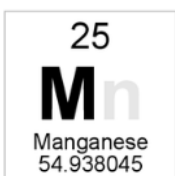
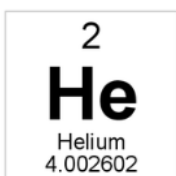
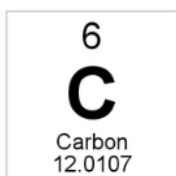
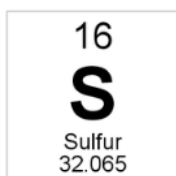
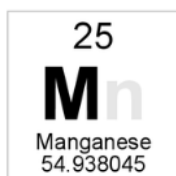


# Energetics SL (answers)

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IB CHEMISTRY SL



## 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:

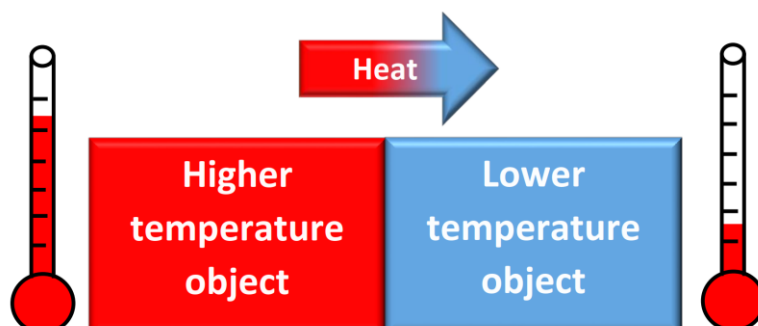
- Calculation of the heat change when the temperature of a pure substance is changed using  $q=mc\Delta T$
- A calorimetry experiment for an enthalpy of reaction should be covered and the results evaluated.

### Guidance:

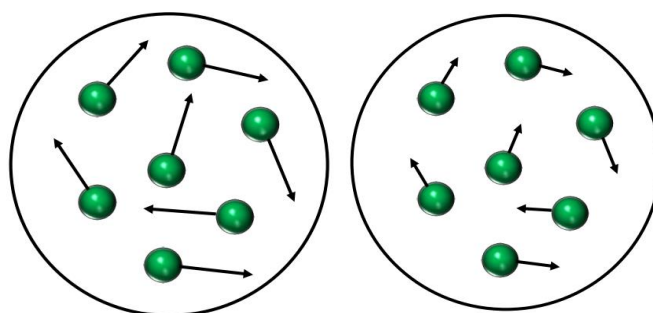
- Enthalpy changes of combustion ( $\Delta H_c^\ominus$ ) and formation ( $\Delta H_f^\ominus$ ) should be covered.
- Consider reactions in aqueous solution and combustion reactions.
- Standard state refers to the normal, most pure stable state of a substance measured at 100 kPa. Temperature is not a 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.

## Heat and temperature

- Heat is the transfer of energy from a higher temperature object to a lower temperature object (measured in J).
- As shown in the diagram below, heat always flows from the hotter object to the cooler object.



- Absolute temperature (in K) is the measure of the average kinetic energy of the particles in a substance.
- Which sample of gas below has the highest temperature? The gas particles in the sample on the left have a higher average kinetic and a higher absolute temperature.



## 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:

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

## Exercises

1) Define the terms heat and absolute temperature.

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

Heat is the transfer of energy from a higher temperature object to a lower temperature object.

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

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 the higher temperature object to the lower temperature object). After one hour, the temperature of the iron and the water will be the same – they will have reached thermal equilibrium.

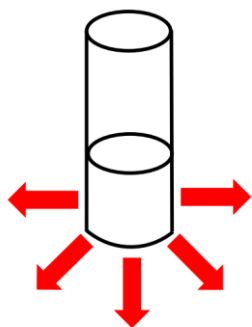
## Enthalpy change ( $\Delta H$ )

- Enthalpy is the heat content of a substance.
- Enthalpy change ( $\Delta H$ ) is the amount of heat absorbed or released in a chemical reaction.
- The standard enthalpy change of reaction ( $\Delta H^\ominus$ ) is the enthalpy change measured under standard conditions (100 kPa and 298 K).

