FACTFILE:
GCE CHEMISTRY

2.8 ENERGETICS

Rewarding Learning

Learning Outcomes

Students should be able to:

2.8.1 define the terms exothermic and endothermic and understand that chemical reactions are usually accompanied by heat changes;

2.8.2 recall standard conditions as 100 kPa and 298 K;

2.8.3 define the term standard enthalpy change, $\Delta H^{\ominus}$;

2.8.4 construct a simple enthalpy level diagram;

2.8.5 define the standard enthalpy of combustion, formation and neutralisation, namely $\Delta_c H^{\ominus}$, $\Delta_f H^{\ominus}$ and $\Delta_n H^{\ominus}$;

2.8.6 recall experimental methods to determine enthalpy changes;

2.8.7 calculate enthalpy changes from experimental data using the equation $q = mc\Delta T$;

2.8.8 demonstrate an understanding of the principle of conservation of energy and define Hess's Law;

2.8.9 construct enthalpy cycles using Hess's Law;

2.8.10 calculate enthalpy changes indirectly using Hess's Law;

2.8.11 define the term average bond enthalpy and calculate the enthalpy change of a reaction using average bond enthalpies;

2.8.12 calculate average bond enthalpies given enthalpy changes of reaction; and

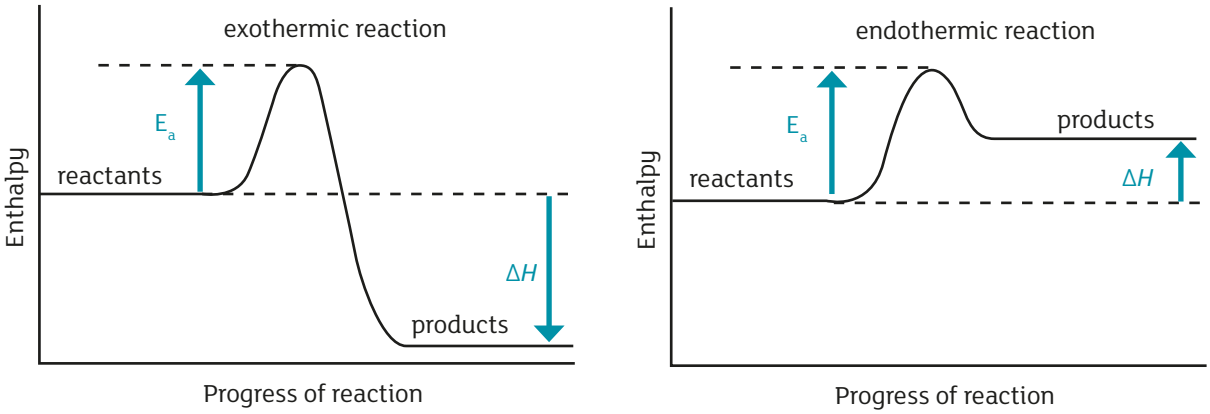
2.8.13 explain why enthalpy changes of reaction calculated using average bond enthalpies differ from those determined using Hess's Law.

Chemical Energetics

Chemical reactions can either release heat energy to their surroundings, exothermic, or heat energy can be transferred to them from the surroundings, endothermic. The amount of heat taken in or given out in a reaction varies with conditions. As reactions usually take place at atmospheric pressure, this leads to the definition of enthalpy, which is described as a heat change at constant pressure.

It is important to note that you cannot measure the actual enthalpy of a substance; you can only measure an enthalpy change. In order to compare the enthalpy changes of a various reactions we must use standard conditions, given as 100 kPa and 298 K.

Enthalpy changes can be represented on enthalpy level diagrams as shown below:



Reactions need an input of energy to get them started because reactant bonds have to be broken before new product bonds can form. The energy required to do this is known as the activation energy of the reaction and is represented by E_a.

- An **endothermic reaction** is one in which the enthalpy of the products is greater than the enthalpy of the reactants.
- An **exothermic reaction** is one in which the enthalpy of the products is less than the enthalpy of the reactants.

Exothermic enthalpy changes have a **negative** sign, **endothermic** enthalpy changes have a **positive** sign. There are a number of standard enthalpy changes which are encountered at GCE level. A standard enthalpy change ΔH^{Θ} is the change in heat energy at constant pressure.

