THERMOCHEMISTRY

(CHEMICAL ENERGETICS / THERMODYNAMICS)

Competency:

The learner analyses and evaluates thermodynamic principles and processes to predict the spontaneity and feasibility of chemical reactions and processes.

This means they should be able to use ideas such as energy changes, heat transfer, and reaction feasibility to explain what happens in real-life situations and predict whether certain processes can occur naturally.

INTRODUCTION

Thermochemistry is the branch of chemistry that studies the **heat energy changes** that occur during (accompany) chemical reactions and physical processes. In everyday life, many activities involve the release or absorption of heat. For example, when charcoal burns in a stove to cook food, heat is released; when ice melts in a drink, heat is absorbed from the surroundings. Understanding how and why these energy changes occur helps learners explain natural phenomena and make informed decisions about energy use.

Thermochemistry investigates what determines the feasibility of a reaction and how far a reaction can go. It therefore explains to us why chemical changes and physical changes as well happen.

Some changes are spontaneous in that they occur by themselves without external help. However, some changes do not occur unless we make them happen.

Spontaneous changes have a driving force taking them in a certain direction and to send such a process in the reverse direction, energy has to be expended. e.g. potassium metal reacts violently when exposed to water to form a hydroxide and a gas. To reverse this process is very difficult.

Some spontaneous process release heat energy while others absorb heat energy. Thus, heat changes are important in the investigation of chemical and physical changes.

The study of these heat changes at constant pressure is called the *enthalpy or heat change* but since the study of heat changes is compared at identical /standard conditions, it's then referred to as *standard enthalpy change*.

Enthalpy changes are also important in the investigation of chemical and physical changes.

Forms of Energy

There are many forms of energy i.e. heat, light, chemical, nuclear etc but all are of only two kinds of i.e. kinetic energy and potential energy.

- Kinetic energy is energy due to molecular motion.
- Potential energy is energy due to position or associated with chemical bonds or arrangement of an object component parts.

Heat energy is a form of kinetic energy associated with the motion of atoms or molecules.

Energy of chemical bonds is a form of potential energy arising from the position of atoms and molecules with respect to one another.

Thus, matter possesses kinetic energy and potential energy as the internal energy of matter. i.e. Internal energy of matter = K.E + P.E

When matter changes from one type to another, in a change of state (physical change) or chemical change, the position of the particles relative to one another also changes.

Energy is taken in (absorbed) or given out (released) by the piece of matter to bring about change.

Therefore, energy can be converted from one form into another and observations on physical changes and chemical reactions are summarized in the first law of thermodynamics which states that '**energy can be changed from one form into another, but it can neither be created nor destroyed'.**

Thermochemistry is closely connected to **thermodynamics**, a field in science that explains how energy is transferred and transformed. By studying thermochemistry, learners develop the ability to determine whether a reaction will release energy (exothermic) or absorb energy (endothermic), and whether it can occur spontaneously under given conditions.

Types of Reactions

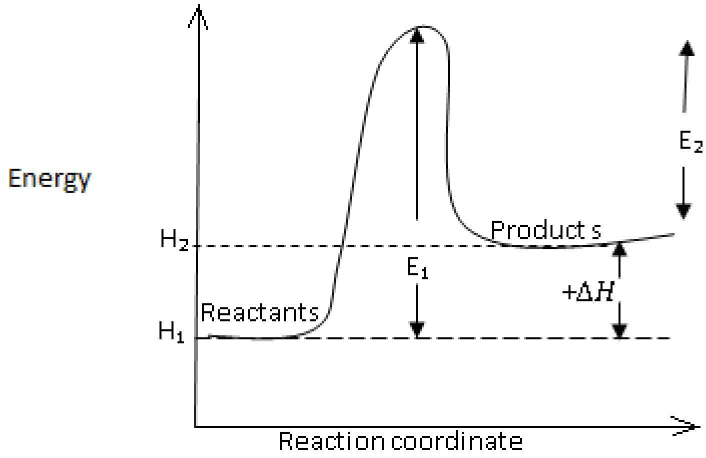
Heat energy leaves a system if the system is at a higher temperature than its surroundings (energy flows out us when we touch a cold object and it flows into us when we step into a hot bath).
Chemical reactions are accompanied by some energy changes which involves evolution or absorption of heat. The heat change involved in also referred to as enthalpy change; ΔH. It can be positive or negative.

$$\Delta H\ reaction\ =\ H_{products}\ -\ H_{reactants}$$

The two types of reactions depending on the heat changes involved include;
(i) **Endothermic reaction.** This is a chemical reaction in which heat is supplied to the system (i.e. the system absorbs heat energy). Thus, the surroundings lose heat and temperature falls. The heat change is denoted as $+\Delta H$.

e.g. $H_2O_{(l)} \longrightarrow H_2O_{(g)}$ $\Delta H = +44\ kJmol^{-1}$

Energy profile (level) diagram



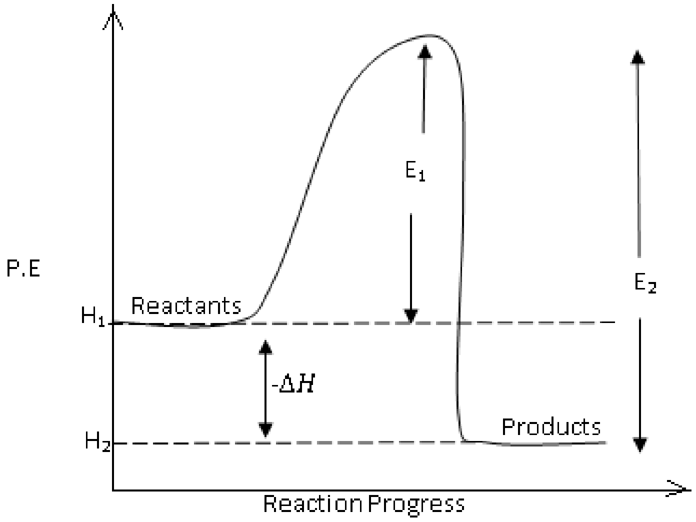
H_2 = Heat energy of products H_1 = Heat energy of reactants
E_2 = activation energy of products E_1 = activation energy of reactants

$$+\Delta H = E_2 - E_1 = H_2 - H_1$$

(ii) **Exothermic reaction.** This is a chemical reaction in which heat is released to the surroundings. Thus, the system loses heat and temperature of the surroundings increases. The heat change is denoted by a $-\Delta H$.

e.g. $CH_{4\,(g)} + 2O_{2\,(g)} \longrightarrow CO_{2\,(g)} + 2H_2O_{(l)}$ $\Delta H = -891\ kJmol^{-1}$

Energy level diagram



H_2 = Heat energy of products H_1 = Heat energy of reactants
E_2 = Activation energy of products E_1 = Activation energy of reactants

$$-\Delta H = E_2 - E_1 = H_2 - H_1$$

Some Real-life Scenarios for Learning Thermochemistry include:
(a)(i) **Cooking with Charcoal or Firewood.**
In many Ugandan homes, charcoal or firewood is used for cooking. When charcoal burns, a chemical reaction occurs between carbon and oxygen, releasing heat energy. Learners can analyse why this reaction is spontaneous once ignited and how the released heat makes cooking possible.
(ii) **Use of Biogas in Homes and Schools.**
Some schools and farms in Uganda use biogas produced from animal waste. The combustion of methane in biogas releases energy used for cooking. Learners can evaluate the thermodynamic principles behind this process and determine why methane combustion is a feasible and efficient energy source.
All combustion reactions are exothermic.
(b) **Cooling Drinks with Ice.**
When ice is placed in a drink, it melts by absorbing heat from the surroundings. Learners can analyse how energy transfer occurs and explain why the process causes the drink to become cold.
Hence melting is endothermic while freezing is exothermic.
(c) **Industrial Production Processes.**
In industries such as cement manufacturing, chemical reactions require large amounts of heat. Learners can examine whether such reactions are feasible and what energy conditions must be met for them to occur.
Other brief scenarios include:
- Reduction of a metal oxide by aluminium is very exothermic (thermit reaction).
- Pouring of water into an acid (concentrated) is a violet exothermic process.
- The process of photosynthesis is endothermic.
- Vapourisation is endothermic.

In conclusion, through these contexts, the competency to **interpret energy changes, predict whether reactions can occur spontaneously, and evaluate the practicality of chemical processes in real-life applications** is developed. This approach helps to connect chemistry concepts to societal needs such as energy efficiency, environmental conservation, and sustainable development.
Ultimately, studying thermochemistry equips one with scientific reasoning skills that enables **analysing energy-related processes in everyday life, agriculture, and industry**, making chemistry meaningful and relevant to Uganda's development.

STANDARD ENTHALPY CHANGES (Standard conditions)

Enthalpy changes are thus stated under standard conditions denoted by ΔH^{θ}.
The standard conditions include; pressure of 1 atmosphere (or 760mmHg or 101325 Pa or 101325 Nm^{-2}); temperature of 298K; 1 molar concentration for solutions (1M).
Standard enthalpy change of a reaction, ΔH^{θ}_{R}, is the enthalpy (heat) change that occurs (heat energy absorbed or evolved) in a reaction between the number of moles (in mole quantities) of reactants shown in the equation at standard conditions.
Factors that affect enthalpy changes include:
- the temperature of the reaction.
- the physical state of reactants and products. e.g.

$$H_{2(g)} + \tfrac{1}{2}O_{2(g)} \longrightarrow H_2O_{(g)\,(steam)} \qquad \Delta H = -242\ kJmol^{-1}$$
$$H_{2(g)} + \tfrac{1}{2}O_{2(g)} \longrightarrow H_2O_{(l)} \qquad \Delta H = -286\ kJmol^{-1}$$

- allotropic modification. e.g. $C_{(s)(graphite)} + O_{2(g)} \longrightarrow CO_{2(g)}$ $\Delta H = -393.4\ kJmol^{-1}$
  $C_{(s)(diamond)} + O_{2(g)} \longrightarrow CO_{2(g)}$ $\Delta H = -395.4\ kJmol^{-1}$
- the pressures of gases.
- the concentration of solutions.

Assumptions in Calorimetry (determination of heat changes) include:

- there is minimal / negligible or no heat loss (transfer of energy) to the surroundings.
- the heat capacity of plastic container or insulated calorimeter is negligible. i.e. no heat is gained by the plastic container or insulated calorimeter.
- the heat energy measured is attributed to that of a complete reaction.
- for dilute aqueous solutions, the density and specific heat capacity is equivalent to that of water.
- whatever chemical reaction is studied, no side reaction is taking place and no impurities are present.

Key Terms and Definitions of Enthalpy Changes Include:

(a) ***Standard enthalpy of formation (Heat/Enthalpy of formation)***. This is enthalpy change that occurs when one mole of a compound (substance) is formed from its constituent (component) elements in their standard states (natural physical states at standard conditions).

(b) ***Standard enthalpy of combustion (Heat/Enthalpy of combustion)***. This is heat energy liberated when one mole of a compound (substance) is completely burnt in oxygen gas (is burnt in excess oxygen) at standard conditions.

(c) ***Atomization energy (Enthalpy of atomization)***. This is the heat energy required (heat absorbed) when one mole of gaseous atoms is formed from its elements in their standard states.

N.B: *Enthalpy of sublimation (sublimation energy)* can be used for the change from the solid state directly to the gaseous state while *enthalpy of vaporization (vapourisation energy)* can be used for the change from the liquid state to the gaseous state.

(d) ***Ionisation energy (Enthalpy of ionization)***. This is the minimum amount of heat energy required to remove one mole of electrons from one mole of gaseous atoms to form one mole of singly charged gaseous cations. This definition in particular is for first ionisation energy.

(e) ***Electron affinity (Enthalpy of electron affinity)***. This is the enthalpy change that occurs when one mole of electrons is added to one mole of gaseous atoms to form one mole of singly charged gaseous anions. This definition in particular is for first electron affinity.

N.B: The 1st electron affinity for most elements especially nonmetals is an exothermic process (heat energy is released) because the electron is being added to a neutral gaseous atom with a strong nuclear attraction for the electron (i.e the electron is attracted by the positive charge on the nucleus of the atom). Hence energy is released.

The 2nd and successive electron affinities are endothermic processes (heat energy is absorbed) because an electron is added to a negatively charged ion which repel each other. Thus, energy is absorbed in order to overcome repulsion and effect the addition of electrons.

(f) ***Bond energy***. This is the enthalpy change that occurs when one mole of covalent bonds is formed or broken.

(g) ***Bond dissociation energy (Bond enthalpy of dissociation)***. This is the heat energy required to breakdown one mole of covalent bonds into its constituent (component) gaseous atoms.

(h) ***Bond energy of formation (Bond enthalpy of formation)***. This is the heat energy given out when one mole of covalent bonds is formed from its constituent gaseous atoms.

(i) ***Lattice energy***. This is heat energy given out when one mole of a solid ionic compound is formed from its constituent (component) gaseous ions. Its value is negative (exothermic). This is also known as the ***lattice energy of formation***.

Alternatively, lattice energy is the heat energy absorbed when one mole of a solid ionic compound is decomposed (broken down) into its constituent (component) gaseous ions. Its value is positive (endothermic). This is also known as the ***lattice energy of dissociation***.

(j) ***Enthalpy of solution (Heat of solution)***. This is the enthalpy change that occurs when one mole of a substance is dissolved in a specific amount of a solvent to form an infinitely dilute solution.

Alternatively, it is the heat change that occurs when one mole of a substance dissolves in water (a solvent) to form an infinitely dilute solution.

(k) ***Hydration energy (Enthalpy of hydration)***. This is the energy liberated when one mole of gaseous ions is completely surrounded by water molecules to form an infinitely dilute solution.

Alternatively, it is the heat energy evolved when one mole of gaseous ions dissolves in water to form an infinitely dilute solution.

(l) ***Enthalpy of neutralization.*** This is heat change that occurs when one mole of hydrogen ions from an acid and one mole of hydroxide ion from a base react to form one mole of water molecules.

(m) ***Enthalpy of displacement.*** This is the enthalpy change that occurs when one mole of less reactive elements is displaced from a compound containing their ions by more reactive elements in their standard states.

(n) ***Enthalpy of precipitation.*** This is the enthalpy change that occurs when one mole of an insoluble substance is formed from its ions in the aqueous state. e.g the enthalpy of precipitation of lead(II) sulphate using aqueous lead(II) nitrate and sodium sulphate is $-50.4\ kJmol^{-1}$.

(o) ***Enthalpy of hydrogenation.*** This is the enthalpy change that occurs when one mole of an unsaturated compound completely reacts with hydrogen gas to form one mole of a saturated compound.

The Laws of Thermochemistry

These laws are based on the principle of conservation of energy.

(a) **The 1st law** is also known as **Lavoiser and Laplace's law.** It states that, '*the quantity of enthalpy/heat used to decompose a given component into its elements is equal in magnitude but opposite in sign to the amount of heat given out when a compound is formed*.

e.g. $HCl_{(g)} \longrightarrow \frac{1}{2}H_{2(g)} + \frac{1}{2}Cl_{2(g)} \quad \Delta H = +92\ kJmol^{-1}$

$\frac{1}{2}H_{2(g)} + \frac{1}{2}Cl_{2(g)} \longrightarrow HCl_{(g)} \quad \Delta H = -92\ kJmol^{-1}$

This law is applied in all other equations. e.g. lattice energy of rubidium.

$RbCl_{(s)} \longrightarrow Rb^{+}_{(g)} + Cl^{-}_{(g)} \quad \Delta H = +675\ kJmol^{-1}$

$Rb^{+}_{(g)} + Cl^{-}_{(g)} \longrightarrow RbCl_{(s)} \quad \Delta H = -675\ kJmol^{-1}$

(b) **The 2nd law** is also known as **Hess' law of constant heat summation.** It states that '*the heat change in a chemical reaction remains the same irrespective of whether the reaction is brought about in one step or through several intermediate steps, but it depends on the initial states of the reactants and the final states of the products*.' The validity of Hess' law can be demonstrated experimentally by measuring the heat change when a reaction is brought about in two or more different ways.

e.g. The result of making a solution of NH_4Cl from 36.5g of HCl gas and 17g of NH_3 and water in two ways is summarized below;

Method I - Mix the two gases to form a white sublimate of ammonium chloride.

i.e. $NH_{3(g)} + HCl_{(g)} \longrightarrow NH_4Cl_{(s)} \quad H_1 = \Delta H = -176.1\ kJmol^{-1}$

- Dissolve the ammonium chloride in water to form an aqueous solution.

i.e. $NH_4Cl_{(s)} + (aq) \longrightarrow NH_4Cl_{(aq)} \quad H_2 = \Delta H = +16.3\ kJmol^{-1}$

Overall reaction $\Delta H = -159.8 kJmol^{-} = H_1 + H_2 = Hess'law$

Method II - Dissolve ammonia gas in water to form ammonia solution.

i.e. $NH_{3(g)} + (aq) \longrightarrow NH_{3(aq)} \quad H_3 = \Delta H = -35.2\ kJmol^{-1}$

- Dissolve hydrogen chloride gas in water to form hydrochloric acid solution.

i.e. $HCl_{(g)} + (aq) \longrightarrow HCl_{(aq)} \quad H_4 = -72.4\ kJmol^{-1}$

- Mix the two solutions to form aqueous ammonium chloride.