## Exothermic and endothermic reactions

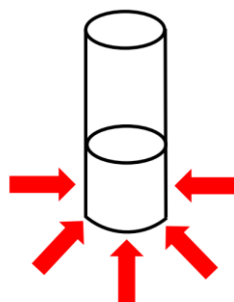
- Chemical reactions can be classified as exothermic or endothermic depending on whether heat is absorbed or released.

### Exothermic reaction



Heat flows from  
the system to the  
surroundings

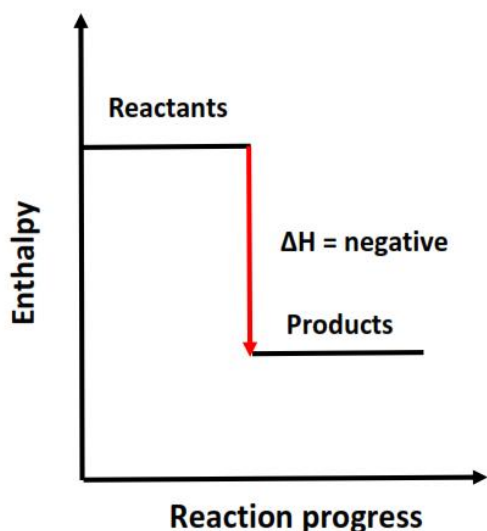
### Endothermic reaction



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).

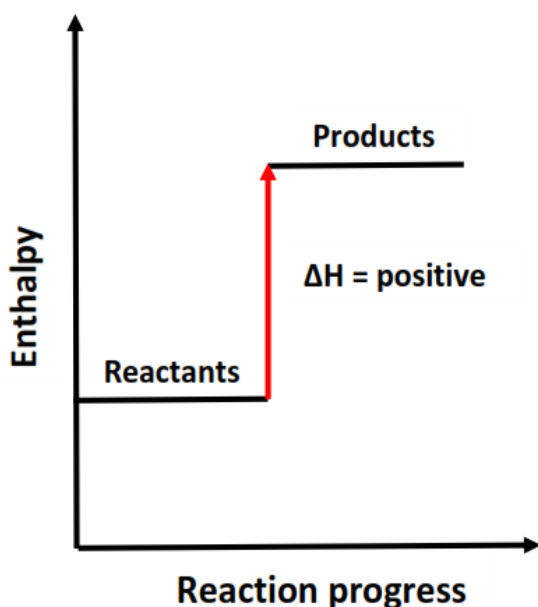
The enthalpy level diagram below is for an exothermic reaction.



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

- 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.
- Examples of exothermic reactions include combustion and neutralisation.

The enthalpy level diagram below is for an endothermic reaction



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

- 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.

- Examples of endothermic reactions include photosynthesis and thermal decomposition reactions.

### Summary:

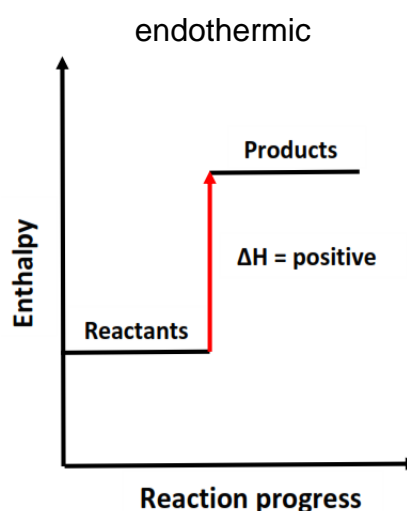
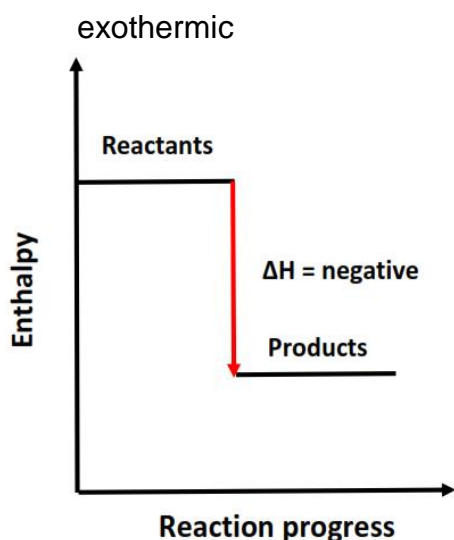
	Exothermic reaction	Endothermic reaction
Enthalpy change ( $\Delta H$ )	negative	positive
Which has lower enthalpy?	products	reactants
Which is more energetically stable?	products	reactants
Is heat absorbed or released?	released	absorbed

### 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, solutions with a concentration of 1.00 mol dm<sup>-3</sup>) with everything in its standard state (the standard state is the normal, most pure stable state of a substance measured at a pressure of 100 kPa).

