A 1.00 g sample of octane ( $C_8H_{18}$ ) is burned in a calorimeter containing 1200.0  $cm^3$  of water at an initial temperature of 25.0  $^{\circ}C$ . After the reaction, the final temperature of the water is 33.0  $^{\circ}C$ . Calculate the enthalpy change, in  $kJ mol^{-1}$ .

$$n(C_8H_{18}) = m \div M = 1.00 \div 114.26 = 8.75 \times 10^{-3} \text{ mol}$$

$$q = mc\Delta T$$

$$q = 1200.0 \times 4.18 \times (33.0 - 25.0)$$

$$q = 40128 \text{ J}$$

For one mole of  $C_8H_{18}$ :

$$-40128 \div 8.75 \times 10^{-3} = -4586057.143 \text{ J mol}^{-1} = -4.59 \times 10^3 \text{ kJ mol}^{-1}$$

(note the negative sign because reaction is exothermic)

- e. In an experiment to measure the enthalpy change of combustion of ethanol ( $C_2H_5OH$ ), a student heated a copper calorimeter containing 100.0  $cm^3$  of water with a spirit lamp and collected the following data.

Initial temperature of water:	20.0 $^{\circ}C$
Final temperature of water:	55.0 $^{\circ}C$
Mass of ethanol burned:	1.78 g

Calculate the enthalpy change of combustion of ethanol in  $kJ mol^{-1}$ .

$$n(C_2H_5OH) = m \div M = 1.78 \div 46.08 = 0.0386 \text{ mol}$$

$$q = mc\Delta T$$

$$q = 100.0 \times 4.18 \times (55.0 - 20.0)$$

$$q = 14630 \text{ J}$$

For one mole of  $C_2H_5OH$ :

$$-14630 \div 0.0386 = -379015.544 \text{ J mol}^{-1} = -379 \text{ kJ mol}^{-1}$$

(note the negative sign because reaction is exothermic)

- f. Use the values in the table below to calculate the percentage error for each example and suggest possible sources of error.

Compound	$\Delta H^{\ominus}_c$ (kJmol <sup>-1</sup> )
C <sub>8</sub> H <sub>18</sub>	-5483
C <sub>2</sub> H <sub>5</sub> OH	-1058

$$\% \text{ error} = \frac{\text{experimental value} - \text{theoretical value}}{\text{theoretical value}} \times 100$$

Percentage error for reaction d:

$$\% \text{ error} = \frac{(-4590) - (-5483)}{(-5483)} \times 100 = 16\%$$

Percentage error for reaction e:

$$\% \text{ error} = \frac{(-379) - (-1058)}{(-1058)} \times 100 = 64\%$$

Possible sources of error include heat loss to the surroundings /incomplete combustion/heat absorbed by the calorimeter.

## 5.2 Hess's law

### Understandings:

- The enthalpy change ( $\Delta H$ ) for a reaction that is carried out in a series of steps is equal to the sum of the enthalpy changes for the individual steps.

### Applications and skills:

- Application of Hess's Law to calculate enthalpy changes.
- Calculation of  $\Delta H$  values using  $\Delta H_f^\ominus$  data
- Determination of the enthalpy change of a reaction that is the sum of multiple reactions with known enthalpy changes.

### Guidance:

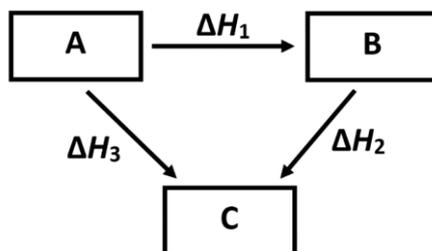
- Enthalpy of formation data can be found in the data booklet in section 12.
- An application of Hess's Law is  $\Delta H^\ominus = \sum \Delta H_f^\ominus \text{ products} - \sum \Delta H_f^\ominus \text{ reactants}$