$NH_{3(aq)} + HCl_{(aq)} \longrightarrow NH_4Cl_{(aq)} \quad H_5 = -52.3kJmol_{-1}$

Overall reaction $\Delta H = -159.9\ kJmol^{-1} = H_3 + H_4 + H_5 = Hess'law$

$\therefore$ according to Hess' law, $H_1 + H_2 = H_3 + H_4 + H_5$

From the above example, a *Born – Haber cycle* or *Energy Level diagram* can be constructed.

A Born – Haber cycle is a technique for applying Hess' law to the standard enthalpy changes which occur when an ionic compound is formed.

$$\begin{tikzcd}
NH_{3(g)} \arrow[d, "H_3\ (aq)"'] & + & HCl_{(g)} \arrow[d, "H_4\ (aq)"'] \arrow[r] & NH_4Cl_{(s)} + (aq) \arrow[r] & NH_4Cl_{(aq)} \\
NH_{3(aq)} & + & HCl_{(aq)} \arrow[urr, "H_5"] & &
\end{tikzcd}$$

From Hess' law, $H_1 + H_2 = H_3 + H_4 + H_5$

Points to note from Hess' law

(i) All heat changes should have the correct sign.
(ii) When a reaction is reversed, the sign should also be reversed.
(iii) Multiplying or dividing a number also multiplies or divides the ΔH by the same number.
(iv) States of products and reactants should be specified because different states have different heat changes.

Standard Enthalpy of Formation (ΔH_f^θ)

This is enthalpy change that occurs when one mole of a chemical compound (substance) is formed from its constituent (component) elements in their standard states.

e.g. $Na_{(s)} + {}^{1}/_{2}\, Cl_{2(g)} \longrightarrow NaCl_{(s)}$ $\quad \Delta H_f^\theta = -411\ kJmol^{-1}$

$2C_{(s)} + H_{2(g)} \longrightarrow C_2H_{2(g)}$ $\quad \Delta H_f^\theta = +227\ kJmol^{-1}$

1 mole of NaCl is formed from solid sodium and gaseous chlorine molecules.

Note:

- All elements in their standard states, *have a value of zero as their enthalpy of formation*.
- Enthalpies of formation of compounds can be *used to calculate the heat change of any chemical reaction*.
- Enthalpies of formation are determined indirectly (calculated) using the Born-Haber cycle or energy level diagram but they are not determined directly (experimentally).
- **The stability of a chemical substance** is explained using the:
  - *enthalpy of formation of the substance*. A negative value shows that a compound is stable while a positive value shows that a compound is not stable.

  The more negative the enthalpy of formation, the more stable the compound is and the reverse is true. Similarly, the more positive the enthalpy of formation, the more unstable the compound is and the reverse is true.
  - *lattice energy* (for ionic compounds). The larger the value of the lattice energy value, the more stable the compound. The reverse is true.
  - *bond energy* (for covalent compounds). The larger the value of the bond energy value, the more stable the compound. The reverse is true.

Sample items:

1. An engineer from National Water and Sewerage Cooperation knows that the energy changes in sewage treatment reactions are important for determining whether the process will require external heating or release energy that can be harnessed. The following data for heats of formation are available:

Substance	$\Delta H^{f\circ}$ (kcal/mol)
$CO_2(g)$	–94.0
$CO(g)$	–26.4
$H_2O(l)$	–68.3

As a chemistry student, use the above data to help the engineer to **find out the heat of reaction at standard temperature and pressure** and explain whether the reaction is **exothermic or endothermic**.

Competency Focus: Learners apply knowledge of enthalpy and thermochemical equations to analyze industrial energy processes.

Possible response: $CO_{(g)} + H_2O_{(l)} \longrightarrow CO_{2(g)} + H_{2(g)}$

ΔH_f^o $\quad -26.4 \quad -68.3 \quad -94.0 \quad 0$

$\Delta H_{rxn} = \sum standard\ enthalpies\ of\ formation\ of\ products - \sum standard\ enthalpies\ of\ formation\ of\ reactants$

$= -94 - (-26.4 + -68.3) \quad = +0.7$ kCal.

The rection is endothermic because the enthalpy change is positive.

2. A group of environmental science students is investigating alternative fuels for household use. They are comparing how much energy different substances release when burned completely in oxygen. One of the fuels they are testing is **benzene (C_6H_6)**, which burns according to the equation:

$C_6H_6(l) + 7½\ O_2(g) \rightarrow 6CO_2(g) + 3H_2O(l)$

To determine how much energy is released during this combustion, the students gather the following data on standard heats of formation:

Substance	**$\Delta H_f°$ (kJ/mol)**
$O_2(g)$	0
$H_2O(l)$	–285
$CO_2(g)$	–393
$C_6H_6(l)$	+49

Task: As a chemistry learner, use the data provided to **find out the heat of combustion of benzene**, and discuss how the value obtained would influence its suitability as a fuel compared to other hydrocarbons.

Competency Focus: Learners apply the law of conservation of energy and knowledge of enthalpy changes to evaluate energy sources for sustainability and efficiency.

Possible response:

$$\begin{array}{ccccccc} C_6H_{6(l)} & + & 7^{1}/_{2}\ O_{2(g)} & \longrightarrow & 6CO_{2(g)} & + & 3H_2O_{(l)} \\ +49 & & 0 & & 6 \text{ x } (-393) & & 3 \text{ x } (-285) \end{array}$$

$\Delta H_{rxn} = \sum standard\ enthalpies\ of\ formation\ of\ products\ - \sum standard\ enthalpies\ of\ formation\ of\ reactants$

$$\Delta H^{o}_{rxn} = [(6 \text{ x} - 393) + (3 \text{ x} - 285)] - [+49 + 0]$$

$$= -3213 - 49 = -3262 \text{ kJmol}^{-1}$$

Benzene would be suitable because its combustion reaction produces a lot of heat energy.

Standard Enthalpy of Combustion (ΔH^{o}_{c})

This is heat energy liberated when one mole of a substance is completely burnt in oxygen gas at standard conditions. *Alternatively*, this is heat energy liberated when one mole of a substance is burnt in excess oxygen gas at standard conditions.

In these reactions, heat is always produced (given out or evolved) and the ΔH^{o}_{c} is exothermic.

Determination of enthalpy of combustion

(a) Experiment to determine the heat of combustion of any substance by measurement

It involves the determination of enthalpy of combustion directly using:

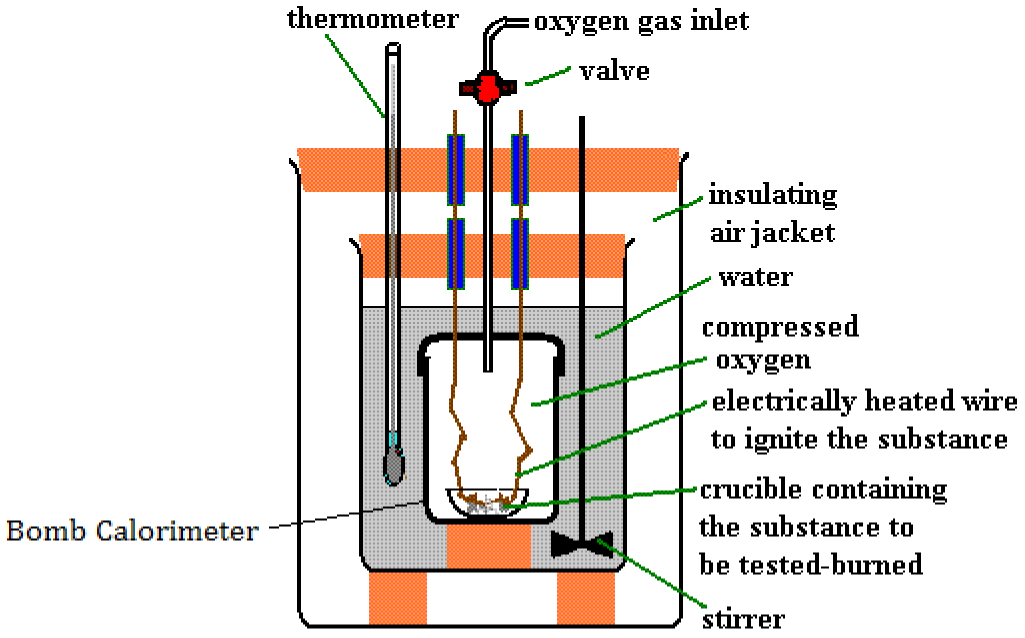
(i) A bomb calorimeter.

Procedure:

Using a weighing scale, a known mass of substance under test is weighed and placed in a platinum crucible at the centre of the bomb. The bomb calorimeter is then filled using a complicated valve system, with excess oxygen gas at a pressure of about 25 atmospheres to ensure rapid and complete combustion. The calorimeter is then immersed in a known volume of water containing a thermometer and stirrer. The initial temperature T_1 of the water is then taken using a thermometer. Combustion of the substance is started by passing an electric current through a thin iron or platinum wire in contact with the substance.

Heat energy will be evolved as the substance burns which raises the temperature of the water. The maximum temperature of the water T_2 is then recorded using a thermometer.

Diagram



Heat Change = mc Δt

Let w = mass of substance
t = rise in temperature $(T_2 - T_1)$
p = heat produced by the burning wire
W = mass of calorimeter
m = mass of water used
c = specific heat capacity of water
M = molar mass of substance
h = heat capacity of calorimeter

Heat evolved = heat gained by water + heat gained by the calorimeter
(produced by burning wire & by burning substance alone)

$$= (m\,c\,t) + (W\,h\,t) = (m\,c + W\,h)\,t \text{ Joules}$$

Heat produced by burning the substance alone = (mc + Wh) t - p Joules

Thus wg of substance yields (mc + Wh)t - p Joules

Mg of substance yields $\frac{[(mc + Wh)t - p]}{w}$ x M joules

Since 1 mole of a substance weighs Mg

∴ 1 mole of substance yields $\frac{[(mc + Wh)t - p]}{w}$ x M joules

$$\Delta H_c^o = -\frac{[(mc + Wh)t - p]}{W \times 1000} \times M \text{ kilo Joules per mole} \quad (\text{kJmol}^{-1})$$

The negative sign is because the reaction is exothermic since there is a temperature rise hence heat energy was evolved.

Note that: A similar experiment is used when during the treatment of results, the heat energy from the burning wire is assumed to be negligible as shown below.

Heat Change = mc Δt

Let w = mass of substance
t = rise in temperature $(T_2 - T_1)$
h = heat capacity of calorimeter
W = mass of calorimeter
m = mass of water used
c = specific heat capacity of water
M = molar mass of substance

Heat evolved = heat gained by water + heat gained by the calorimeter (produced by the burning substance)

$$= (m\,c\,t) + (W\,h\,t) = (m\,c + W\,h)t \text{ Joules}$$

Heat produced by burning the substance alone = (mc + Wh)t Joules

Thus wg of substance yields (mc + Wh)t Joules

Mg of substance yields $\frac{[(mc + Wh)t]}{w}$ x M joules

Since 1 mole of a substance weighs Mg

∴ 1 mole of substance yields $\frac{[(mc + Wh)t]}{W}$ x M joules

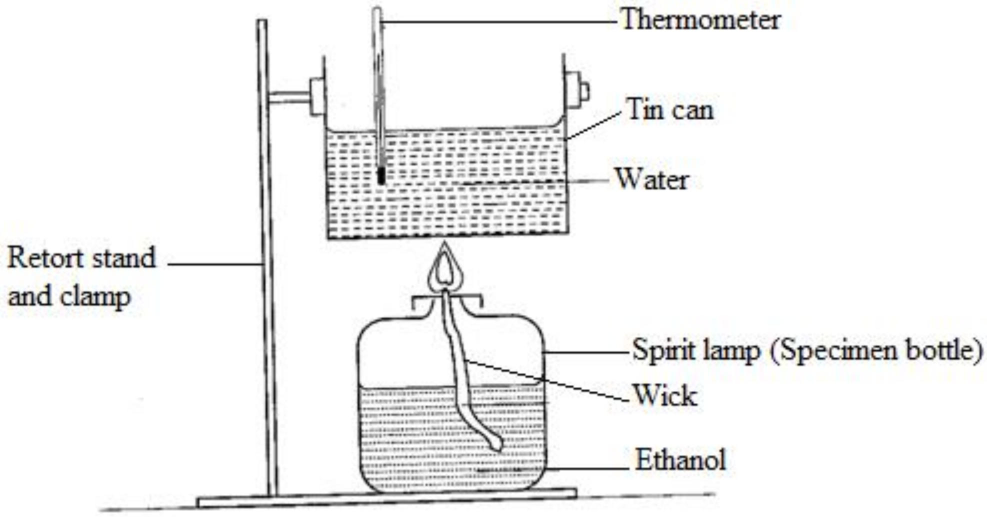
$\Delta H^{o}_{c} = -\frac{[(mc + Wh)t]}{W \times 1000}$ x M kilo Joules per mole (kJmol^{-1})

(ii) A spirit lamp. This is used only for liquids.

e.g. Experiment to determine the heat of combustion of an alcohol such as ethanol:

An ethanol spirit lamp with a wick inserted in it is commonly used. Other requirements include; a thermometer, retort stand and clamp, tin can, water, and a lighter.

The apparatus is set up as shown below.



Procedure:

- A thin-walled tin can is filled with a known volume, Vcm^3 of water using a measuring cylinder and the initial temperature of the water is recorded, T_1 ^{0}C using a thermometer.
- Ethanol is added to a spirit lamp using a measuring cylinder and a wick is fitted in a cork through the top.
- Using a weighing scale, the mass M_1 of the spirit lamp with its contents is taken and recorded.
- The wick is lit in order for the flame to heat the water in the tin can.
- The water is stirred carefully with the thermometer and after a convenient temperature rise, the flame is blown off and the temperature reached is recorded, T_2 ^{0}C.
- Using a weighing scale, the mass M_2 of the spirit lamp and its contents is taken after cooling and recorded.

Points to take note in this experiment include:

i) A tin can is used because it is a good conductor of heat and transmits most of the heat directly into the water.
ii) A wire gauze is not used to distribute the flame because it would absorb some of the heat.
iii) Usually, a shield is used because it reduces heat loss to the surrounding.
iv) It is assumed that all the heat energy produced by the burning alcohol is absorbed by the water.

Errors that occur during the experiment include:

(i) Usually heat energy is lost to the surroundings and some to the tin (copper) can holding the water.
Thus, the values of the enthalpy of combustion calculated through treatment of results obtained in the experiment are usually lower than the actual values.
(ii) Alcohols like ethanol gradually absorb water from the atmosphere. Hence errors in mass loss of contents in the flask after burning occurs.

(iii) There is formation of soot on the surface of the tin can. This implies that ethanol does not burn completely and the maximum enthalpy change for complete combustion is not obtained.
(iv) Human error usually occurs during reading of the temperature on the thermometer scale.
(v) Loss of ethanol through evaporation.

Treatment of results:

- Mass of ethanol that burnt = $(M_1 - M_2)$ g
- Density of water = 1 gcm^{-3} hence mass of water = density x volume = 1 x V = V g
- Temperature change (rise) of water, $\theta = (T_2 - T_1)\ ^0C$
- Specific heat capacity of water, c = 4.2 $J/g/^0C$

Heat energy produced by the burning ethanol = heat energy gained by the water in the tin can.
Heat gained by water = mass of water x specific heat capacity of water x temperature change
= $V \times c \times \theta$ Joules

Then $(M_1 - M_2)$ g of ethanol burn to produce $(V \times 4.2 \times \theta)$ Joules of heat energy
M_r g of ethanol burn to produce $\left(\frac{V \times 4.2 \times \theta}{M_1 - M_2}\right) \times M_r$ Joules, where M_r is the molar mass of ethanol

Thus, the heat of combustion of ethanol = $\left(\frac{V \times 4.2 \times \theta}{M_1 - M_2}\right) \times M_r$ J/mol

$$= -\frac{(V \times 4.2 \times \theta)}{(M_1 - M_2) \times 1000} \times M_r \text{ kJ/mol}$$

The units of energy are in kJ/mol hence the joules are converted to kilo joules.
The negative (-) sign is because there was a temperature rise of water in the tin can as heat energy was produced (liberated). Hence the reaction is exothermic.