2) Sketch an enthalpy level diagram for an exothermic reaction and an endothermic reaction.



3) Which statements about exothermic reactions are correct?

- I. They have negative  $\Delta H$  values.
- II. The products have a lower enthalpy than the reactants.
- III. The products are more energetically stable than the reactants.

- A. I and II only                      B. I and III only  
C. II and III only                    D. I, II and III

D

4) Which statements are correct for an endothermic reaction?

- I. The system absorbs heat.
- II. The enthalpy change is positive.
- III. The products are more energetically stable than the reactants

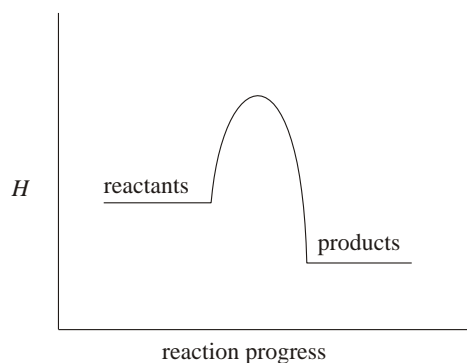
- A. I and II only                      B. I and III only  
C. II and III only                    D. I, II and III

A

5) 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**

6) According to the enthalpy level diagram below, what is the sign for  $\Delta H$  and what term is used to refer to the reaction?

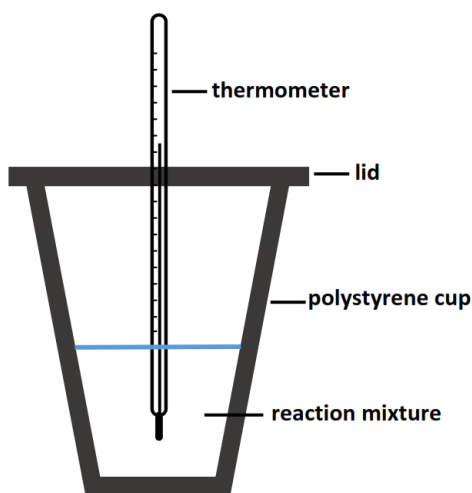




	$\Delta H$	reaction
A.	positive	endothermic
B.	negative	exothermic
C.	positive	exothermic
D.	negative	endothermic

**B**

### Calculating enthalpy changes



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

$$q = mc\Delta T$$

$q$  = heat (J)

$m$  = mass of water (or solution)

$c$  = specific heat capacity ( $J/g^{\circ}C$ )

$\Delta T$  = change in temperature

#### Some important definitions:

**Standard enthalpy change of neutralization ( $\Delta H^{\ominus}_n$ )** - the enthalpy change when a strong acid and base are reacted together to form one mole of water under standard conditions (with everything in their standard states).

**Standard enthalpy change of combustion ( $\Delta H^{\ominus}_c$ )** - the enthalpy change when one mole of a compound is burned in excess oxygen under standard conditions (with everything in their standard states).

**Standard conditions:** 100 kPa, 298 K

**Standard state** is the normal, most pure stable state of a substance measured at a pressure of 100 kPa.

**Density of water:**  $1 \text{ g cm}^{-3}$       **Specific heat capacity of water:**  $4.18 \text{ J g}^{-1}\text{C}^{-1}$

**Exercises:**

Calculate the enthalpy change of neutralization for the following reactions.

1.  $200.0 \text{ cm}^3$  of  $0.400 \text{ mol dm}^{-3}$  HCl is added to with 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  $27.7 \text{ }^\circ\text{C}$ .