### Syllabus checklist

Objective	I am confident with this	I need to review this	I need help with this
Outline the meaning of the standard enthalpy change of formation ( $\Delta H_f^\ominus$ )			
Calculate enthalpy changes using standard enthalpy change of formation values			
Determine the enthalpy change of a reaction that is the sum of multiple reactions with known enthalpy changes			
Interpret enthalpy cycles to calculate enthalpy changes			

## Hess's law

- Hess's law states:  
"The enthalpy change for a chemical reaction is independent of the route by which the chemical reaction occurs."
- This means that regardless of the route, if the starting point and end point are the same, the enthalpy change will also be the same.
- Enthalpy cycles are an application of Hess's law.
- Whichever route is taken to get to the products C, the enthalpy change is the same.

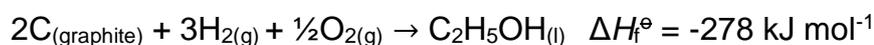


$$\Delta H_3 = \Delta H_1 + \Delta H_2$$

## Standard enthalpy change of formation ( $\Delta H_f^\ominus$ )

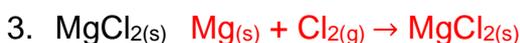
- The standard enthalpy change of formation ( $\Delta H_f^\ominus$ ) is the enthalpy change when one mole of a substance is formed from its elements in their standard states under standard conditions.
- Standard state is the normal, most pure stable state of a substance measured at a pressure of 100 kPa.

### Examples:



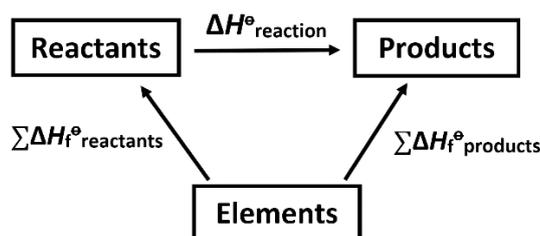
- Note that fractional coefficients are used because the equation is written for the formation of one mole of product.

**Exercise:** Write equations for the formation of one mole of the following compounds.



### Calculating $\Delta H$ for a reaction using $\Delta H_f^\ominus$ values

- Using the enthalpy cycle below and applying Hess's law, we arrive at the following equation:

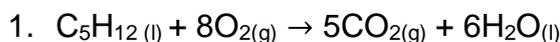


$$\Delta H^\ominus = \sum \Delta H_f^\ominus \text{ products} - \sum \Delta H_f^\ominus \text{ reactants}$$

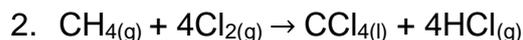
- $\Delta H_f^\ominus$  values can be found in table 12 of the data booklet.
- Elements have a standard enthalpy change of formation of zero.

**Exercise:** Calculate the enthalpy change for each of the reactions below, given the data in the table.

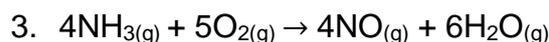
	$\Delta H_f^\ominus$ (kJ mol <sup>-1</sup> )
CO <sub>2(g)</sub>	-394
CH <sub>4(g)</sub>	-74.9
H <sub>2</sub> O <sub>(g)</sub>	-241.8
HCl <sub>(g)</sub>	-92.3
C <sub>5</sub> H <sub>12(l)</sub>	-173
CCl <sub>4(l)</sub>	-95.7
NH <sub>3(g)</sub>	-45.9
NO <sub>(g)</sub>	+90.3



$$\Delta H = [(5 \times -394) + (6 \times -286)] - (-173) = -3513 \text{ kJ mol}^{-1}$$



$$\Delta H = [(4 \times -92.3) + (4 \times -95.7)] - [(4 \times -74.9)] = -452 \text{ kJ mol}^{-1}$$