1. Standard enthalpy change of formation : $\Delta_f H^{\Theta}$
The enthalpy change when one mole of a compound is formed from its elements under standard conditions. For example,

$C(s) + 2H_2(g) \rightarrow CH_4(g)$ $\quad\quad$ $\Delta_f H^{\Theta} = 650\ kJ\ mol^{-1}$

$\frac{1}{2}H_2(g) + \frac{1}{2}Cl_2(g) \rightarrow HCl(g)$ $\quad\quad$ $\Delta_f H^{\Theta} = -22\ kJ\ mol^{-1}$

2. Standard enthalpy change of combustion : $\Delta_c H^{\Theta}$
The enthalpy change when one mole of a substance is completely burnt in oxygen under standard conditions. For example,

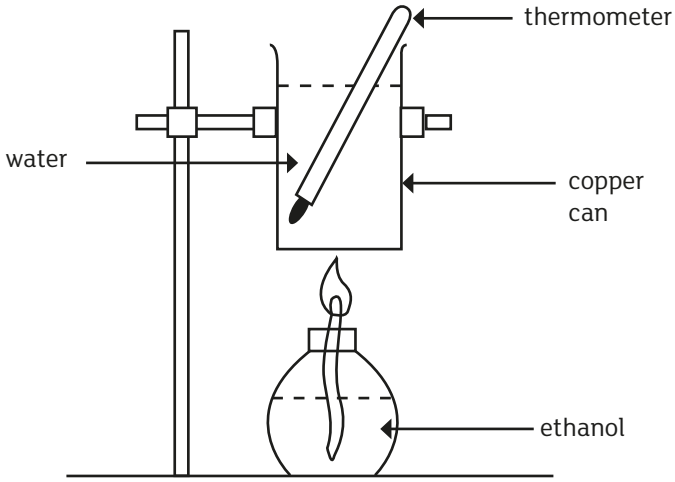
$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$ $\quad\quad$ $\Delta_c H^{\Theta} = -1368\ kJ\ mol^{-1}$

3. Standard enthalpy change of neutralisation : $\Delta_n H^{\Theta}$
The enthalpy change when one mole of water is produced in a neutralisation reaction under standard conditions. For example,

$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$ $\quad\quad$ $\Delta_n H^{\Theta} = -57.1\ kJ\ mol^{-1}$

Experimental methods to determine enthalpy changes

Calorimetry is the experimental determination of enthalpy changes. For example, to determine the enthalpy of combustion of a fuel, the heat evolved can be used to heat a known volume of water:



We cannot measure heat directly but we can calculate it if we take the temperature, mass of the substance and type of substance into account. The amount of heat needed to raise the temperature of 1g of a substance by 1K (which is the same as 1 °C) is called the **specific heat capacity**, c. The value for water is 4.18 J g^{-1} K^{-1} (or 4.2 is sometimes used). The amount of heat is given by:

Heat (q) = mass of water (m) x specific heat capacity (c) x temperature rise (ΔT)

This can be simplified as q = mcΔT

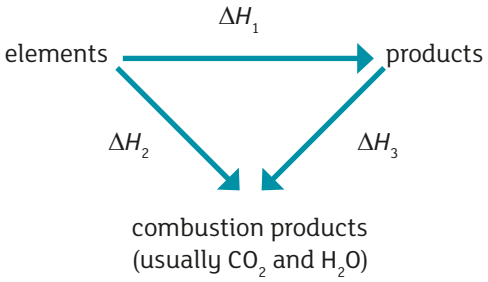
This is converted to an enthalpy change by scaling to one mole of the fuel. Other enthalpy changes, such as the enthalpy of neutralisation, can be experimentally determined by measuring the temperature of the reaction mixture directly.

Other enthalpy changes cannot be determined by experiment. They can be determined indirectly by applying Hess's Law:

Hess's Law states that the enthalpy change for a reaction is independent of the route taken, provided the initial and final conditions are the same. Hess's Law is an application of the principle of conservation of energy which states that energy cannot be created or destroyed but it can change from one form into another.

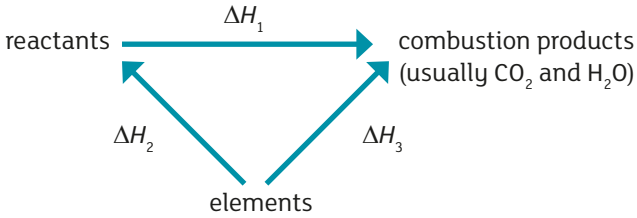
Enthalpy changes can be calculated using a Hess's Law cycle. There are two common types

1. Enthalpy cycles using enthalpies of combustion. The general cycle looks like:



Applying Hess's Law: $\Delta H_1 = \Delta H_2 - \Delta H_3$

2. Enthalpy cycles using enthalpies of formation. The general cycle looks like:



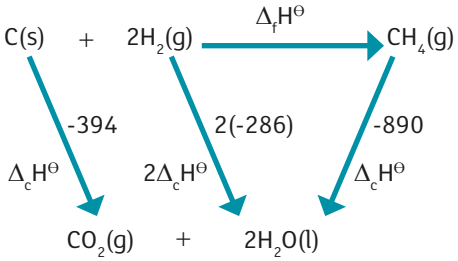
Applying Hess's Law: $\Delta H_1 = -\Delta H_2 + \Delta H_3$