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

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

$$q = mc\Delta T$$

$$q = 400 \times 4.18 \times (27.7 - 25.1)$$

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

$$\text{For one mole of water produced: } 4347.2 \div 0.0800 = 54340 \text{ J mol}^{-1} = -54.3 \text{ kJ mol}^{-1}$$

2.  $400.0 \text{ cm}^3$  of  $0.500 \text{ mol dm}^{-3}$  hydrochloric acid was reacted with  $400.0 \text{ cm}^3$  of  $1.00 \text{ mol dm}^{-3}$  potassium hydroxide and the temperature rose by  $3.20 \text{ }^\circ\text{C}$ .



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

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

HCl is the limiting reactant – ratio of HCl to  $\text{H}_2\text{O}$  is 1:1

$$q = mc\Delta T$$

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

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

$$\text{For one mole of water produced: } 10700.8 \div 0.200 = 53504 \text{ J mol}^{-1} = -53.5 \text{ kJ mol}^{-1}$$

3. 250.0 cm<sup>3</sup> of NaOH and 500.0 cm<sup>3</sup> of HCl are mixed. Both solutions had concentrations of 0.500 mol dm<sup>-3</sup>. The temperature increased by 2.10°C.



$$n(\text{NaOH}) = 0.500 \times (250.0 \div 1000) = 0.125 \text{ mol}$$

$$n(\text{HCl}) = 0.500 \times (500.0 \div 1000) = 0.250 \text{ mol}$$

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

$$q = mc\Delta T$$

$$q = 750.0 \times 4.18 \times (2.10)$$

$$q = 6583.5 \text{ J}$$

$$\text{For one mole of water produced: } 6583.5 \div 0.125 = 52668 \text{ J mol}^{-1} = -52.7 \text{ kJ mol}^{-1}$$

The literature value for the standard enthalpy change of neutralization is -57.0 kJ zmol<sup>-1</sup>. For each of the above examples, calculate the percentage error and suggest possible sources of error.

Percentage error for reaction a:

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

Percentage error for reaction b:

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

Negative percentage error means that the experimental value is lower than the theoretical/literature value.

Percentage error for reaction c:

$$\% \text{ error} = \frac{(-52.7) - (-57.0)}{(-57.0)} \times 100 = -8\%$$

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>).

Calculate the enthalpy change of combustion ( $\Delta H_c$ ) for the following reactions.

1. A 1.00 g sample of octane ( $C_8H_{18}$ ) is burned in a bomb calorimeter containing 1200.0 cm<sup>3</sup> of water at an initial temperature of 25.0°C. After the reaction, the final temperature of the water is 33.0°C.

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

$$q = mc\Delta T$$

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

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

$$\text{For one mole of } C_8H_{18}: 40128 \div 8.75 \times 10^{-3} = 4586057.143 \text{ J mol}^{-1}$$

$$= - 4.59 \times 10^3 \text{ kJ mol}^{-1}$$

2. A 0.500 g sample of naphthalene ( $C_{10}H_8$ ) is burned in a calorimeter containing 650.0 cm<sup>3</sup> of water at an initial temperature of 20.0°C. After the reaction, the final temperature of the water is 26.4°C.

$$n(C_{10}H_8) \quad n = m \div M = 0.500 \div 128.18 = 3.90 \times 10^{-3}$$

$$q = mc\Delta T$$

$$q = 650.0 \times 4.18 \times (26.4 - 20.0)$$

$$q = 17388.8 \text{ J}$$

$$\text{For one mole of } C_{10}H_8: 17388.8 \div 3.90 \times 10^{-3} = 4458666.667 \text{ J mol}^{-1}$$

$$= - 4.46 \times 10^3 \text{ kJ mol}^{-1}$$

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

Initial temperature of water:	20.0 °C
Final temperature of water:	55.0 °C
Mass of ethanol burned:	1.78 g

Calculate the enthalpy change of combustion per mole of ethanol.

$$n(C_2H_5OH) \quad n = m \div M = 1.78 \div 46.08 = 0.0386$$

$$q = mc\Delta T$$

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

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

$$\text{For one mole of C}_2\text{H}_5\text{OH: } 14630 \div 0.0386 = 379015.544 \text{ J mol}^{-1} = -379 \text{ kJ mol}^{-1}$$

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>10</sub> H <sub>8</sub>	-5140
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 a:

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

Percentage error for reaction b:

$$\% \text{ error} = \frac{(-4460) - (-5140)}{(-5140)} \times 100 = -13\%$$

Percentage error for reaction c:

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

Negative percentage error means that the experimental value is lower than the theoretical/literature value.