Alternatively,

Moles of ethanol burnt = $\left(\frac{M_1 - M_2}{M_r}\right)$ moles
$\left(\frac{M_1 - M_2}{M_r}\right)$ moles of ethanol burn to evolve $(V \times c \times \theta)$ Joules
1 mole of ethanol burns to evolve $\left(\frac{V \times 4.2 \times \theta}{M_1 - M_2}\right) \times M_r$ Joules
Thus, the heat of combustion of ethanol = $-\frac{(V \times 4.2 \times \theta)}{(M_1 - M_2) \times 1000} \times M_r$ kJ/mol

N.B; ethanol is an example of a fuel. The following are qualities of a good fuel.
i) It must produce a lot of heat.
ii) It must be cheap and must burn easily.
iii) It should not burn with much smoke (pollutant gas).
iv) It should be easily transported with little or no fire risks.
v) It should be easily safely stored.

Item: During a science exhibition at a local secondary school, a group of students conducted an experiment to demonstrate how energy is released during the combustion of sugar (sucrose). They completely burned **48.36g of sucrose** in a calorimeter, heating **2.35 kg of water** from **24.5°C to 97.5°C**. The students aimed to calculate the **molar heat of combustion** of sucrose ($C_{12}H_{22}O_{11}$), using the known values: Density of water = 1 g/cm^3, Specific heat capacity of water = 4.2 J/g°C, C=12, H=1, O=16

Tasks: As a student in the group;
(i) Describe an experiment that can be used to determine the enthalpy of combustion of sugar (sucrose). (Diagram is **not** required) (06 scores)

(ii) Use the data from the experiment to **calculate the molar heat of combustion of sucrose**. Show your working clearly. (03 scores)

(iii) After comparing the experimental value with the theoretical value from textbooks, the students noticed that their result was lower. **Explain** why the molar heat of combustion determined in the experiment might be **lower than the theoretical value**. (02 scores)

(b) Determination of enthalpy of combustion indirectly.

This is achieved by applying the laws of thermochemistry in various calculations such use of the Born Haber cycle and Energy Level diagram.

Sample items involving ΔH_f^θ and ΔH_c^θ include:

1. A group of chemistry students at Mengo Senior School is investigating how much energy is involved when different fuels are formed or burned. One of the fuels they are studying is **ethyne (C_2H_2)**, a gas often used in welding because of its high flame temperature.

To understand how much energy is required to form ethyne from its elements, the students collect the following data from standard tables of thermochemical reactions:

Reaction	Equation	ΔH° (kJ mol⁻¹)
(I)	$C(s) + O_2(g) \rightarrow CO_2(g)$	-394
(II)	$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$	-286
(III)	$C_2H_2(g) + 2\frac{1}{2}O_2(g) \rightarrow 2CO_2(g) + H_2O(l)$	-1300

Task: As a chemistry learner, use the data above, help the students **calculate the enthalpy of reaction for the formation of ethyne** from its elements according to the equation: $2C(s) + H_2(g) \longrightarrow C_2H_2(g)$

Explain the steps you take and indicate whether the overall reaction is **endothermic or exothermic**.

Answer approach one:

Equation (I) shows both the ΔH_c^θ of carbon and ΔH_f^θ of carbondioxide.

Equation (II) shows both the $\mathbf{\Delta H_c^\theta}$ of hydrogen and $\mathbf{\Delta H_f^\theta}$ of water.

The standard enthalpy content of a substance is equal to the standard enthalpy of formation of the substance from its elements in their standard states.

$$C_2H_{2(g)} + 2^1/_2O_{2(g)} \longrightarrow 2CO_{2(g)} + H_2O_{(l)} \qquad \Delta H_3^\theta = -1300\ \text{kJmol}^{-1}$$

$$\Delta H^\theta{}_f(C_2H_2) + \Delta H^\theta{}_f(O_2) \longrightarrow \Delta H^\theta{}_f(CO_2) + \Delta H^\theta{}_f(H_2O)$$

$$\Delta H_f^\theta(C_2H_2) + 0 \longrightarrow 2(-394) + (-286) \qquad \Delta H_3^\theta = -1300\ \text{kJmol}^{-1}$$

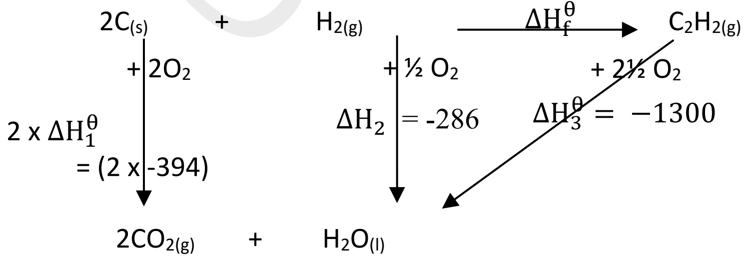
From, $\begin{pmatrix}\text{standard enthalpy}\\\text{change of reaction}\end{pmatrix} = \begin{pmatrix}\text{standard enthalpy}\\\text{change of products}\end{pmatrix} - \begin{pmatrix}\text{standard enthalpy}\\\text{change of reactants}\end{pmatrix}$

$$-1300 = [2(-394) + (-286)] - [\Delta H_f^\theta(C_2H_2)]$$

$$\Delta H_f^\theta(C_2H_2) = +226\ \text{kJmol}^{-1}$$

Answer approach two: **Application of Hess' Law**

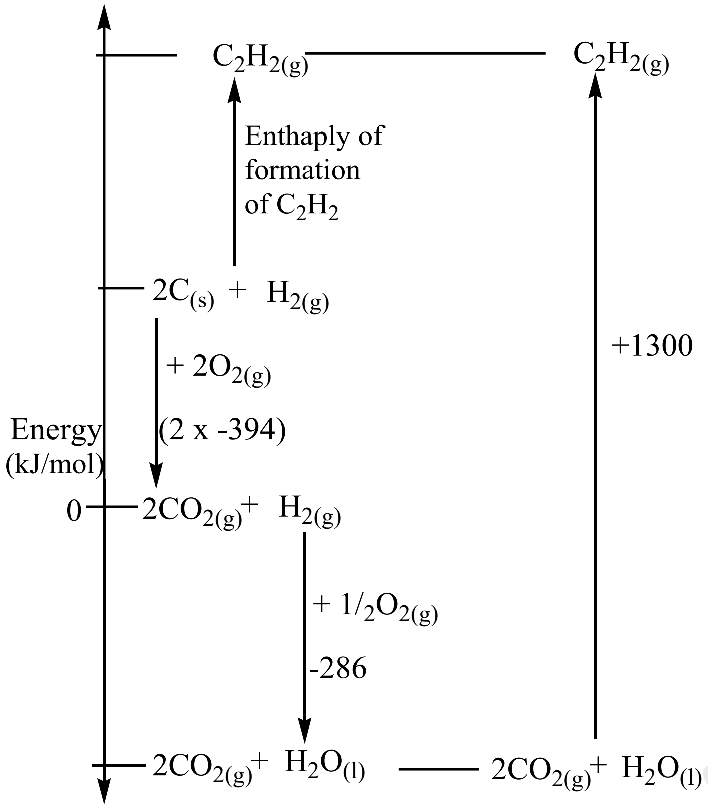
Construction of a born Haber cycle and apply Hess's law



$$\Delta H_f^\theta + -1300 = (2 \times -394) + (-286)$$

$$\Delta H_f^\theta = (2 \times -394) + (-286) + 1300 \qquad = +226\ \text{kJmol}^{-1}$$

Alternatively, construction of the energy level diagram and apply Hess's law



Enthalpy of formation, ΔH_f^θ of ethyne = (2 x -394) + (-286) + (+1300) = +226 kJmol^{-1}

Answer approach three: **Application of the first law of thermochemistry**

Multiply equation (I) by 2

$2C_{(s)} + 2O_{2(g)} \longrightarrow 2CO_{2(g)}$ $\Delta H_4^\theta = (2 \times -394) = -788$ kJmol^{-1} -------- (IV)

Add equation (II) to equation (IV)

$2C_{(s)} + 2O_{2(g)} \longrightarrow 2CO_{2(g)}$ $\Delta H_4^\theta = -788$ kJmol^{-1}

$+ \ H_{2(g)} + {}^1/_2O_{2(g)} \longrightarrow H_2O_{(l)}$ $\Delta H_2^\theta = -286$ kJmol^{-1}

$2C_{(s)} + H_{2(g)} + 2{}^1/_2O_{2(g)} \longrightarrow 2CO_{2(g)} + H_2O_{(l)}$ $\Delta H_5 = -1074$ kJmol^{-1} --- (V)

Then reverse equation (III) and add it to equation (V)

$2C_{(s)} + H_{2(g)} + 2{}^1/_2O_{2(g)} \longrightarrow 2CO_{2(g)} + H_2O_{(l)}$ $\Delta H_5^\theta = -1074$ kJmol^{-1}

$+ \ 2CO_{2(g)} + H_2O_{(l)} \longrightarrow C_2H_{2(g)} + 2{}^1/_2O_{2(g)}$ $\Delta H_3^\theta = +1300$ kJmol^{-1}

$2C_{(s)} + H_{2(g)} \longrightarrow C_2H_{2(g)}$ $\Delta H_6^\theta = +226 \text{ kJmol}^{-1}$

2. Hannah is a secondary school student who is passionate about renewable energy and fuel efficiency. During her holiday, she visited a local biofuel plant that uses organic waste to produce propan-1-ol as a sustainable fuel. While touring the facility, she found out that understanding the energy content of biofuels is key to evaluating their usefulness compared to traditional fuels like petrol and diesel.

The plant chemist challenged Hannah to use data from the lab to calculate the enthalpy of formation of propan-1-ol. She is given the following standard enthalpies of combustion (in kJ mol^{-1}): $C_3H_7OH_{(l)}$ = -2010; $C_{(s)}$ = -394; $H_{2(g)}$ = -286

Task: As a chemistry learner, help Hannah to;

(a) Use different methods to calculate the standard enthalpy of formation of propan-1-ol (C_3H_7OH).

(b) Explain why the calculations of enthalpy of formation are useful for the biofuel plant's operations and decision-making.

For a biofuel plant, calculating the enthalpy of formation is important because it:

- **Aids to indirectly determine energy changes in reactions:** By using Hess's Law, enthalpies of formation allow us to calculate the enthalpy change (ΔH) for any chemical reaction without having to measure it directly.

This is especially useful when direct measurement is impractical or dangerous.

e.g. Calculating the energy released when a fuel like propane is burned.

- **Compares stability of compounds:** A more negative ΔH_f^0 means the compound is more stable relative to its elements. This helps chemists and engineers select materials that are thermodynamically stable under given conditions.
- **Helps assess energy efficiency**: It enables factory chemists to determine how much energy is needed or released when producing biofuels from raw materials. This helps in choosing the most energy-efficient production methods.
- **Supports economic decisions**: By comparing enthalpy values of different potential biofuels, the plant can decide which fuel offers the best balance of production cost and energy output.
- **Enables environmental analysis**: Calculating the enthalpy of formation helps in estimating greenhouse gas emissions and sustainability, supporting decisions that align with environmental regulations.
- **Assists in process design and safety**: Knowing energy changes involved helps in designing reactors and controlling temperatures to avoid hazards.

(c) Suggest one real-life application where knowing the enthalpy of combustion of a substance is useful.

One real-life application where knowing the enthalpy of combustion is useful is in **designing and operating internal combustion engines** (e.g., in cars or generators). Engineers use this value to determine the energy output of different fuels, helping optimize engine performance and fuel consumption efficiency.

For example, bioethanol can be evaluated against petrol to determine whether switching fuels would be energy-efficient and environmentally beneficial.

3. Enthalpy of Combustion of Alkanes and Alcohols.

A team of chemistry students at Mengo Senior School is working on a **bioenergy research project** to compare the energy content of different fuels. They are also to identify the most efficient fuel that could be used for both domestic and school laboratory heating purposes. Their chemistry teacher encourages them to investigate how the amount of energy released when burning different types of fuels depends on the number of carbon atoms in each molecule.

They are particularly interested in understanding how the **enthalpy of combustion** varies with the number of carbon atoms in straight-chain **alkanes** and **primary alcohols**.

The learners collect data on the *standard enthalpies of combustion* (ΔH°_n) for a series of straight-chain alkanes and primary alcohols as shown in the table below.

Alkanes	***$\Delta H^\circ c$ (kJ mol^{-1})***	***Alcohols***	***$\Delta H^\circ c$ (kJ mol^{-1})***
CH_4	-890.0	CH_3OH	-666.6
C_2H_6	-1560.7	C_2H_5OH	-1333.3
C_3H_8	-2219.2	C_3H_7OH	-2017.7
C_4H_{10}	-2878.0	C_4H_9OH	-2670.0

The team plans to use this information to predict the efficiency of longer-chain fuels such as **hexane** and **heptan-1-ol**.

Task:

(a) As a chemistry learner, help the students to **plot a graph** on the same axes of the enthalpy of combustion ($\Delta H^\circ c$) of both **alkanes** and **alcohols** against the **number of carbon atoms**.

(b) Use the graph to help students **estimate** the enthalpy of combustion of **hydrogen**, **water**, **hexane**, and **heptan-1-ol**.

(c) The learners notice that the graph for alkanes seems to follow a straight-line pattern. Help the students to use the slope and intercept of this line to calculate the enthalpy of combustion of heptane (C_7H_{16}).

(d) As a learner of chemistry, explain the shape of the two graphs and what this implies about the relationship between the number of carbon atoms in a molecule and its energy content.

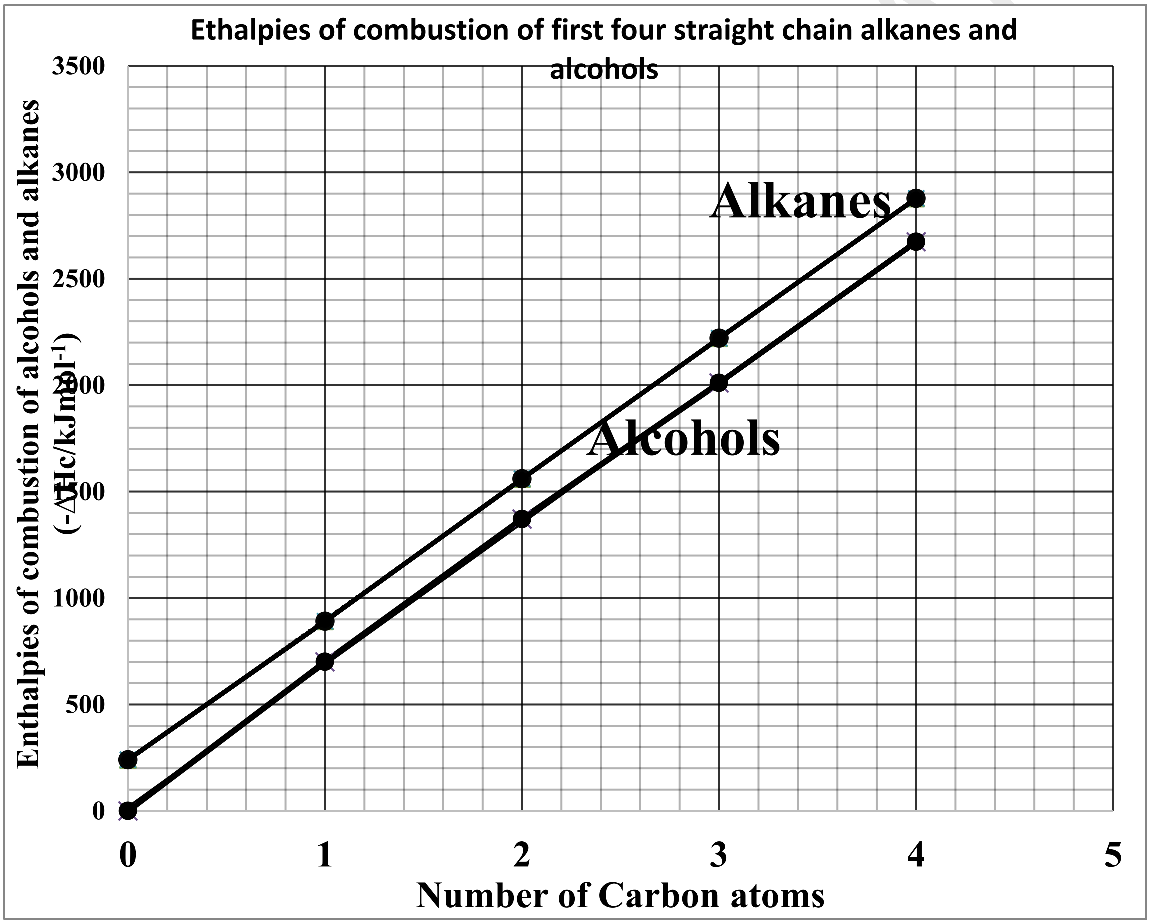
Theory:

(a) The enthalpies of combustion increase with the increase in the number of carbon atoms. This is because there is addition of a successive constant methylene group ($-CH_2$).