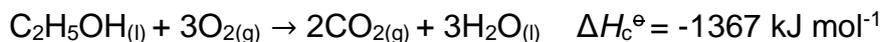


$$\Delta H = [(6 \times -241.8) + (4 \times +90.3)] - [(4 \times -45.9)] = -906 \text{ kJ mol}^{-1}$$

### Standard enthalpy change of combustion ( $\Delta H_c^\ominus$ )

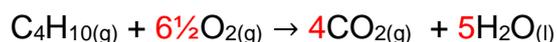
- The standard enthalpy change of combustion ( $\Delta H_c^\ominus$ ) is the enthalpy change when one mole of a substance is burned completely in oxygen under standard conditions.

#### Example:



- Note that the equation is balanced for one mole of reactant and the oxygen is assumed to be the excess reactant.

Write an equation for the complete combustion of one mole of butane,  $\text{C}_4\text{H}_{10}$ .

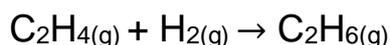


### Calculating $\Delta H$ using $\Delta H_c^\ominus$ values

- The equation used to calculate the  $\Delta H$  of a reaction using  $\Delta H_c^\ominus$  values is given below.

$$\Delta H^\ominus = \sum \Delta H_c^\ominus \text{ reactants} - \sum \Delta H_c^\ominus \text{ products}$$

**Example:** Determine the  $\Delta H$  of the following reaction using the values given in the table.



	$\Delta H_c^\ominus$
$\text{C}_2\text{H}_4$	-1411
$\text{H}_2$	-286
$\text{C}_2\text{H}_6$	-1561

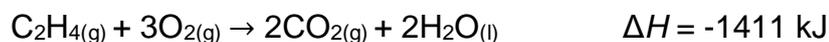
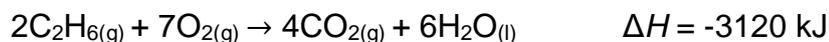
$$\Delta H^\ominus = (-1411 + -286) - (-1561)$$

$$\Delta H^\ominus = -136 \text{ kJ mol}^{-1}$$

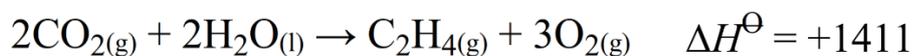
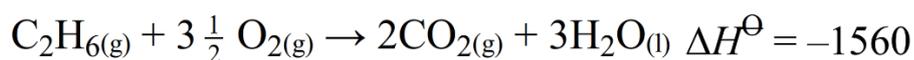
### Calculating the enthalpy change of a reaction that is the sum of multiple reactions with known enthalpy changes

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

#### Example:



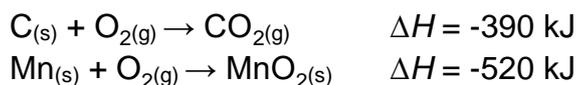
Calculate the enthalpy change for the following reaction:



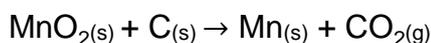
$$\Delta H = -1560 + 286 + 1411 = 137 \text{ kJ}$$

## Exercises:

1. Using the equations below:

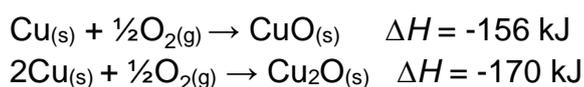


Determine the  $\Delta H$  (in kJ) for the following reaction.

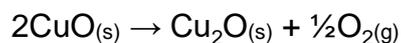


$$\Delta H = +130 \text{ kJ mol}^{-1}$$

2. Using the equations below:

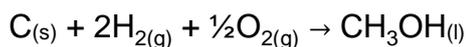


Determine the  $\Delta H$  (in kJ) for the following reaction.