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

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

Definition: The enthalpy change when one mole of a substance is formed from its elements in their standard states under standard conditions.

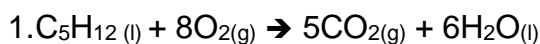
- Elements have a standard enthalpy change of formation of zero.

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

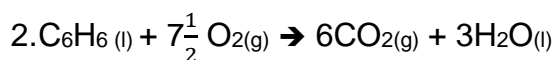
### Exercise:

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

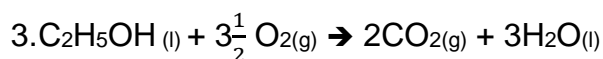
	$\Delta H_f^\ominus$ (kJmol <sup>-1</sup> )
CO <sub>2</sub> (g)	-394
H <sub>2</sub> O(l)	-286
C <sub>2</sub> H <sub>5</sub> OH(l)	-278
C <sub>5</sub> H <sub>12</sub> (l)	-173
C <sub>6</sub> H <sub>6</sub> (l)	+49.0



$$\Delta H^\ominus = (5 \times -394 + 6 \times -286) - (-173) = (-3686) - (-173) = -3513 \text{ kJ mol}^{-1}$$



$$\Delta H^\ominus = (6 \times -394 + 3 \times -286) - (+49.0) = (-3222) - (+49.0) = -3271 \text{ kJmol}^{-1}$$

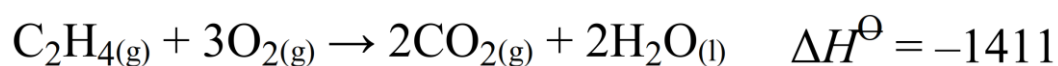
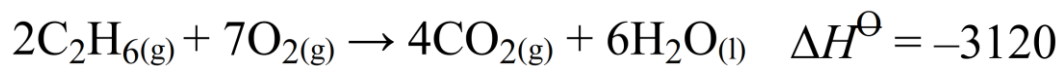


$$\Delta H^\ominus = (2 \times -394 + 3 \times -286) - (-278) = (-1646) - (-278) = -1368 \text{ kJmol}^{-1}$$

### Hess's law

- The enthalpy change for a chemical reaction is independent of the route by which the chemical reaction occurs.

### Example:



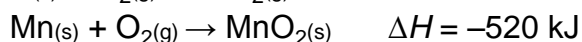
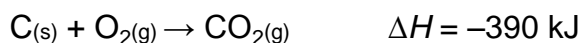
Calculate the enthalpy change for the following reaction:



See video for solution.

### Exercises:

1. Using the equations below:

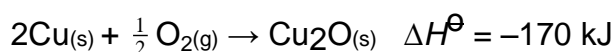
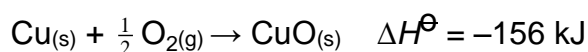


what is  $\Delta H$  (in kJ) for the following reaction?

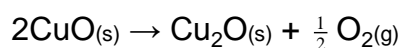


$$\Delta H = +130 \text{ kJ mol}^{-1}$$

2. Using the equations below

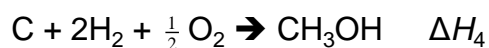


what is the value of  $\Delta H^\ominus$  (in kJ) for the following reaction?

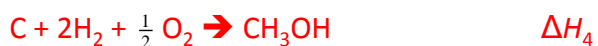
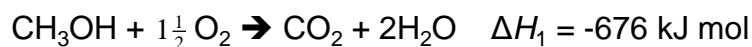


$$\Delta H = +142 \text{ kJ mol}^{-1}$$

3. Calculate the enthalpy change,  $\Delta H_4$  for the reaction



using Hess's Law and the following information.