Thus, as the number of carbon and hydrogen atoms in the organic based compounds increase, the products formed proportionally increase after complete combustion. Hence more new bonds are formed and so the heat energy liberated increases. This also makes the two graphs shown above to become parallel.

(b) A plot of *enthalpy of combustion* against number of carbon atoms shows that the graph for:
(i) alcohols passes through the origin because when n=0, the general molecular formula for alcohols, $C_nH_{2n+1}OH$ or $C_nH_{2n+2}O$ gives the formula H_2O of which water does not burn.
(ii) alkanes has a vertical intercept because when n=0, the general molecular formula for alkanes, C_nH_{2n+2} gives the formula H_2 of which hydrogen burns with a blue flame to produce energy.
The enthalpy of combustion for any alkane can also be obtained (calculated) from the slope and the y-intercept using a relationship of the linear plots. i.e.
Plotting a graph of enthalpy of combustion, ΔH_c *(on y-axis)* against number of carbon atoms, n *(on x-axis)* has the graph in the form $y = mx + c$
then ΔH_c of alkane = slope x number of carbon atoms in the alkane + y-intercept

(c) The plot of *enthalpy of combustion* for alkanes and that for alcohols on the same axes shows that the graph for alkanes is above that of alcohols. This because one mole of an alkane uses more oxygen to completely burn as compared to one mole of an alcohol with the same number of carbon atoms.



(d) A plot of *enthalpy of formation* against number of carbon atoms for alkanes and that for alcohols on the same axes shows that the graph for:
(i) alcohols has a vertical intercept because when n=0, the general molecular formula for alcohols, $C_nH_{2n+1}OH$ or $C_nH_{2n+2}O$ gives the formula H_2O which water. In relation to enthalpy of formation; one mole of water is formed from one mole of hydrogen gas and half a mole of oxygen gas as the constituent elements in their standard states and hence there is an enthalpy change that takes place.

(ii) alkanes passes through the origin because when n=0, the general molecular formula for alkanes, C_nH_{2n+2} gives H_2 which hydrogen gas. In relation to enthalpy of formation; in order to form one mole of hydrogen gas, the starting chemical species is one mole of hydrogen gas and hence there is no enthalpy change that occurs.

(e) The plot of *enthalpy of formation* for alkanes and that for alcohols on the same axes shows that the graph for alcohols is above that of alkanes.
This because one mole of an alcohol is formed from more atoms (i.e. carbon, hydrogen and oxygen) as compared to one mole of an alkane (i.e. carbon and hydrogen) with the same number of carbon atoms.
Below are some enthalpies of formation values for straight chain alkanes and primary alcohols.

Alkanes	ΔH_f^θ *kJmol*$^{-1}$	*Alcohols*	ΔH_f^θ *kJmol*$^{-1}$
CH_4	-74.80	CH_3OH	-251.30
C_2H_6	-84.70	C_2H_5OH	-277.15
C_3H_8	-104.70	C_3H_7OH	-303.00
C_4H_{10}	-125.60	C_4H_9OH	-332.80

4. During a chemistry research project at Makerere University, a group of students decided to investigate the stability of different organic compounds that contain double bonds. They were particularly interested in understanding why **benzene** behaves differently from other unsaturated hydrocarbons such as **cyclohexene**.
In their study, they collected the following data on the **standard enthalpies of combustion** for some substances:

Substance	**Standard enthalpy of combustion, ΔHc (kJmol⁻¹)**
Hydrogen	-285
Benzene	-3280
Cyclohexene	-3725
Cyclohexane	-3920

Using this information, the students set out to analyze the **enthalpy of hydrogenation** of cyclohexene and benzene to understand the relative stability of these compounds.
Task: As a chemistry learner;
(a) Use the data provided to calculate the **enthalpy of hydrogenation** of **cyclohexene** and **benzene**.
(b) From your results in (a), **comment on the relative stability** of **benzene** compared to **alkenes (cyclohexatriene)**.

Determination of Lattice Energy

Lattice energy is the heat energy released or given off when one mole of solid ionic compounds (ionic crystal lattices) is formed from its constituent gaseous ions at standard conditions.

i.e. $M^+_{(g)} + X^-_{(g)} \longrightarrow MX_{(s)} \quad \Delta H = \text{negative}$

e.g. $Na^+_{(g)} + Cl^-_{(g)} \longrightarrow NaCl_{(s)} \quad \Delta H = -761\ kJmol^{-1}$

Alternatively, lattice energy can be defined as the amount of heat energy required to decompose (breakdown) one mole of solid ionic compounds (ionic crystal lattices) into its constituent gaseous ions at standard conditions.

i.e. $MX_{(s)} \longrightarrow M^+_{(g)} + X^-_{(g)} \quad \Delta H = \text{positive}$

e.g $NaCl_{(s)} \longrightarrow Na^+_{(g)} + Cl^-_{(g)} \quad \Delta H = +761\ kJmol^{-1}$

Factors that determine or affect lattice energy:

Lattice energy is given by the equation; $U_L = \frac{M\,NZ^+\,Z^+\,e^2\ (1 - {}^1/_n)}{r^+\,r^-}$

where U_L = Lattice energy M = Madelung's constant N = Avogadro's number
Z^+ = Charge on cation Z^- = Charge on anion e = electronic charge
r^+ = radius of cation r^- = radius of anion n = Born exponent

From the above expression, it shows that lattice energy basically depends on two factors:
(i) The *ionic charge*.
- *For cations,* the higher the ionic charge, the more ionic the compound.

Therefore, the stronger the attraction for the oppositely charged ions (the stronger the ionic bond) and hence the higher the lattice energy. The reverse is true.

e.g. NaCl ($Na^+ Cl^-$) -761 kJmol^{-1} $MgCl_2$ ($Mg^{2+} Cl^-$) -2489 kJmol^{-1}

The magnesium ions more strongly attract the chloride ions than the sodium ions hence magnesium chloride has greater lattice energy.

However, for cations with high polarising power; The higher the ionic charge, the higher the charge density and the higher the polarising power. Thus, the less ionic or the more covalent the compound.

Therefore, the weaker the attraction for the oppositely charged ions (the weaker the ionic bond) and hence the lower the lattice energy. The reverse is true.

- *For anions,* the higher the ionic charge, the more ionic the compound.

Therefore, the stronger the attraction for the oppositely charged ions (the stronger the ionic bond) and hence the higher the lattice energy. The reverse is true.

However, for anions with high polarizabity; The higher the ionic charge, the higher the polarizability (the higher the ability to be polarized). Thus, the less ionic or the more covalent the compound.

Thus, the weaker the attraction for the oppositely charged ions (the weaker the ionic bond) and hence the lower the lattice energy. The reverse is true.

(ii) The *ionic radius*. Lattice energy is inversely proportional to the radius of the ion.

For cations, the larger the ionic radius, the lower the charge density and the lower the polarising power. Thus, the less ionic or the more covalent the compound.

Thus, the weaker the attraction for the oppositely charged ions (the weaker the ionic bond) and hence the lower the lattice energy. The reverse is true.

For anions, the larger the ionic radius, the higher the polarizability (the higher the ability to be polarized). Thus, the more ionic or the less covalent the compound.

Thus, the stronger the attraction for the oppositely charged ion (the stronger the ionic bond) and hence the higher the lattice energy. The reverse is true.

e.g.	NaF	-915 kJmol^{-1}	NaCl	-776 kJmol^{-1}
	NaBr	-742 kJmol^{-1}	NaI	-699 kJmol^{-1}

In the above case, the charges on both the cations and the anions are of the same magnitude (singly positively and negatively charged) but the radius of the anions increases from fluoride to the iodide ion.

The increase in anionic radii progressively down the group leads to increased polarization by the cation (polarizability) decreasing the ionic nature of the salt.

Hence there is greatest of oppositely charged ions in sodium fluoride and least in sodium iodide.

Determination of lattice energy of Ionic Compounds using the Born Haber Energy Cycle and the Energy Level Diagram:

The Born-Haber cycle is employed and Energy Level Diagram in the determination at standard conditions. Hess' Law is applied in both cases.

e.g. Formation of sodium chloride crystal from sodium metal and chlorine gas.

$Na_{(s)} + {}^{1}/{}_{2}Cl_{2(g)} \longrightarrow NaCl_{(s)}$ ΔH_f^θ, Standard enthalpy of formation of sodium chloride

The reaction occurs by means of five steps each involving different energy terms.

(i) Conversion of solid to gaseous sodium atom. i.e. atomization of sodium.

$Na_{(s)} \longrightarrow Na_{(g)}$ ΔH_s^θ (or ΔH_a^θ) = +108 kJmol^{-1}

Sublimation or atomization energy of sodium.

(ii) Ionisation of gaseous sodium atom to give sodium ion (removal of an electron).

$Na_{(g)} \longrightarrow Na^+_{(g)} + e^-$ $\Delta H_{I.E}^\theta = +500\ kJmol^{-1}$

Ionization energy of sodium i.e 1st ionization energy

(iii) Dissociation of chlorine molecules into atoms.

$^1/_2Cl_{2\ (g)} \longrightarrow Cl_{(g)}$ $\quad {}^1/_2\ \Delta H^{\theta}_{BDE} = {}^1/_2$ Bond Dissociation Energy

i.e. Bond dissociation energy (BDE) of chlorine molecules = +242.2 kJmol^{-1}.

(atomisation energy of chlorine = $^1/_2$BDE of chlorine) = +121.1 kJmol^{-1}.

(iv) Ionisation of chlorine atoms to chloride ion (addition of an electron). This is called electron affinity.

$Cl_{(g)} + e^- \longrightarrow Cl^-_{\ (g)}$ $\quad \Delta H^{\theta}_{E_A} = -349$ kJmol^{-1} = Electron affinity of chlorine

(v) Reaction between $Na^+_{(g)}$ ion and $Cl^-_{(g)}$ ion to give $NaCl_{(s)}$.

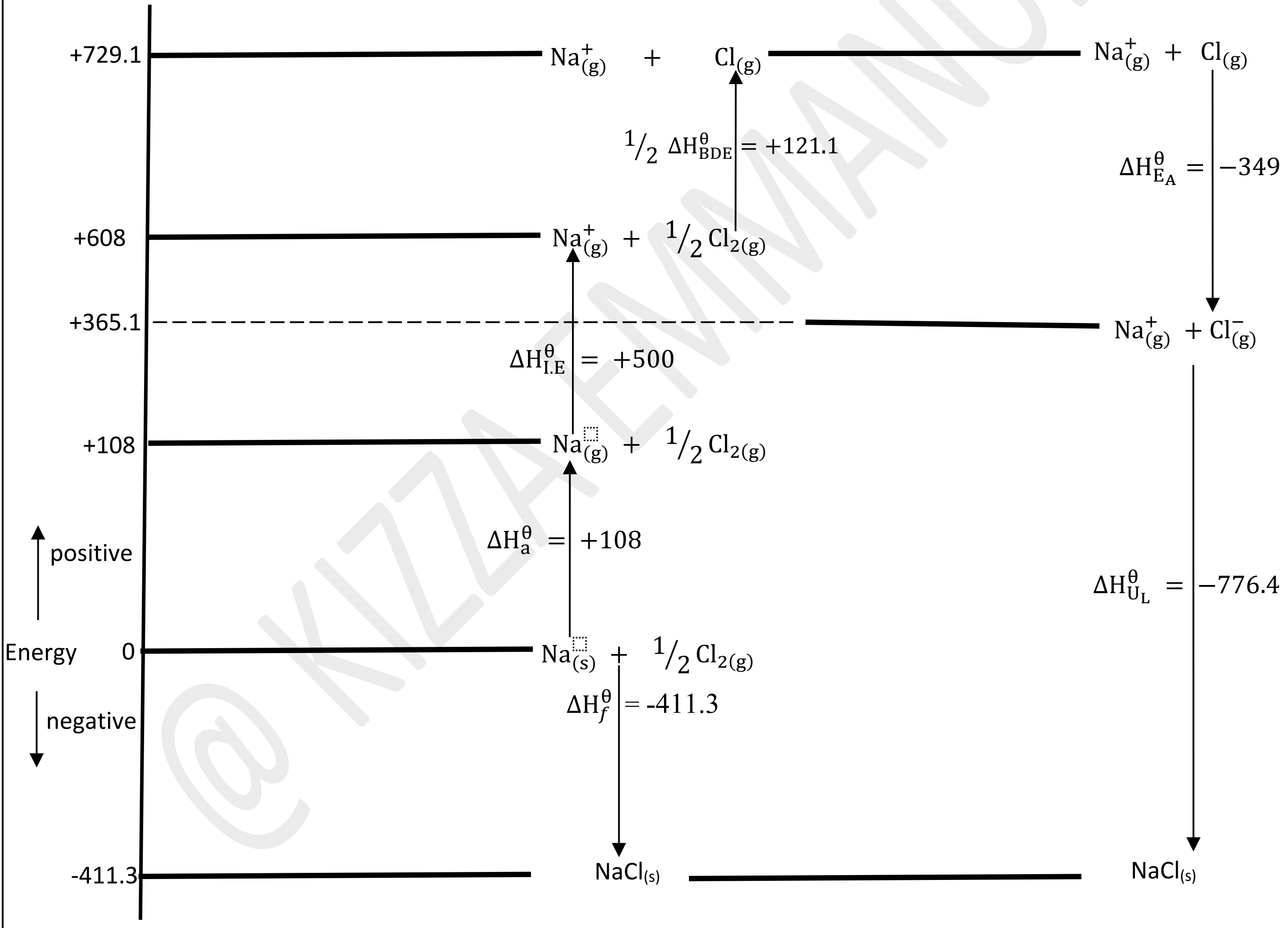
$Na^+_{(g)} + Cl^-_{(g)} \longrightarrow NaCl_{(s)}$ $\quad U_L(\Delta H^{\theta}_{U_L}) = -776.4\text{kJmol}^{-1}$

Lattice energy of sodium chloride.

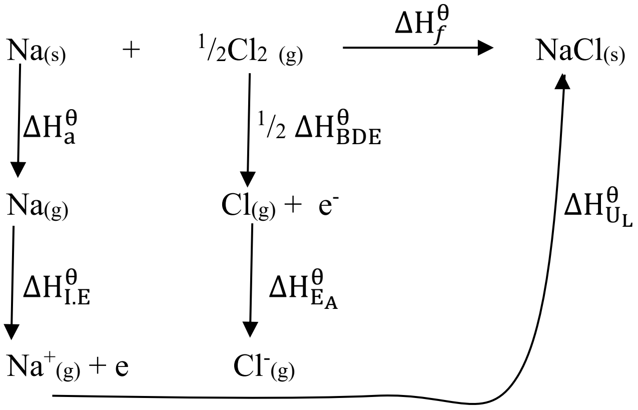
Overall energy change thus, $\Delta H^{\theta}_f = \Delta H^{\theta}_s + \Delta H^{\theta}_{I.E} + {}^1/_2\,\Delta H^{\theta}_{BDE} + \Delta H^{\theta}_{E_A} + \Delta H^{\theta}_{U_L}$

$= 108 + 500 + 121.1 - 349 - 776.4 = -396.3$ kJmol^{-1}

Energy Level Diagram is:



The Born Haber cycle is:



Note that; for salts that contain atoms with more than one electron to be removed in their outermost shell, one electron is removed at a time. e.g. For magnesium in magnesium chloride.

$Mg_{(g)} \longrightarrow Mg^+_{(aq)} + e^-$ $\Delta H^\theta_{I.E_1}$ = 1st ionization energy

$Mg^+_{(g)} \longrightarrow Mg^{2+}_{(aq)} + e^-$ $\Delta H^\theta_{I.E_2}$ = 2nd ionization energy

Similarly, salts that contain atoms with more than one electron to be added in their outermost shell, one electron is added at a time. e.g For oxygen in sodium oxide.