Example: Calculate the standard enthalpy change of formation of methane given the following enthalpy changes of combustion:

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$ $\quad \Delta_c H^\Theta = -890$ kJ mol^{-1}

$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$ $\quad \Delta_c H^\Theta = -286$ kJ mol^{-1}

$C(s) + O_2(g) \rightarrow CO_2(g)$ $\quad \Delta_c H^\Theta = -394$ kJ mol^{-1}



Applying Hess's Law: $\Delta_f H^\Theta(CH_4(g)) = +(-394) + 2(-286) - (-890) = -76$ kJ mol^{-1}

Enthalpy changes for reactions involving covalent molecules can be determined by considering bond enthalpies, defined as the energy required to break one mole of a given bond averaged over many compounds. A chemical reaction can be viewed as a breaking of all the reactant bonds to give separate atoms followed by the forming of products as| atoms recombine. The enthalpy change for the reaction is given by:

$\Delta H = \Sigma$ bond enthalpies of reactants $- \Sigma$ bond enthalpies of products

Σ means summation i.e. all the bond energies are added together

Enthalpy changes calculated using bond enthalpies differ from those determined using Hess's Law as the bond enthalpy values are averaged across compounds containing the bond.

Credits

Pg. 2 *cc* © Taken from Chemistry for CCEA AS Level;

Revision Questions

1 Many enthalpy changes, including ethalpies of combustion, can be determined by experiment. Others can be deduced using Hess's Law.

(a) (i) State Hess's Law.

.....

..... [2]

(ii) Define the term **standard enthalpy of combustion**.

.....

.....

..... [2]

(iii) Use the standard enthalpies of information given in the table below to calculate the enthalpy change for the complete combustion of methane:

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$$

Compound	$\Delta_f H^{\ominus}$/kJ mol^{-1}
CO_2	-394
H_2O	-286
CH_4	-75

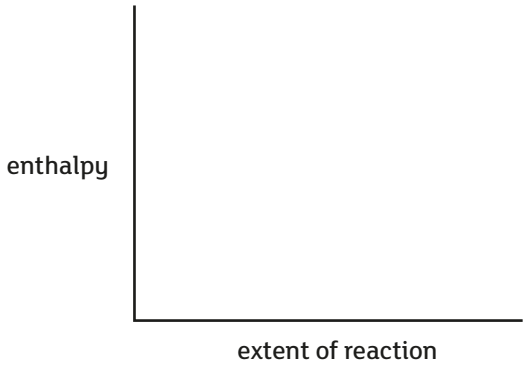
.....

.....

..... [2]

(b) Enthalpy changes can also be calculated using average bond enthalpies. When bond enthalpies were used to estimate the enthalpy of reaction for the complete combustion of methane, it was found to have a value of –698 kJ mol^{-1}.

(i) Complete the enthalpy level diagram for this reaction.



[1]

(ii) What is meant by the term **average bond enthalpy?**

.....

..... [2]

(iii) Using the estimated enthalpy of reaction and the bond enthalpies given in the table below, calculate the bond enthalpy of the C–H bond in methane.

bond	bond enthslpy/kJ mol^{-1}
C═O	743
O═O	496
O—H	463

.....

.....

.....

..... [3]

2 0.47 g of hydrocarbon was completely burnt in air. The heat produced raised the temperature of 200g of water by 28.2 °C. The standard enthalpy of combustion of the hydrocarbon is −2220 kJ mol^{-1}. The specific heat capacity of water is 4.2 $Jg^{-1}\,°C^{-1}$.

Which one of the following is the molar mass of the hydrocarbon?

A. 40

B. 44

C. 185

D. 199

3 Some "ice packs" used to treat sports injuries contain ammonium nitrate and water. A capsule containing the water is broken, the ammonium nitrate dissolves and the temperature falls.

(a) The following results were obtained when ammonium nitrate was added to some water.

Mass of water = 100 g
Mass of ammonium nitrate = 5.0 g
Initial temperature = 25.0 °C
Final temperature = 24.1 °C

Specific heat capacity of water = 4.2 $J\,g^{-1}\,°C^{-1}$

(i) Calculate the enthalpy change taking place.

.....

..... [2]

(ii) How many moles of ammonium nitrate were used?

.....

..... [1]

(iii) Calculate the molar enthalpy change for dissolving the ammonium nitrate in 100 g of water.

.....

..... [1]

(b) One manufacturer makes an "ice pack" containing 120 g of water.
What mass of ammonium nitrate will be needed to produce a temperature fall 25.0 °C?

.....

.....

..... [2]