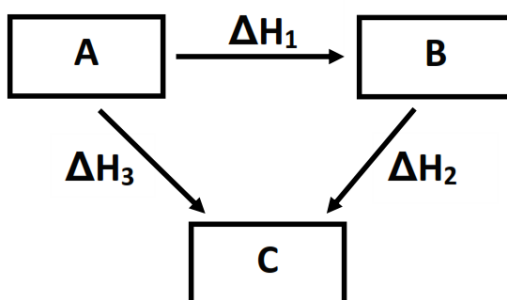
$$\Delta H_4 = -202 \text{ kJ}$$

## Enthalpy cycles

- Enthalpy cycles are an application of Hess's law.

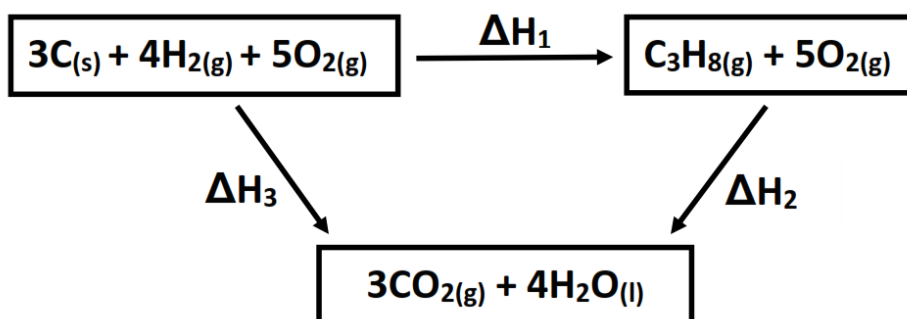


- Whichever route is taken to get to the products C, the enthalpy change is the same.



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

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



$$\Delta H_1 = -105 \text{ kJmol}^{-1}$$

$$\Delta H_2 = -2219 \text{ kJmol}^{-1}$$

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

$$\Delta H_3 = -105 + -2219$$

$$\Delta H_3 = -2324 \text{ kJmol}^{-1}$$

### 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.

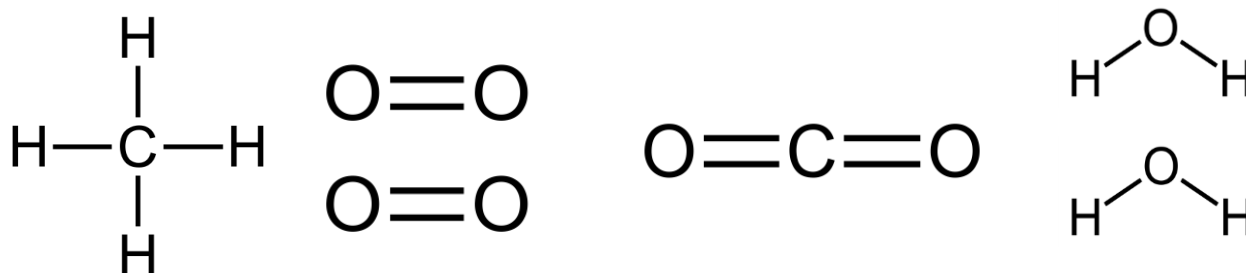
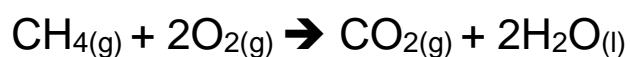
**Bond enthalpies**

- 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:



See video for solution.

### Exercises:

1. Which statement about bonding is correct?

- A. Bond breaking is endothermic and requires energy.
- B. Bond breaking is endothermic and releases energy.
- C. Bond making is exothermic and requires energy.
- D. Bond making is endothermic and releases energy.

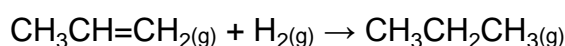
A

2. Which is correct about energy changes during bond breaking and bond formation?

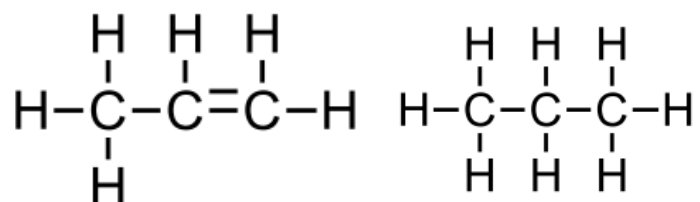
	Bond breaking	Bond formation
A.	exothermic and $\Delta H$ positive	endothermic and $\Delta H$ negative
B.	exothermic and $\Delta H$ negative	endothermic and $\Delta H$ positive
C.	endothermic and $\Delta H$ positive	exothermic and $\Delta H$ negative
D.	endothermic and $\Delta H$ negative	exothermic and $\Delta H$ positive

C

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



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



- (a) Enthalpy changes can be determined using average bond enthalpies. Define the term *average bond enthalpy*.