$O_{(g)} + e^- \longrightarrow O^-_{(g)}$ $\Delta H^\theta_{E_{A_1}}$ = 1st electron affinity

$O^-_{(g)} + e^- \longrightarrow O^{2-}_{(g)}$ $\Delta H^\theta_{E_{A_2}}$ = 2nd electron affinity

Sample items

Item 1: Magnesium Oxide Formation.

A team of materials scientists at a university research laboratory is investigating the formation of magnesium oxide (MgO), a compound widely used as a refractory material in furnaces due to its high thermal stability.

To understand why MgO is so stable, the team decides to calculate its enthalpy of formation from first principles using the Born–Haber cycle. They obtain the following thermochemical data from reference tables:

Process	**Equation**	**Enthalpy Change (kJ mol⁻¹)**
Sublimation	$Mg(s) \rightarrow Mg(g)$	+152.7
1st Ionization energy	$Mg(g) \rightarrow Mg^+(g) + e^-$	+738
2nd Ionization energy	$Mg^+(g) \rightarrow Mg^{2+}(g) + e^-$	+1451
Bond dissociation energy	$O_2(g) \rightarrow 2O(g)$	+496
1st Electron affinity	$O(g) + e^- \rightarrow O^-(g)$	–121.7
2nd Electron affinity	$O^-(g) + e^- \rightarrow O^{2-}(g)$	+866
Lattice energy (Option 1)	$Mg^+(g) + O^-(g) \rightarrow MgO(s)$	–1246
Lattice energy (Option 2)	$Mg^{2+}(g) + O^{2-}(g) \rightarrow MgO(s)$	–3933

Task: As a chemistry student with knowledge about thermochemistry, help the chemists to:

(a) Use the data to draw the **Born–Haber cycle** and **Energy Level diagram** for the formation of magnesium oxide.

(b) Find out the **enthalpy of formation** (ΔHf°) of MgO(s) using each of the two lattice energies provided.

(c) Use your results, **state and explain** which lattice energy leads to the formation of the more stable oxide.

Item 2: Electron Affinity of Chlorine

A chemical engineer working in a fertilizer manufacturing plant needs to understand the energy changes involved when potassium reacts with chlorine to form potassium chloride (KCl).

To estimate the electron affinity of chlorine, the engineer decides to construct a Born–Haber cycle using the following experimental data:

Process	**Equation**	**Enthalpy Change (kJ mol⁻¹)**
Sublimation of K	$K(s) \rightarrow K(g)$	+90
Ionization energy of K	$K(g) \rightarrow K^+(g) + e^-$	+420
Bond dissociation energy of Cl_2	$Cl_2(g) \rightarrow 2Cl(g)$	+244
Lattice energy of KCl	$K^+(g) + Cl^-(g) \rightarrow KCl(s)$	–706
Enthalpy of formation of KCl	$K(s) + \frac{1}{2}Cl_2(g) \rightarrow KCl(s)$	–436

Task: As a chemistry student with knowledge about thermochemistry, help the engineer to:

(a) Construct a Born–Haber cycle for the formation of potassium chloride.

(b) Use the data provided to calculate the electron affinity (EA) of chlorine.

(c) Illustrate the processes involved on an energy level diagram, showing where the electron affinity step fits in.

Determination of the Bond Energy of Covalent Compounds using the Born Haber Cycle:

Standard Bond Enthalpy (Bond Energy) is the heat change that occurs when one mole of covalent bonds is either broken to give its constituent gaseous atoms or formed from its constituent gaseous atoms at standard conditions.

When bonds are broken, energy is required or absorbed. However, when bonds are formed, energy is given out. i.e. There is a certain amount/quantity of energy associated with every type of chemical bond and this energy is absorbed when the bond is broken but evolved when it is formed.

Determination of the bond energy in a diatomic molecule due to formation of bonds, involves measurement of the heat of formation of the molecule.

On the other hand, determination of the bond energy is a molecule due to the breakage of bonds, involves measurement of the heat of dissociation of the molecule.

Thus, **bond dissociation energy** is the heat energy required to breakdown one mole of covalent bonds into gaseous atoms.

The bond length determines the strength of a covalent bond and hence the bond energy. The longer the bond, the weaker the covalent bond and the lower the bond energy. The reverse is true.

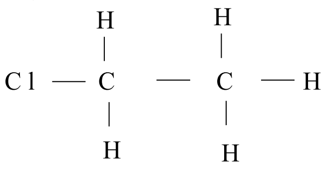
Factors affecting bond energy include:

- Atomic radius. The larger the radius of an atom in a covalent bond, the longer the bond and hence the weaker the covalent bond.
- Electron density. The higher the number of bonding electrons between two covalently bonded atoms, the shorter the bond and hence the stronger the covalent bond.
- Electronegativity. The higher the electronegativity, the more polar the bond such that the polar ends more strongly attract each other. Thus, the shorter the bond and hence the stronger the covalent bond.

Some bond energies of common bonds are given below;

Bond	**Bond Energy (kJmol⁻¹)**
H – H	436
C – C	348
C = C	615
C – H	414
O – H	463
C – O	360
C = O	743
Cl – Cl	242
H – Cl	431
C – Cl	329
O = O	496
C ≡ C	839

With poly atomic molecules, the bond energies can be obtained by thermo-chemical measurements as shown below: e.g. In chloroethane, CH_3CH_2Cl



$$\begin{array}{ccccccc} & & H & & H & & \\ & & | & & | & & \\ Cl & — & C & — & C & — & H \\ & & | & & | & & \\ & & H & & H & & \end{array}$$

There are 5(C – H) bonds, 1(C – C) and 1(C – Cl) bonds, when one mole of chloroethane is atomized (broken down to form gaseous atoms). The sum of these bond energies is the bond energy of the molecule.

These bond energies can also be used to calculate the heat or enthalpy of any reaction from the following formula;

Enthalpy of reaction = $\begin{pmatrix} sum\ of\ bond\ energies \\ of\ bonds\ broken \end{pmatrix} - \begin{pmatrix} sum\ of\ bond\ energies \\ of\ bonds\ formed \end{pmatrix}$

i.e. $\Delta H = (\sum bond\ energies\ of\ reactants) - (\sum bond\ energies\ of\ products)$

Sample Items:

1. At Namilyango College, a group of Senior Five learners in the chemistry club are participating in a school energy project. The club wants to understand how fuels release energy when used for cooking in homes around Kampala. Their teacher explains that many fuels are hydrocarbons, and understanding the **enthalpy of formation** of these compounds helps scientists predict how much energy they can store and release.

During one meeting, the learners decide to study **ethane (C_2H_6)**, a simple hydrocarbon found in natural gas. They want to estimate the **standard enthalpy of formation of ethane** from its elements in their standard states. The teacher guides them to use **bond enthalpies** and **enthalpies of atomization** to determine the energy changes involved when atoms combine to form the molecule.

The learners gather the following data from their chemistry reference books:

- Average bond enthalpies:
    - C – C bond = **348 kJ mol⁻¹**
    - C – H bond = **413 kJ mol⁻¹**
- Standard enthalpies of atomization:
    - C(s) → C(g) = **718 kJ mol⁻¹**
    - ½H_2(g) → H(g) = **218 kJ mol⁻¹** (per mole of hydrogen atoms formed)

The students recall that the **standard enthalpy of formation** refers to the enthalpy change when **one mole of a compound is formed from its elements in their standard states**.

Task: As a chemistry student, use the information provided to:

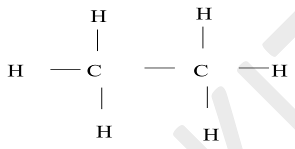
(a) Determine the total energy required to atomize the elements needed to form one mole of ethane.

(b) Calculate the total energy released when the bonds in ethane are formed.

(c) Hence, calculate the **standard enthalpy of formation of ethane (C_2H_6)**.

(d) State whether the formation of ethane is **exothermic or endothermic**, giving a reason for your answer.

Response:



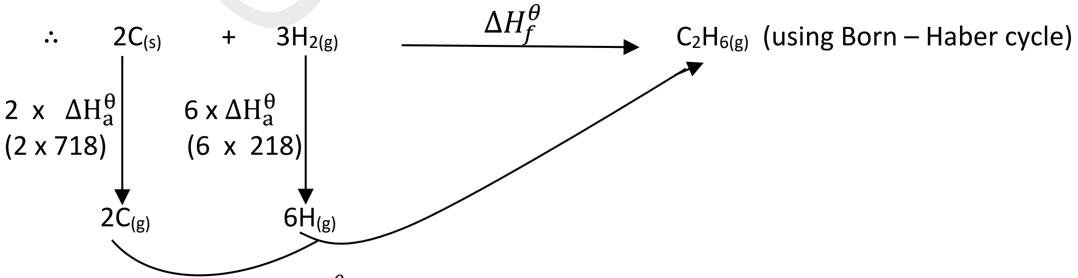
has 1 C – C bond = (1 x 348) = 348 kJmol^{-1}

6 C – H bonds = (6 x 413) = 2478 kJmol^{-1}

Total standard enthalpy of bonds = 348 + 2478 = 2826 kJmol^{-1}

Thus, this is the heat given out when atoms combine in their gaseous state to form ethane gas. It is exothermic.

i.e. $2C_{(g)} + 6H_{(g)} \longrightarrow C_2H_{6(g)} \quad \Delta H^{\theta}_{BE} = -2826\ \text{kJmol}^{-1}$



Using Hess's law; ΔH_f^{θ} = (2 x 718) + (6 x 218) + (-2826) = -82 kJmol^{-1}

Alternative method: Using heats of formation of products and reactants

$2C_{(g)}$ + $6H_{(g)}$ $\longrightarrow$ $C_2H_{6(g)}$ $\Delta H^{\theta}_{reaction}$ = -2826 kJmol^{-1}

ΔH^{θ}_{f} = (2 x 718) ΔH^{θ}_{f} = (6 x 218) ΔH^{θ}_{f} = ?

$$\Delta H^{\theta}_{reaction} = \sum \Delta H^{\theta}_{f} \text{ products} - \sum \Delta H^{\theta}_{f} \text{ reactants}$$

-2826 = ΔH^{θ}_{f} of C_2H_6 - $[(2 \text{ x } 718) + (6 \text{ x } 218)]$

ΔH^{θ}_{f} of C_2H_6 = - 82kJmol^{-1}

2. At **Rubaga Girls' Secondary School** in **Kampala**, the Senior Five chemistry class is carrying out a project on how industrial chemists determine the energy changes involved in chemical reactions. Their teacher explains that understanding reaction energies helps industries design safer and more efficient chemical processes.

During the lesson, the students investigate the reaction between **propanone (acetone)** and **hydrogen cyanide** to form a cyanohydrin compound, which is useful as an intermediate in organic synthesis. The reaction occurring in the gaseous state is represented as:

$$(CH_3)_2C = O(g) + HCN(g) \rightarrow (CH_3)_2C(OH)CN(g)$$

To estimate the **standard enthalpy change of the reaction**, the students decide to use **mean bond enthalpies**. Their teacher reminds them that the enthalpy change of a reaction can be approximated using the relationship:

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

The following **mean standard bond enthalpies** are provided to help the students complete their investigation:

Bond	**Mean bond enthalpy (kJ mol⁻¹)**
C = O	743
C – H	413
C – O	360
C – C	348
O – H	463

Task: As a chemistry student who is a member of the chemistry project group:

(a) Identify the bonds **broken** and the bonds **formed** during the reaction.

(b) Using the given bond enthalpies, **calculate the standard enthalpy change (ΔH)** for the reaction.

(c) State whether the reaction is **endothermic or exothermic** and briefly explain your reasoning.

Response:

Bonds broken are: 1 C–H bond = (1 x 413) = 413 kJmol^{-1}
1 C=O bond = (1 x 743) = 743 kJmol-1

Total enthalpy of bonds broken 413 + 743 = 1156 kJmol^{-1}

Bonds formed are: 1 C – O bond = (1 x 360) = 360 kJmol^{-1}
1 C – C bond (1 x 348) = 348 kJmol^{-1}
1 O - H bond (1 x 463) = 463 kJmol^{-1}

Total enthalpy of bonds formed = 360 + 348 + 463 = 1172 kJmol^{-1}

Note that; when bonds are broken, heat is absorbed and the enthalpy change will be positive.

$\Delta H = +1156$ kJmol^{-1}

However, when bonds are formed, heat energy is released and enthalpy change will be negative.

$\Delta \text{H} = -1171$ kJmol^{-1}

∴ ΔH reaction = Total enthalpy of bonds broken and formed

= +1156 + (-1171) = -15 kJmol^{-1}.

Alternative method:

$$\Delta H_{reaction} = \left(\sum \begin{matrix} of\ bond\ energies \\ of\ bonds\ broken \end{matrix}\right) - \left(\sum \begin{matrix} bond\ enegies \\ of\ bonds\ formed \end{matrix}\right)$$

= 1156 - 1171 = -15 kJmol^{-1}

<u>Enthalpy of Solution/Dissolution</u> and **<u>Hydration/Solvation Energy</u>**

Enthalpy of solution is enthalpy change that occurs when one mole of an ionic compound is dissolved in a specific amount of water (a solvent) to form an infinitely dilute solution.

Alternatively, it is the heat change that occurs when one mole of an ionic compound dissolves in water (a solvent) to form an infinitely dilute solution.

Hydration energy (Enthalpy of hydration) is the heat energy evolved when one mole of gaseous ions is completely surrounded by water molecules to form an infinitely dilute solution.

Alternatively, it is the heat energy evolved when one mole of gaseous ions dissolves in water to form an infinitely dilute solution.

Enthalpy of solution of affected by the magnitude of:

- the lattice energy. The higher the lattice energy, the less easily it can be overcome by the hydration energy and hence the lower the enthalpy of solution. The reverse is true.
- the hydration energy. The higher the hydration energy, the more easily it can overcome the lattice energy and hence the higher the enthalpy of solution. The reverse is true.

Hydration energy is affected or influenced by two factors:

- the *ionic charge*. The higher the ionic charge, the higher the hydration energy. This is because when the charge on the ion is high, the stronger the attraction of the ion for water molecules and more heat energy is released. The reverse is true.
- the *ionic radius*. The larger the ionic radius, the lower the hydration energy. This is because when the radius of the ion is large, the weaker the attraction of the ion for water molecules and less heat energy is evolved. The reverse is true.

Note that: The amount of solvent used determines whether the enthalpy changes involved in dissolution of an ionic salt in water is either positive or negative.

Dissolution of ionic salts involves two processes;

(i) Separation of solid into gaseous ions

$$MX_{(s)} \longrightarrow M^+_{(g)} + X^-_{(g)} \quad \Delta H_{U_L} = \text{positive (endothermic)}$$

(ii) Hydration of the gaseous ions

$$\left.\begin{aligned} M^+_{(g)} + (aq) &\longrightarrow M^+_{(aq)} \\ X^-_{(g)} + (aq) &\longrightarrow X^-_{(aq)} \end{aligned}\right\} \Delta H_{hydration} = \text{negative (exothermic)}$$

$$\Delta H_{solution} = \Delta H_{U_L} + \Delta H_{hydration}$$

Using the Born – Haber cycle to apply Hess' Law,

$$\begin{tikzcd} MX_{(s)} \arrow[r, "\Delta H_{solution}"'] \arrow[d, "\Delta H_{U_L}"'] & M^+_{(aq)} + X^-_{(aq)} \\ M^+_{(g)} + X^-_{(g)} \arrow[ur, "\Delta H_{hydration}"'] & \end{tikzcd}$$

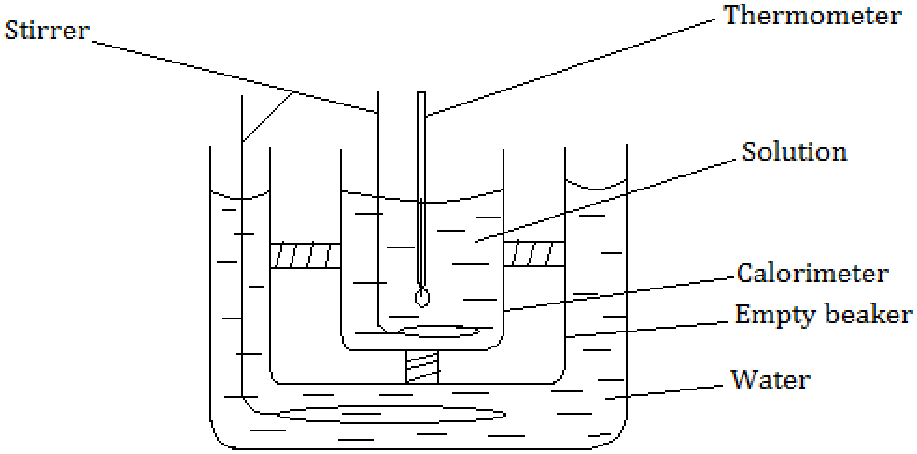
For any ionic compound to dissolve in water, the lattice energy of the system should be overcome or outweighed (cancelled out). Thus, the hydration energy should be numerically greater in order to give an overall negative value.