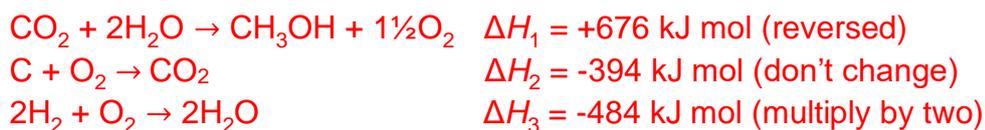
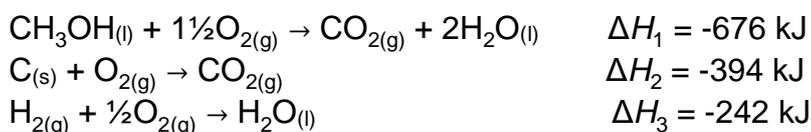


$$\Delta H = +142 \text{ kJ mol}^{-1}$$

3. Calculate the enthalpy change,  $\Delta H$ , for the reaction:



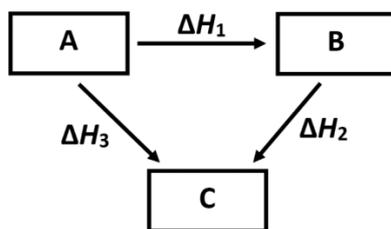
using Hess's Law and the following information.



$$\Delta H = -202 \text{ kJ}$$

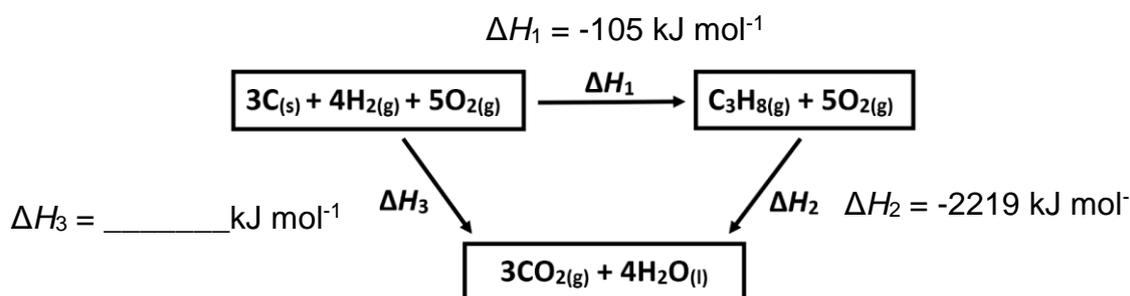
## Enthalpy cycles

- Enthalpy cycles are an application of Hess's law.
- Whichever route is taken to get to C, the enthalpy change is the same.



$$\Delta H_3 = \Delta H_1 + \Delta H_2$$

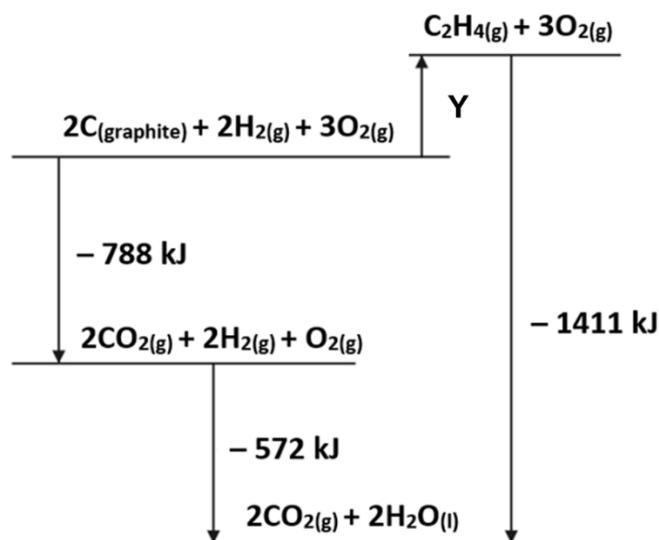
**Example 1** Using the values below for  $\Delta H_1$  and  $\Delta H_2$ , calculate  $\Delta H_3$



$$\Delta H_3 = \Delta H_1 + \Delta H_2$$

$$\Delta H_3 = -105 + -2219 = -2324 \text{ kJ}$$

**Example 2** Calculate the value of Y using the values given in the enthalpy cycle.



$$Y = -788 + (-572) + 1411 = +51 \text{ kJ mol}^{-1}$$

### 5.3 Bond enthalpies

#### Understandings:

- Bond-forming releases energy and bond-breaking requires energy.
- Average bond enthalpy is the energy needed to break one mole of a bond in a gaseous molecule averaged over similar compounds.