The energy needed to break 1 mole of bonds in a gaseous molecule (average calculated from a range of compounds).

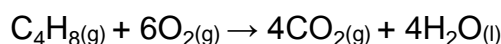
- (b) Determine a value for the hydrogenation of propene using information from Table 12 of the Data Booklet. State and explain if the reaction is exothermic or endothermic.

**Bonds broken:**  
 6 C-H (414 kJ)      1 H-H (436 kJ)  
 1 C=C (614 kJ)  
 1 C-C (346 kJ)

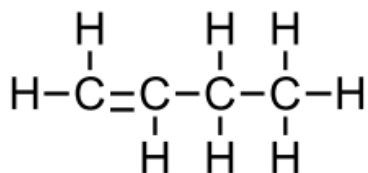
**Bonds made:**  
 8 C-H (414 kJ)  
 2 C-C (346 kJ)

$\Delta H = \sum(\text{bonds broken}) - \sum(\text{bonds made})$   
 $\Delta H = 3880 - 4004$   
 $\Delta H = -124 \text{ kJ}$   
 Reaction is exothermic (negative  $\Delta H$ )

4. Use the information from Table 12 of the Data Booklet to calculate the enthalpy change for the complete combustion of but-1-ene, according to the following equation.



Structure of but-1-ene:



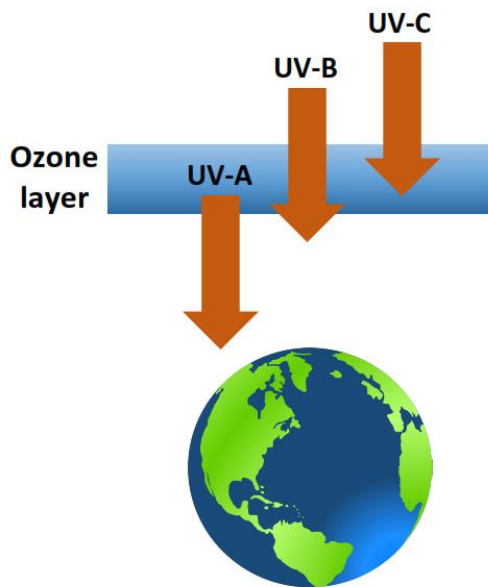
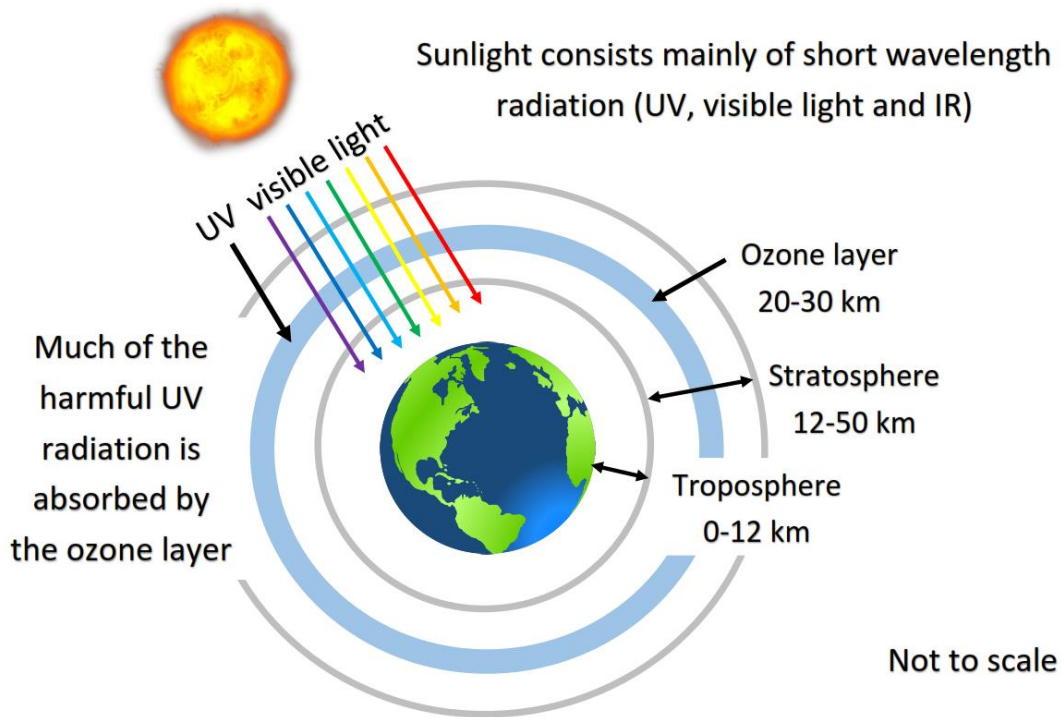
**Bonds broken**  
 1 C=C, 2 C-C, 8 C-H, 6 O=O  
 $(614) + (2 \times 346) + (8 \times 414) + (6 \times 498) = 7606 \text{ kJ}$