For soluble salts, the enthalpy of solution is usually negative since is dissolution is spontaneous. The more negative the value for enthalpy of solution, the more soluble the ionic salt.

For sparingly soluble salts (insoluble salts), the enthalpy of solution is usually positive since is dissolution is not very spontaneous. The less positive the value for enthalpy of solution, the more sparingly soluble (less insoluble) the ionic salt.

Experiment to determine the enthalpy of solution of an ionic salt:

Using a measuring cylinder, a known amount (volume/mass) of water, w_g is placed in an insulated calorimeter (plastic container) and its initial temperature, T_1 is read and recorded using a thermometer.
A known mass, m_g of the salt was weighed using a weighing scale and then added to the water in the insulated calorimeter. The mixture was stirred using the thermometer (or shaken thoroughly) and the lowest (or highest) / minimum (or maximum) temperature, T_2 is read and recorded.
The apparatus is set as shown in the diagram below.



Assumptions:
Density of solution = density of water = $1 gcm^{-3}$.
Specific heat capacity of solution = specific heat capacity of water = $4.2\ Jg^{-1}\ {}^0C^{-1}$.
Heat capacity of the insulated calorimeter is negligible (No heat gained by insulated calorimeter).
Treatment of results:
Let; Temperature change be Δt i.e. $(T_2 - T_1)$ or $(T_1 - T_2)$
Thus, the heat of solution = heat lost or gained by water.
Heat change by m_g of substance = $(w + m) \times 4.2 \times \Delta t$ joules.
If RFM of substance is equal to M
m_g of substance causes an energy change of $(w + m)\ c\ \Delta t$ joules
M_g of substance causes an energy change of $\frac{(w+m) \times 4.2 \times \Delta t}{m} \times M$ Joules
Thus heat change by Mg of substance = $\frac{(w+m) \times 4.2 \times \Delta t}{m} \times M$ Joules
But 1 mole of a substance weighs Mg
$\therefore$ heat of solution = + or - $\frac{(w+m) \times 4.2 \times \Delta t}{m \times 1000} \times M$ $kJmol^{-1}$
The heat of solution has either a negative enthalpy change if there was a temperature rise or a positive enthalpy change if there was a temperature fall.
Note that:
1. The effect of temperature on solubility of an ionic compound depends on whether the enthalpy of solution is endothermic or exothermic.
(a) If the enthalpy of solution is endothermic (has a positive value), it is favoured by high temperature. Thus, increase in temperature, increases the solubility of the solute. The reverse is true.
(b) If the enthalpy of solution is exothermic (has a negative value), it is favoured by low temperature. Thus, increase in temperature, decreases the solubility of the solute. The reverse is true.
2. (a) The enthalpy of solution of *hydrated* sodium sulphate to be **lower (more exothermic or less endothermic)** than that of *anhydrous* sodium sulphate. i.e. i.e. Hydrated sodium sulphate
($\Delta H \approx -3.9$ kJ/mol) has a lower enthalpy of solution than anhydrous sodium sulphate ($\Delta H \approx +2.4$ kJ/mol)
Explanation:
Anhydrous sodium sulphate (Na_2SO_4) dissolves in water with an enthalpy of solution that includes:

- Breaking the ionic lattice of the anhydrous solid (lattice energy which is always endothermic).
- Hydration of the dissolved ions (hydration energy which is always exothermic).

Hydrated sodium sulphate (e.g., $Na_2SO_4 \cdot 10H_2O$) already has water of crystallization in its structure. To dissolve:

- The crystal lattice of the hydrate breaks apart but this lattice is *weaker* (lower lattice energy) than the anhydrous lattice because the water molecules disrupt the strong ionic forces. Hence less energy is absorbed to overcome the weaker interactions between the hydrated ions and the water molecules in the crystal.
- The water of crystallization is already partly hydrating the ions, so less energy is needed to fully hydrate them in solution.

Because the hydrate has a weaker lattice and requires less energy to separate its ions, the **overall enthalpy of solution is lower in magnitude** than for the anhydrous salt.

🔎 In simpler terms: ✅ **Reason in one sentence**:

The enthalpy of solution of hydrated sodium sulphate is lower because its crystal lattice is weaker and partly pre-hydrated, requiring less energy to break apart and hydrate in water.

(b) **Hydrated copper(II) sulphate dissolves in water with absorption of heat energy** (i.e., why its enthalpy of solution is **endothermic**).

i.e. Anhydrous $CuSO_4$ dissolves exothermically (–66.5 kJ/mol) while hydrated $CuSO_4 \cdot 5H_2O$ dissolves endothermically (+11.7 kJ/mol)

✅ **Key idea**: When a salt dissolves, two main energy changes happen:

The **lattice enthalpy** (breaking the crystal lattice into separate ions) – this always **requires** energy (endothermic).

The **hydration enthalpy** (hydrating the free ions with water) – this **releases** energy (exothermic).

✅ For **hydrated copper(II) sulphate**: Even though it's already partly hydrated in the solid (the Cu^{2+} is coordinated to water molecules in the crystal), these interactions are quite strong in the crystal structure.

To dissolve, these interactions in the solid must be **broken** and this requires energy.

The hydration energy (energy liberated) when the ions and water molecules rearrange in solution **is less** than the energy required to break the solid's structure.

✅ **Reason it absorbs heat overall**: The energy required to **separate the hydrated ionic lattice** (which includes both ionic bonds and hydrogen-bonded lattice water) is **greater** than the energy released when these ions are hydrated by bulk water.

Therefore, the overall **enthalpy of solution is positive** (endothermic).

✅ **Simple explanation**: Hydrated copper(II) sulphate dissolves with absorption of heat because more energy is needed to break apart its strongly bonded crystal structure (including water of crystallisation) than is released when the separated ions become hydrated in water.

✅ **In one line**: *It absorbs heat because breaking the crystal structure needs more energy than is released by hydration.*

Experimental determination of enthalpy of solution by use of a graph:

Usually, a known mass of the given salt (solute) is added to a known volume of water (solvent) whose initial temperature has been noted and recorded.

The temperature of the solution is recorded at regular time intervals with constant stirring using a thermometer.

The temperature begins to rise for exothermic reactions, and then it either remains constant or decreases slightly, due to cooling.

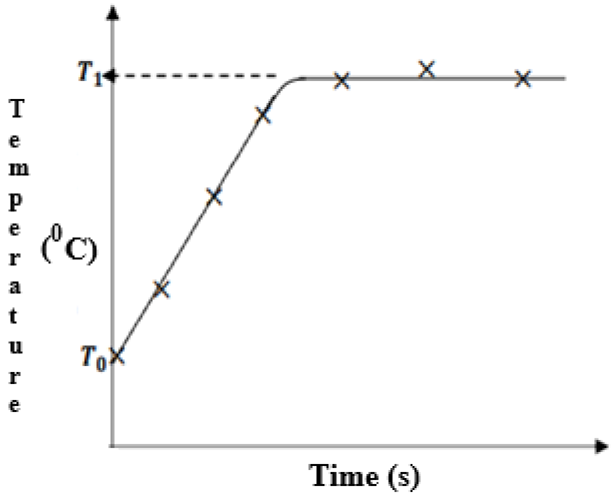
For endothermic reaction, the temperature begins to fall due to absorption of heat from the surroundings, and then it either remains constant or increases slightly due to maximum amount of heat absorbed.

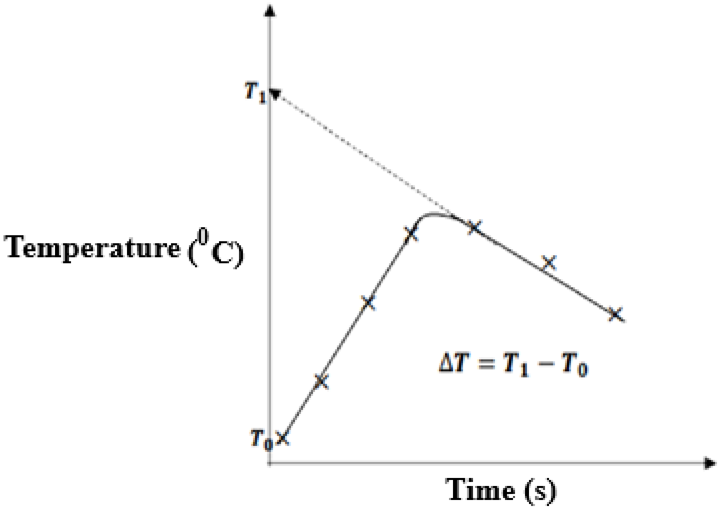
Procedure:

Using a measuring cylinder, a *known volume, Vcm^3* of water is transferred into an *insulated calorimeter (plastic beaker or container)* and the *initial temperature, T_1 ^0C* of the solution is read using a thermometer and recorded. Using a weighing scale, a *known mass of the solute* is weighed and then added to the water in the plastic container and a stop clock is simultaneously started. The mixture is *stirred* thoroughly using a thermometer (or gently *shaken*) and the temperature is read and recorded every 30 seconds for a period of 4 minutes. The time and temperature are recorded in a table.

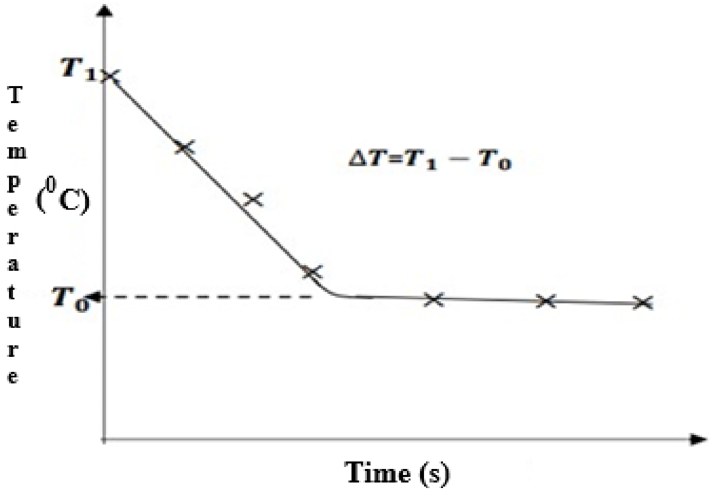
A graph of temperature against time is plotted.

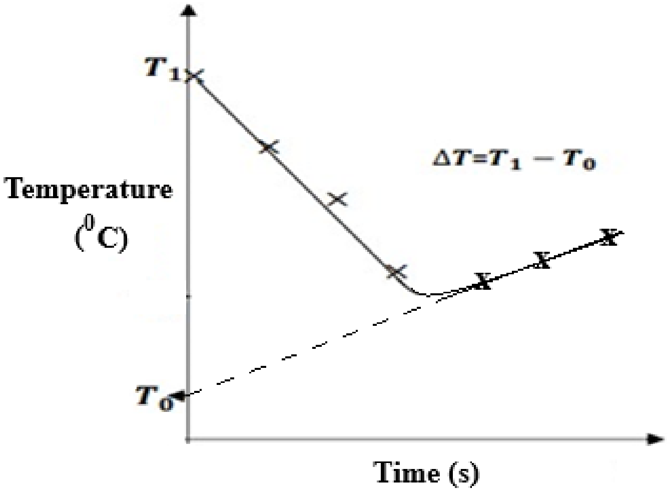
(a) For exothermic reactions.





(b) For endothermic reactions.





Trial items:

1. During a chemistry practical session, students were investigating the energy changes that occur when a salt dissolves in water. The teacher provided **5.35 g of ammonium chloride (NH_4Cl)** and **100 cm^3 of water** in a plastic beaker. The students stirred the mixture gently with a thermometer and recorded the temperature changes over time as shown in the table below.

Time (seconds)	**0**	**20**	**40**	**60**	**80**	**120**	**140**	**160**	**180**	**200**
Temperature (°C)	25	23	21.0	19.2	18.0	17.4	17.5	18.4	19.2	20.0

Tasks: As a chemistry student, based on the experiment above:

(a) Explain what is meant by the **enthalpy of solution**.

(b) Describe **two energy factors** that influence the enthalpy of solution when a salt dissolves in water.

(c) Use the data provided to:

(i) Plot a **graph of temperature (°C) against time (seconds)** to show the change in temperature during dissolution.

(ii) Find out the **molar enthalpy of solution of ammonium chloride**, given that; Density of water = 1.0 g/cm³; Specific heat capacity of the solution = 4.2 J g^{-1} $°C^{-1}$; Molar mass of NH_4Cl = 53.5 g/mol;

(iii) State any **assumptions** you make in your calculations.

(d) After observing the results, the students noticed that heating the solution caused more NH_4Cl to dissolve.

Explain the **effect of increasing temperature** on the solubility of ammonium chloride, and give a reason for your answer.

2. A group of learners investigated energy changes involved when salts dissolve in water. They carried out experiments using **anhydrous copper(II) sulphate** and **hydrated copper(II) sulphate ($CuSO_4{\cdot}5H_2O$)** to determine their respective enthalpies of solution.

When **4.0 g of anhydrous $CuSO_4$** was added to **50 g of water**, the temperature increased by **8.0°C**.

When **4.0 g of hydrated $CuSO_4{\cdot}5H_2O$** was added to **50 g of water**, the temperature **dropped from 24.5°C to 23.6°C**.

(Specific heat capacity of solution = 4.2 J g^{-1} K^{-1})

Task:

(a) Before conducting the experiment, the teacher asked the learners to recall a key law used in determining enthalpy changes.

State **Hess's Law of constant heat summation**.

(b) The students were also asked to define key energy terms related to ionic compounds.

(i) Explain the term **hydration energy** and write a chemical equation to illustrate it.

(ii) Explain the term **lattice energy** and write a chemical equation to illustrate it.

(c) Using your knowledge of chemical bonding, explain **two factors** that affect the value of the lattice energy of an ionic compound.

(d) Using the experimental data provided; find out the **enthalpy of solution** of:

(i) Anhydrous copper(II) sulphate.

(ii) Copper(II) sulphate-5-water.

(Assume density of solution = 1.0 g/cm³)

(e) From your results in (d) above:

(i) State which one of the two copper(II) salts is **more soluble in water**, and explain your reasoning.

(ii) Using a **Born–Haber cycle**, determine the **hydration energy** of anhydrous copper(II) sulphate.

3. During a cold morning in the school laboratory, a group of students is asked to investigate which salt could be used in a chemical hand warmer. They are given **0.5 g of calcium chloride ($CaCl_2$)** and **100 cm³ of water at 24.0°C**. The mixture is stirred, and the temperature is recorded at regular time intervals.

The data they obtained are shown below:

Time (minutes)	0	½	1	1½	2	2½	3	3½	4	4½	5	5½	6
Temperature (°C)	24.0	51.0	55.5	57.5	57.0	56.5	56.0	55.0	54.0	53.5	52.5	52.0	51.0

The **specific heat capacity** of the solution is **4.2 J g^{-1} $°C^{-1}$**, and the **density of water** is **1.0 g cm^{-3}**.

Task: As a chemistry student;

(a) Plot a **graph of temperature (°C)** of the solution against **time (minutes)** using the data provided.

(b)(i) From your graph, determine the **maximum temperature rise** during the dissolution process.

(ii) Using the data and given constants, calculate the **molar enthalpy of solution** of calcium chloride.

(Assume the mass of the solution = mass of water + mass of $CaCl_2$)

(iii) In a previous experiment, another group tested a different salt that released less heat when dissolved. Based on your results and calculations, **compare the heat of dissolution of calcium chloride** with the other salt and **justify which substance is better for a hand warmer.**

Standard Enthalpy of Neutralisation

This is heat change that occurs when one mole of hydrogen ions from an acid and one mole of hydroxide ion from a base react to form one mole of water at standard conditions.

Experiment to determine the enthalpy of neutralization:

The procedure can also be used to determine the enthalpy of any reaction that takes place in solution.

e.g To determine the enthalpy of neutralization of a monobasic strong acid against a strong base.