#### Applications and skills:

- Calculation of the enthalpy changes from known bond enthalpy values and comparison of these to experimentally measured values.
- Sketching and evaluation of potential energy profiles in determining whether reactants or products are more stable and if the reaction is exothermic or endothermic.
- Discussion of the bond strength in ozone relative to oxygen in its importance to the atmosphere.

#### Guidance:

- Bond enthalpy values are given in the data booklet in section 11.
- Average bond enthalpies are only valid for gases and calculations involving bond enthalpies may be inaccurate because they do not take into account intermolecular forces.

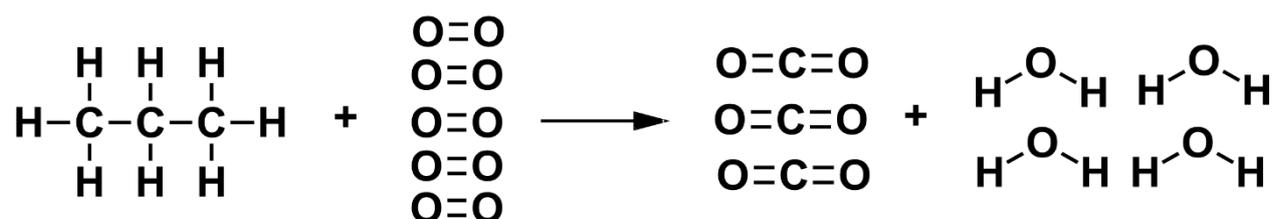
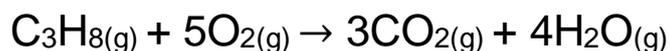
<b>Objective</b>	<b>I am confident with this</b>	<b>I need to review this</b>	<b>I need help with this</b>
Describe the energy changes involved in bond breaking and bond making			
Calculate enthalpy changes using average bond enthalpy values			
Outline why enthalpy change values calculated using average bond enthalpies differ from the actual value			
Describe the bonding in bonding and ozone			
Explain the importance of ozone in the atmosphere			
Explain the difference in the wavelength required to break the bonds in ozone and oxygen			

## 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:



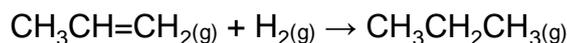
$$\Delta H = \sum (\text{bonds broken}) - \sum (\text{bonds made})$$

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

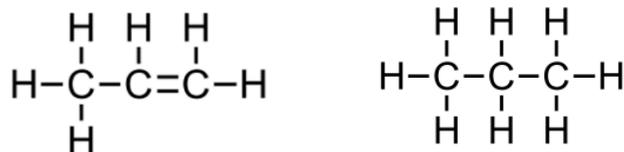
$$\Delta H = 6494 - 8528 = -2034 \text{ kJ mol}^{-1}$$

## Exercises:

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



Structures of propene (left) and propane (right):



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:

6 C-H (414 kJ)

1 C=C (614 kJ)

1 C-C (346 kJ)

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

$$\Delta H = 3880 - 4004$$

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

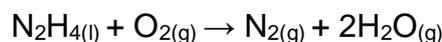
Reaction is exothermic (negative  $\Delta H$ )

Bonds formed:

8 C-H (414 kJ)

2 C-C (346 kJ)

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



Bonds broken:

1  $\times$  158 kJ

4  $\times$  391 kJ

1  $\times$  498 kJ

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

$$\Delta H = 2220 - 2797$$

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

Bonds formed:

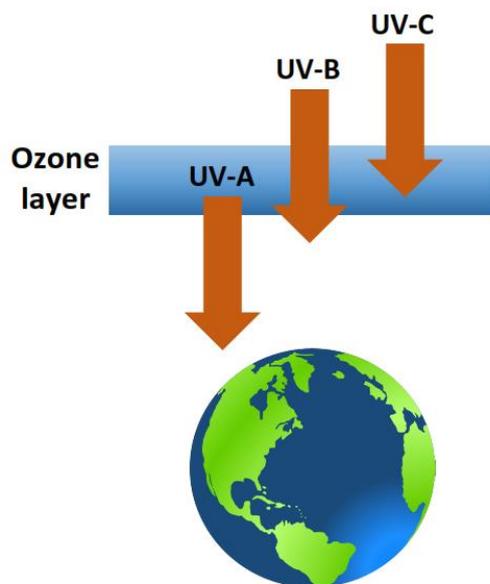
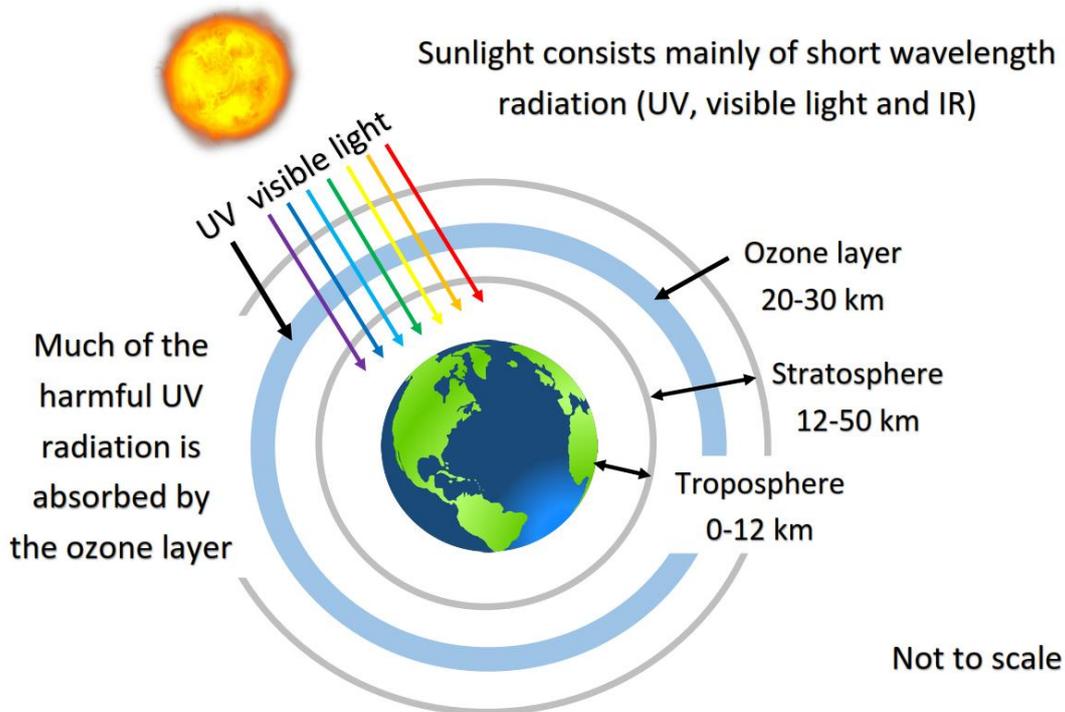
1  $\times$  945 kJ

4  $\times$  463 kJ

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.