**Bonds made**  
 8 C=O, 8 O-H  
 $(8 \times 804) + (8 \times 463) = 10136 \text{ kJ}$

$\Delta H = \sum(\text{bonds broken}) - \sum(\text{bonds made})$   
 $\Delta H = 7606 - 10136$   
 $\Delta H = -2530 \text{ kJ}$   
 Reaction is exothermic (negative  $\Delta H$ )

5. State two reasons why enthalpy changes calculated using average bond enthalpies can be inaccurate.
- Average bond enthalpies apply to the gaseous state only.
  - The actual values for the specific bonds may be different to the average values.

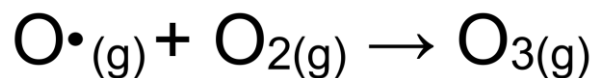
### Ozone in the atmosphere



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

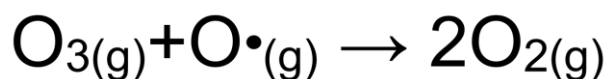
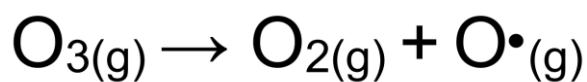
## Formation of ozone

- 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.

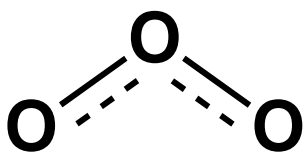


### **Destruction of ozone**

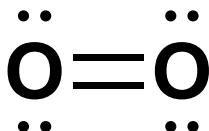
- The ozone molecule dissociates in the presence of low energy UV radiation to form O<sub>2</sub> and an oxygen free radical.
- The free radical then reacts with ozone to form molecular oxygen.



### **Bond strength in ozone and oxygen**



The bonds in ozone are intermediate in strength and length 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 <242nm forming two oxygen free radicals.
- The bond in O<sub>3</sub> is broken by UV radiation with a wavelength of <330nm forming an oxygen molecule and an oxygen free radical.

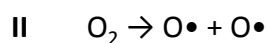
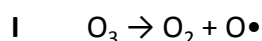
Molecule	Bond order	$\lambda$ UV radiation (nm)
O <sub>2</sub>	2	<242
O <sub>3</sub>	1.5	<330

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

Reaction II requires a shorter wavelength (higher energy).  $\text{O}_2$  has a double bond (bond order of 2) and  $\text{O}_3$  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  $\lambda = 300 - 330 \text{ nm}$  contributes to photochemical decomposition of ozone but not to the formation of ozone from oxygen.

The double bond in  $\text{O}_2$  is stronger than the bonds in  $\text{O}_3$   
Bond order in  $\text{O}_2$  is 2 and bond order in  $\text{O}_3$  is 1.5  
The energy of light depends on wavelength (higher energy has lower wavelength)  
The energy of light with  $\lambda = 300-330 \text{ nm}$  is insufficient to break the bond in  $\text{O}_2$  but sufficient to break bonds in  $\text{O}_3$