Using a measuring cylinder, a known volume V_1 cm^3 of a monobasic acid of known concentration M_1 is transferred into a plastic container or insulated calorimeter and its initial temperature T_1 ^{0}C is read using a thermometer and noted.

Using a measuring cylinder, a known volume V_2 cm^3 of a strong base with a known concentration M_2 is placed into another container and its initial temperature T_2 ^{0}C is using a thermometer and noted. The amount of the strong base (alkali) must be equal to or more than the acid.

The alkali is then transferred to the acid as rapidly as possible, the mixture stirred using the thermometer (or shaken thoroughly) and the highest (maximum / final steady) temperature T_3 ^{0}C, during the reaction is read and noted.

Assumptions made:

- the density of solution to be the density of water = 1gcm^{-3}. This is because water is the major product of neutralization.
- the specific heat capacity, c of the solution is equal to that of water = 4.2 Jg^{-1} ^{0}C^{-1}
- the specific heat capacity of the plastic container is negligible.
- a plastic beaker or insulated container is used because it is a poor conductor that is able to minimize loss of heat energy to the surroundings unlike a glass or metallic beaker.

Treatment of results:

Initial temperature = $(\frac{T_1 + T_2}{2})$ ^{0}C   Rise in temperature = $T_3 - \left(\frac{T_1 + T_2}{2}\right) = t$ ^{0}C (exothermic)

Mass of solution = (density x volume) of solution = 1 x ($V_1 + V_2$) = ($V_1 + V_2$)g

Heat change = mass of solution x specific heat capacity x temperature rise

= m x c x t = ($V_1 + V_2$) x 4.2 x t Joules

But 1000cm^3 of solution contain M_1 moles of acid

V_1 cm^3 of solution contain $\left(\frac{M_1 \text{ x } V_1}{1000}\right)$ moles of acid

1 mole of monobasic acid contains 1 mole of hydrogen ions

Moles of hydrogen ions = $\left(\frac{M_1 \text{ x } V_1}{1000}\right)$ moles

Then $\left(\frac{M_1 \text{ x } V_1}{1000}\right)$ moles of hydrogen ions from the acid produce ($V_1 + V_2$) x 4.2 x t Joules

1 mole of hydrogen ions produces $\left((V_1 + V_2)\text{x } 4.2 \text{ x t} \Big/ \frac{M_1 \text{ x } V_1}{1000}\right)$ Joules

$\Delta H^{\theta}_{\text{neutralisation}}$of monobasic acid by the base $= -\left(\frac{(V_1 + V_2) \text{ x } 4.2 \text{ x t}}{M_1 \text{ x } V_1}\right)$ kJmol^{-1}

Note that:

(a) In **a neutralisation (an enthalpy) experiment**, the increase in temperature depends on the:

(i) **Strength of acid/Base**. e.g.

- Nitric acid is a strong acid and thus it is fully ionised. Hence there is a higher temperature rise.
- Ethanoic acid is a weak acid and thus is partially ionised. Hence there is a lower temperature rise.

(ii) **Concentration of acid/Base.** The higher the concentration, the higher the temperature.

(b) The **shape of graph in a neutralisation experiment** gives a "rise–plateau–fall" trend. This is because when the acid/base is neutralised, the extra base/acid added just dilutes and cools the mixture.

(c) The **enthalpies of neutralization of strong acids by strong bases or strong bases by strong acids are almost constant and higher**. e.g.

(i) Reaction of the nitric acid, HNO_3 with potassium hydroxide, KOH is -57.3 kJmol^{-1}.

(ii) Reaction of sulphuric acid, H_2SO_4 with sodium hydroxide, NaOH is -61.3 kJmol^{-1}.

(iii) Reaction of nitric acid, HNO_3 with sodium hydroxide, NaOH is -57.1 kJmol^{-1}

This is because they are strong acids and strong bases which undergo complete ionisation.

i.e. $NaOH_{(aq)} \longrightarrow Na^+_{(aq)} + \overline{O}H_{(aq)}$ $HCl_{(aq)} \longrightarrow H^+_{(aq)} + Cl^-_{(aq)}$

Thus, the only reaction between strong acids and strong bases is involving hydrogen ions from the acid and hydroxide ions from the base. $H^{+}_{(aq)} + {}^{-}OH_{(aq)} \longrightarrow H_2O_{(l)}$

(d) **The enthalpies of neutralization of weak acids / bases by strong acids / bases (or weak acids / bases) or the reverse are not constant and are lower**. e.g.

(i) Reaction of the weak hydrocyanic acid, HCN with sodium hydroxide, NaOH has enthalpy of neutralization as -12.0 $kJmol^{-1}$.

(ii) Reaction of hydrochloric acid, HCl with the weak ammonia, NH_3 has enthalpy of neutralization as -51.5 $kJmol^{-1}$.

This is because weak acids / bases do not undergo complete ionisation (partially ionize).

Considering the reaction above between the weak hydrocyanic acid and sodium hydroxide:

The reaction involves further ionization of the weak hydrocyanic acid an endothermic process (energy is absorbed to effect the complete ionization of the weak acid, $\Delta H = +45.1\ kJmol^{-1}$) and then later energy is liberated during the neutralization reaction ($\Delta H = -57.1\ kJmol^{-1}$).

As a result, the overall (total) enthalpy change, $\Delta H = +45.1 + -57.1 = -12.0\ kJmol^{-1}$.

Experimental determination of enthalpy of neutralisation by use of a graph:

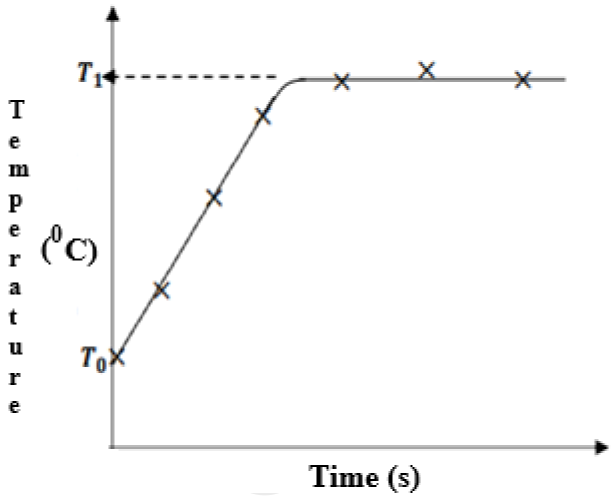
Procedure:

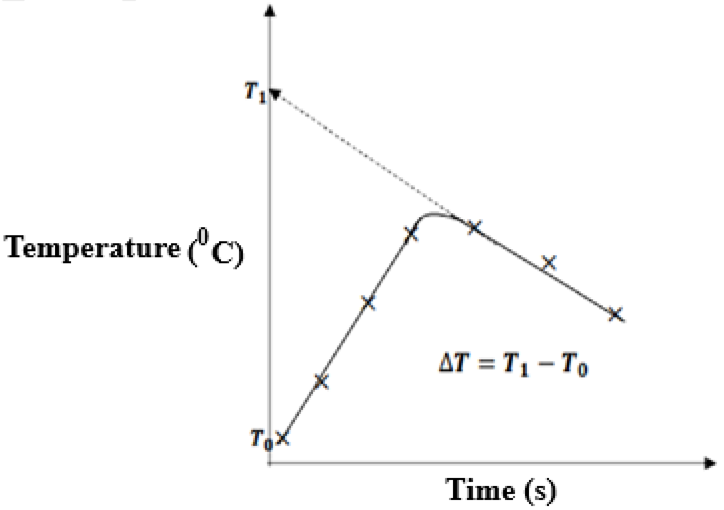
Using a measuring cylinder, a known volume V_1 cm^3 of a monobasic acid of known concentration M_1 is transferred into a plastic container or insulated calorimeter and its initial temperature T_1 ^0C is read using a thermometer and noted.

Using a measuring cylinder, a known volume V_2 cm^3 of a strong base with a known concentration M_2 is placed into another container and its initial temperature T_2 ^0C is using a thermometer and noted. The amount of the strong base (alkali) must be equal to or more than the acid.

The alkali is then transferred to the acid in the plastic container as rapidly as possible, the mixture stirred using the thermometer and a stop clock is simultaneously started. The mixture is *stirred* thoroughly using a thermometer (or gently *shaken*) and the temperature is read and recorded every 30 seconds for a period of 4 minutes. The time and temperature are recorded in a table.

A graph of temperature against time is plotted.





Alternative procedure:

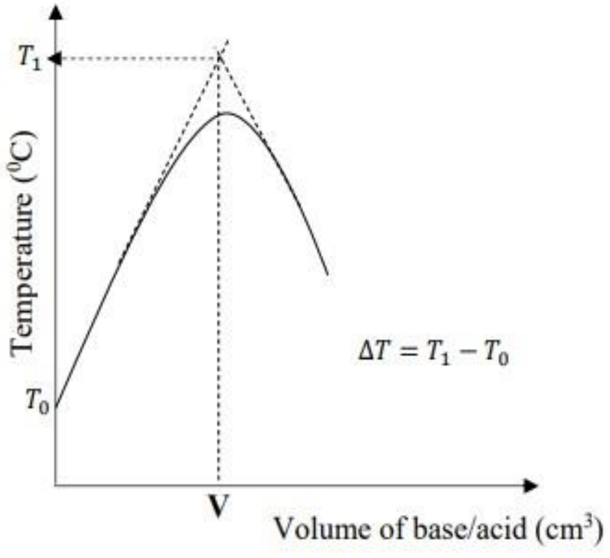
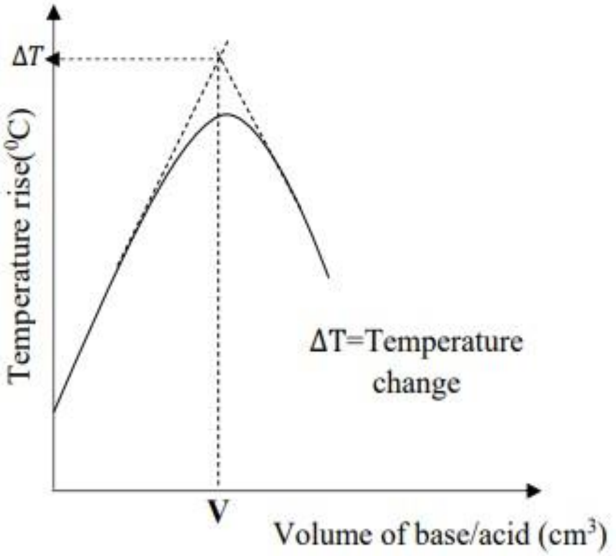
(a) The initial temperature T_1 ^0C of a monobasic acid with a known concentration M_2 in a plastic container is read and recorded using a thermometer.

The initial temperature T_2 ^0C of a strong base with a known concentration M_2 in another container is also read and recorded using a thermometer.

(b) Using a burette, a known volume V_1 (e.g. 25) cm^3 of the monobasic acid of known concentration M_1 is transferred into a plastic container or insulated calorimeter.

(c) Using a measuring cylinder, at once a known fixed volume 10.0 cm^3 of the strong base is added at uniform intervals to the acid in the plastic container. The mixture is *stirred* thoroughly using a thermometer (or gently *shaken*) and the final maximum temperature is read and recorded for readings up to 10.0 cm^3.

The results are recorded in a table and a graph of temperature (or temperature rise) against volume of base added is plotted.

<table>
<tr><th>(i) A graph of temperature against volume of alkali/base added.</th><th>(ii) A graph of temperature rise against volume of alkali (or acid added).</th></tr>
<tr><td>
</td><td>
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Sample Items:

1. In Hoima district, a local water treatment company investigated various samples of acidic water containing hydrochloric acid obtained from factories and the company often neutralizes the acidic water before discharge. Understanding the molar heat of neutralization of the acidic water helps chemical engineers from the water treatment company to design safe processes for discharge avoiding overheating and the best process is one that produces less heat of neutralization.

Okoth, a new trainee engineer in the company was confused on whether to add sodium hydroxide solution or ammonia solution to the acidic waters before discharge. The laboratory technician helped Okoth by carrying out experiments on the acidic water and obtained the following data;

Table I;

Volume of NaOH solution (cm^3)	10	20	30	40	50	60	70	80
Final temperature (^{0}C)	27.8	29.9	31.5	32.2	32.0	31.0	30.0	29.0
Initial temperature (^{0}C)	25.0	25.0	25.0	25.0	25.0	25.0	25.0	25.0
Temperature rise (^{0}C)								

Table II;

Volume of ammonia solution (cm^3)	10	20	30	40	50	60	70	80
Final temperature (^{0}C)	26.4	27.8	28.6	29.0	28.8	28.0	27.8	27.0
Initial temperature (^{0}C)	25.0	25.0	25.0	25.0	25.0	25.0	25.0	25.0
Temperature rise (^{0}C)								

During the experiments, the laboratory technician measured 30cm^3 of a 1M solution of the acidic water into a plastic beaker and added a 1M solution of alkali separately using a burette and temperature recorded as shown in the tables above.

Task: As a S.5 student with knowledge about thermochemistry, help Okoth to;

(a) Understand what is meant by the term molar heat of neutralization and write the ionic equation representing the process. (02 scores)

(b) Plot on the same axes, the graph of temperature rise against volume of alkali solution added. (08 scores)

(c) Determine with EVIDENCE, the best approach to neutralize the acidic waters before discharge.
(Density of solution is $1 gcm^{-3}$, specific heat capacity of solution is $4.2 Jg^{-1}\,^{0}C^{-1}$) (10 scores)

2. Hope Secondary School in Gulu District is planning to set up a small community laboratory to help farmers test soil acidity. Before launching the project, the school's science club conducts an investigation to compare the heat energy released when **a strong acid (nitric acid, HNO_3)** and **a weak acid (acetic acid, CH_3COOH)** react with **sodium hydroxide (NaOH)**.

The laboratory technician told the science club students to use the following procedure:

- Measure **25 cm^3** of **1 M acid** into a polystyrene cup (first using HNO_3, then repeating with CH_3COOH).
- Add **1 M NaOH** in **steps of 5 cm^3** up to **40 cm^3**.
- After each addition, gently stir and record the **maximum temperature** reached.

The room temperature is **25 ^{0}C**. The table below shows the results obtained:

Volume of NaOH added (cm^3)	**Max Temp with HNO_3 (^{0}C)**	**Max Temp with CH_3COOH (^{0}C)**
0	25.0	25.0
5	28.0	26.5
10	31.5	28.0
15	35.0	30.0
20	38.0	32.0
25	39.5	33.0
30	38.5	32.5
35	37.0	31.5
40	35.0	30.0

However, the science club members needed your assistance on data interpretation and analysis.

Task: As a chemistry student;

(a)(i) **On the same axes,** plot an appropriate graph for both nitric acid and ethanoic acid.

(ii) From your graph, determine; the **volume of NaOH at which maximum temperature** occurs for each acid and explain why one acid produces more heat than the other.

(iii) Calculate the enthalpy of neutralisation of sodium hydroxide separately by each acid.

(b) Since the school hopes to use this information to help farmers correct soil acidity.

(i) Explain to the farmers why knowing **how much heat is released** during neutralisation may be useful when designing safe soil-treatment procedures.

(ii) If a farmer accidentally adds too much lime, $Ca(OH)_2$ to soil, **predict**; what will happen to soil pH, how it may affect crop growth and how the farmer can correct the imbalance

(c)(i) Identify **three safety precautions** that learners must observe when handling acids and bases during this experiment.

(ii) Suggest **two environmentally responsible practices** the school should follow when disposing of the used acid-base mixtures.

Standard Enthalpy of Displacement

This is the enthalpy change that occurs when one mole of elements in the standard state is displaced from compounds containing its ions.

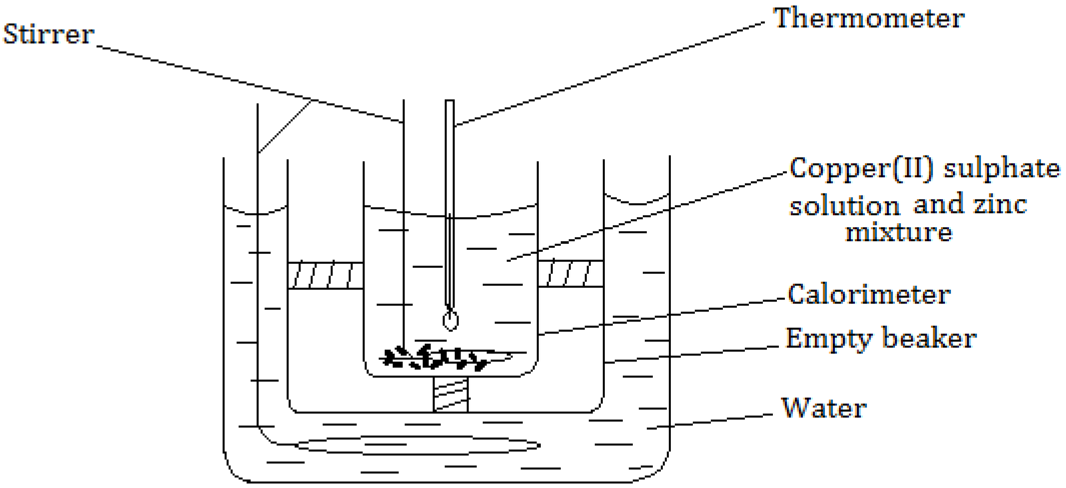
Experiment to determine the enthalpy of displacement:

e.g. Determination of the enthalpy of displacement of copper in the reaction between copper(II) sulphate solution and zinc.