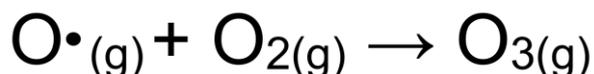
## Ozone in the atmosphere



Ozone absorbs UV radiation in the range of 200 nm-315 nm. This corresponds to the higher energy UV radiation (UV-B and UV-C) which can cause damage to living tissue.

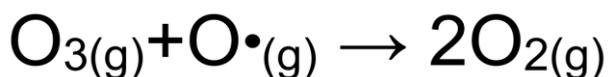
### Formation of ozone

- The oxygen molecule dissociates in the presence of high energy UV radiation to form two oxygen free radicals.
- The free radicals then react with molecular oxygen to form ozone.



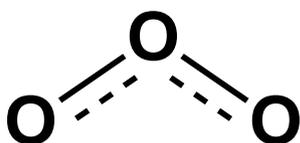
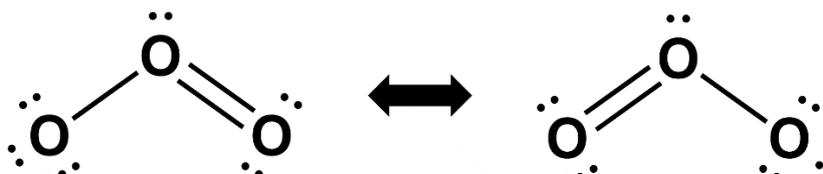
### Destruction of ozone

- The ozone molecule dissociates in the presence of lower energy UV radiation to form  $\text{O}_2$  and an oxygen free radical.
- The free radical reacts with ozone to form molecular oxygen.

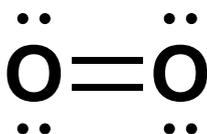


### Bond strength in ozone and oxygen

- Ozone,  $\text{O}_3$ , exists as two resonance structures.



The bonds in ozone are of equal length and strength; intermediate between a single and a double bond (bond order 1.5).



The oxygen molecule has a double covalent bond between the oxygen atoms (bond order 2).

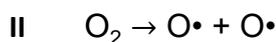
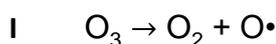
- The double bond in O<sub>2</sub> is stronger than the bonds in O<sub>3</sub> – it requires more energy to break the bond in O<sub>2</sub> than in O<sub>3</sub>
- The double bond in O<sub>2</sub> is broken by UV radiation with a wavelength of <242 nm forming two oxygen free radicals.
- The bond in O<sub>3</sub> is broken by UV radiation with a wavelength of <330 nm forming an oxygen molecule and an oxygen free radical.

Molecule	Bond order	λ UV radiation (nm)
O <sub>2</sub>	2	<242
O <sub>3</sub>	1.5	<330

- Stronger double bond in O<sub>2</sub> requires higher energy UV radiation (shorter wavelength) to break.
- Weaker bond in O<sub>3</sub> requires lower energy (longer wavelength) to break.

### Exercises:

1. The following reactions take place in the ozone layer by the absorption of ultraviolet light.



State and explain, by reference to the bonding, which of the reactions, **I** or **II**, requires a shorter wavelength of UV light.

Reaction II requires a shorter wavelength (higher energy). O<sub>2</sub> has a double bond (bond order of 2) and O<sub>3</sub> has an intermediate bond between a double and single bond (bond order of 1.5). The double bond in oxygen is stronger than the intermediate bonding in ozone.

2. Explain why UV radiation with a wavelength of between 300 and 330 nm contributes to photochemical decomposition of ozone but not to the formation of ozone from oxygen.

The double bond in O<sub>2</sub> is stronger than the bonds in O<sub>3</sub>  
Bond order in O<sub>2</sub> is 2 and bond order in O<sub>3</sub> is 1.5

The energy of light depends on wavelength (higher energy has lower wavelength)

The energy of light with  $\lambda = 300\text{--}330$  nm is insufficient to break the bond in  $\text{O}_2$  but sufficient to break bonds in  $\text{O}_3$