Using a measuring cylinder, a *known volume, Vcm^3* of copper(II) sulphate solution with a *known concentration, Y moldm^{-3}* is transferred into an *insulated calorimeter (plastic beaker or container)* and the *initial temperature, T_1 ^{0}C* of the solution is read using a thermometer and recorded.

Using a weighing scale, an *excess* of zinc powder is weighed and then added to the copper(II) sulphate solution and the mixture *stirred* thoroughly using a thermometer (or gently *shaken*) and the *final (maximum or highest) temperature, T_2 ^{0}C* reached is read from the thermometer and recorded.

The apparatus is set as shown in the diagram below.



Treatment of results:

Assumptions:

- the density of solution to be the density of water = 1gcm^{-3}.
- the specific heat capacity, c of the solution is equal to that of water = 4.2 Jg^{-1} ^{0}C^{-1}
- the specific heat capacity of the plastic container is negligible.
- a plastic beaker or insulated container is used because it is a poor conductor that is able to minimize loss of heat energy to the surroundings unlike a glass or metallic beaker.

Temperature change be Δt i.e. ($T_2 - T_1$)

Mass of copper(II) sulphate solution = density x volume = 1 x V = Vg

Thus, heat energy liberated on displacement of copper
= heat gained by the solution mixture.

Enthalpy change of displacement of copper by m_g of zinc = (V x 4.2 x Δt) Joules.

Moles of copper(II) sulphate solution = ($\frac{Y}{1000}$ x V) moles

$$Zn(s) + CuSO_4(aq) \longrightarrow Cu(s) + ZnSO_4(aq)$$

Moles of copper displaced = ($\frac{Y}{1000}$ x V) moles

Either:

($\frac{Y}{1000}$ x V) moles of copper are displaced with a heat change of (V x 4.2 x Δt) Joules

1 mole of copper is displaced with a heat change of $\frac{(V \times 4.2 \times \Delta t)}{Y \times V}$ x 1000 Joules

Hence enthalpy of displacement = - $\frac{(4.2 \times \Delta t)}{Y}$ kJmol^{-1}

Or

If 1 mole of copper weighs 63.5g, then

($\frac{Y}{1000}$ x V x 63.5)g of copper are displaced with an energy change of (V x 4.2 x Δt) Joules

63.5_g of copper are displaced with an energy change of $\frac{(V \times 4.2 \times \ \Delta t)}{Y \times V}$ x 1000 Joules

∴ Enthalpy of displacement of copper = - $\frac{(4.2 \times \Delta t)}{Y}$ kJ mol^{-1}

Or

Let w_g of copper be displaced by zinc and 1 mole of copper weighs 63.5g, then
w_g of copper are displaced with an energy change of (V x 4.2 x Δt) Joules
63.5_g of copper are displaced with an energy change of $\frac{(V \times 4.2 \times \ \Delta t)}{w}$ x 63.5 Joules

∴ Enthalpy of displacement of copper = - $\frac{(V \times 4.2 \times \Delta t)}{w \times 1000}$ x 63.5 kJ mol^{-1}

Alternatively; Using a measuring cylinder, a known volume, Vcm^3 of copper(II) sulphate solution with a known concentration, Y moldm^{-3} is transferred into an insulated calorimeter (plastic container) and the initial temperature T_1 of the solution is read using a thermometer and recorded. Using a weighing scale an excess of zinc powder, m_g is weighed and then added to the copper(II) sulphate solution. The mixture stirred thoroughly using a thermometer and the final (maximum or highest) temperature, T_2 reached is read and recorded.

Assumptions:

- the density of solution to be the density of water = 1gcm^{-3}.
- the specific heat capacity, c of the solution is equal to that of water = 4.2 Jg^{-1} ^{0}C^{-1}
- the specific heat capacity of the plastic container is negligible.
- a plastic beaker or insulated container is used because it is a poor conductor that is able to minimize loss of heat energy to the surroundings unlike a glass or metallic beaker.

Temperature change be Δt i.e. ($T_2 - T_1$)
Mass of copper(II) sulphate solution = density x volume = 1 x V = Vg
Thus, heat energy liberated on displacement of copper
= heat gained by the solution mixture.
Molar mass of copper is 64g
Moles of copper(II) sulphate solution = ($\frac{Y}{1000}$ x V) moles
Moles of zinc that reacted = ($\frac{Y}{1000}$ x V) moles
Molar mass of zinc is 65g
Mass of zinc that reacted = ($\frac{Y}{1000}$ x V x 65) g
Enthalpy change of displacement of copper by m_g of zinc = [V + ($\frac{Y}{1000}$ x V x 65)] x 4.2 x Δt Joules.

$$Zn(s) + CuSO_4(aq) \longrightarrow Cu(s) + ZnSO_4(aq)$$

Moles of copper metal displaced = ($\frac{Y}{1000}$ x V) moles
($\frac{Y}{1000}$ x V) moles of copper when displaced there is an energy change of (V + m) x 4.2 x Δt Joules
1 mole of copper when displaced there is an energy change of $\frac{[V + (\frac{Y}{1000} \times V \times 65)] \times 4.2 \times \ \Delta t}{Y \times V}$ x 1000 Joules

∴ Enthalpy of displacement of copper = $\frac{[V + (\frac{Y}{1000} \times V \times 65)] \times 4.2 \times \Delta t}{Y \times V}$ kJmol^{-1}

N.B: Enthalpy of displacement has either a negative enthalpy change if there was a temperature rise or a positive enthalpy change if there was a temperature fall.

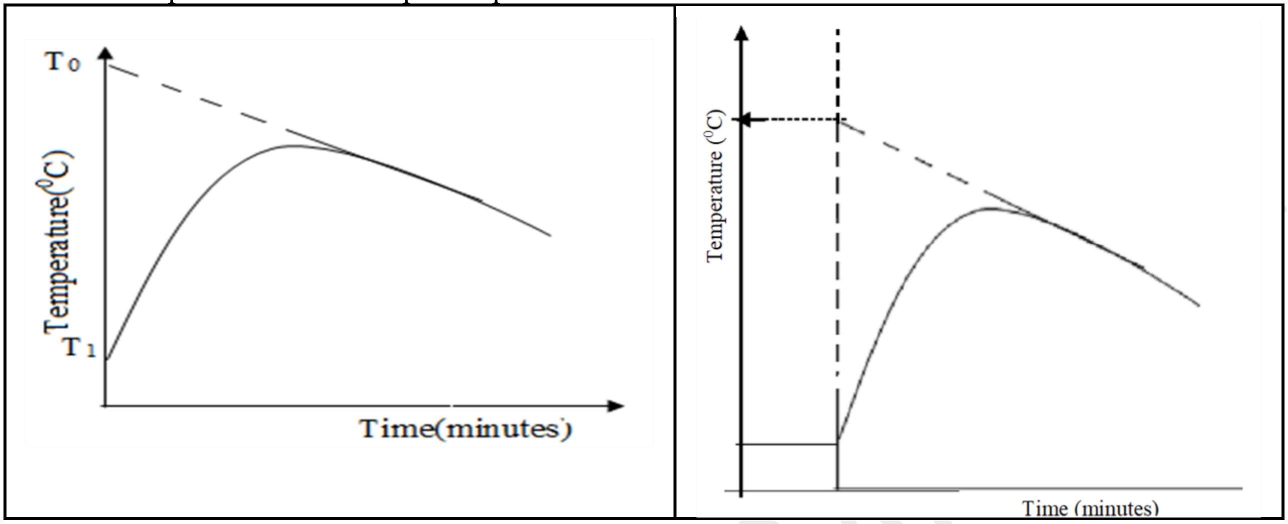
Experimental determination of enthalpy of displacement by use of a graph:

Using a measuring cylinder, a *known volume, Vcm*^{3} of copper(II) sulphate solution with a *known concentration, Y moldm*$^{-3}$ is transferred into an *insulated calorimeter (plastic beaker or container)* and the *initial temperature,* T_1 ^0C of the solution is read using a thermometer and recorded.

Using a weighing scale, an *excess* of zinc powder is weighed and then added to the copper(II) sulphate solution and a stop clock is simultaneously started. The mixture is *stirred* thoroughly using a thermometer (or gently *shaken*) and the temperature is read and recorded every 30 seconds for a period of 4 minutes. The time and temperature are recorded in a table.

A graph of temperature against time is plotted.

The final maximum temperature or temperature change attained in an exothermic reaction can be determined from the graph by extrapolation of the graph backwards using dotted lines starting from the dropping point after the observed plotted maximum point up to the vertical axis.



However, the final minimum temperature or temperature change attained in an endothermic process (e.g. enthalpy of solution of ammonium chloride) can be determined from the graph by extrapolation of the graph backwards using dotted lines starting from the rising point after the observed plotted minimum point down to the vertical axis.

This is done for a graph in which temperature or temperature change is plotted against time.

Sample Items:

1. (a) During a chemistry lesson, your teacher demonstrates a reaction in which zinc metal is added to a blue copper(II) sulphate solution. After a few minutes, a reddish-brown solid is formed, and the temperature of the mixture increases. Using the above scenario, explain what is meant by the *enthalpy of displacement* in this reaction.

(b) A group of students wants to investigate how much heat energy is released when zinc reacts with copper(II) sulphate solution. They are provided with a polystyrene cup, thermometer, measuring cylinder, weighing balance, zinc powder, and copper(II) sulphate solution.

Describe how the students can carry out an experiment to determine the *enthalpy of displacement* for the reaction between zinc and copper(II) sulphate.

(c) After performing the experiment, the students recorded the following data:

- Volume of copper(II) sulphate solution used = 250 cm^3
- Initial temperature of the solution = 25.0 °C
- Final temperature = 31.4 °C
- Amount of zinc powder added = 0.1 moles
- Density of solution = 1 g cm^{-3}
- Specific heat capacity of solution = 4.2 J g^{-1} $°C^{-1}$

Using the information provided, calculate the *enthalpy change of displacement* for the reaction between zinc and copper(II) sulphate.

2. James and Jennifer are S.5 Chemistry students in a certain school, who were given an assignment by their teacher. The assignment was, to investigate which of the two unknown metals, M and T is above the other in the reactivity series and would displace it from its aqueous solution, by carrying out appropriate experiments on the metals.

The two students carried out separate experiments, with James using 25 cm^3 of 1M copper(II) sulphate solution and Jennifer using 20 cm^3 of 0.1M copper(II) sulphate solution.

James investigated metal M and Jennifer investigated metal T.

The results obtained by James are as shown below:

Time / minutes	0.0	1.0	2.0	3.0	4.0	5.0	6.0	7.0
Temperature / ^{0}C	27.0	46.0	60.0	69.5	67.5	65.0	62.0	59.5

Jennifer discovered from her experiment that the temperature of the solution rose from 25.0 to 32.0 °C.

Support: Both metals, M and T form divalent cations (i.e. with formula, M^{2+} and T^{2+} respectively), density of water = 1gcm^{-3}, specific heat capacity of water = 4.2 Jg^{-1}K^{-1}.

Task: As a chemistry student,

(a) Describe the procedure which James followed in carrying out his experiment. (05 scores)

(b) By making appropriate manipulation of the results of each student, find the molar enthalpy change of their respective reactions. Hence, deduce which of the two metals is more reactive. (15 scores)

Standard Enthalpy of Precipitation

This is the enthalpy change that occurs when one mole of an insoluble substance is formed from its ions in the aqueous state at standard conditions.

e.g - the enthalpy of precipitation of lead(II) sulphate from lead(II) nitrate solution and sodium sulphate solution is -50.4 kJmol^{-1}.

Standard Enthalpy of Hydrogenation

This is the enthalpy change that occurs when one mole of unsaturated molecules completely reacts with hydrogen gas to form saturated compounds at standard conditions.

Revision Items:

1. During a visit to **Kavera Community Welding Workshop in Masaka**, Senior Five learners observe that workers use ethylene gas ($CH_2{=}CH_2$) mixed with hydrogen in a controlled burner to produce ethane (CH_3CH_3) for testing flame characteristics. The chemistry teacher tells the class that determining the enthalpy change of this reaction helps understand the energy released during industrial hydrogenation.

You are provided with the following average bond enthalpies (kJ mol^{-1}):

Bond	Enthalpy (kJ/mol)
C–H	416
C=C	612
C–C	348
H–H	436

Task: As a chemistry student, use the data above to calculate the standard enthalpy change for the reaction:

$$CH_2{=}CH_{2(g)} + H_{2(g)} \longrightarrow CH_3CH_{3(g)}$$

2. At the **Kampala Industrial Research Laboratory**, S.6 learners are shown how chemists estimate the energy required to form benzene, a compound used in making plastics and detergents. The laboratory technician gives students the following data:

- Standard enthalpy of formation of benzene (experimental): **+49.2 kJ mol^{-1}.**
- Mean bond enthalpies (kJ mol^{-1}):
  - C–C: 348
  - C=C: 615
  - C–H: 412
- ΔH° for vaporisation of carbon: **718 kJ mol^{-1}.**
- ΔH° for atomisation of hydrogen (per mole of H atoms): **217.5 kJ mol^{-1}.**

Tasks: As a chemistry learner,

(a) Use the information provided to calculate the **theoretical standard enthalpy of formation** of benzene.
(b) Compare your calculated value with the experimental value (+49.2 kJ mol^{-1}) and comment on the difference.

3. At **Ntare School Chemistry Club**, learners are analysing different fuels for their efficiency. One of the fuels is **methane**, commonly used in school laboratories. To better understand its combustion behaviour, they need to estimate the **bond dissociation energy (BDE)** of the C–H bond in methane.
They are provided with the following thermodynamic data:

- Standard enthalpy of formation of methane: **$\Delta H°f = -74$ kJ mol^{-1}**.
- Standard enthalpy of atomisation of carbon: **$\Delta H°a = 720$ kJ mol^{-1}**.
- Bond dissociation energy of H–H: **436 kJ mol^{-1}**.

Task: As a chemistry student, use the information above to calculate the **average C–H bond dissociation energy** in methane.

4. During a field study at a **local pharmaceutical plant in Kampala**, chemistry students observe a process where an alcohol **propan-1-ol ($CH_3CH_2CH_2OH$)** is converted into an ether **methoxyethane ($CH_3CH_2OCH_3$)** as part of a synthesis route. The learners are asked to estimate the **enthalpy change** involved in converting the alcohol to the ether.
The bond enthalpies (kJmol^{-1}) are:

Bond	**Energy (kJ/mol)**
C–C	336
C–H	412
C–O	340
O–H	426

Task: As a chemistry learner, analyse the data:

$$CH_3CH_2CH_2OH_{(l)} \longrightarrow CH_3CH_2OCH_{3(l)}$$

5. A research team is working on understanding the energy changes that occur when fuels are burned. They focus on alkanes, since these are common components of petrol and other fuels. The scientists want to measure and compare their combustion properties.
(a) One student on the team asks: *"Before we continue, can someone remind me what is meant by the term enthalpy of combustion?"*
Task: Respond to the student's question.
(b) To collect accurate data, the group decides to measure the enthalpy of combustion of hexane in the laboratory. Imagine you are part of this team.
Task: Using a simple spirit lamp, describe the experiment you would set up and how you would ensure the results are reliable.
(c) The team then looks at the following data they have compiled for straight-chain alkanes:

Number of carbon atoms	**1**	**2**	**3**	**4**	**5**
Enthalpy of combustion (kJ/mol^{-1})	–890	–1560	–2220	–2877	–3509

The group now discusses their findings:
(i) They decide to draw a graph of enthalpy of combustion (y-axis) against the number of carbon atoms (x-axis). **Task:** Plot a suitable graph.
(ii) From the graph, the team wants to estimate the enthalpy of combustion of **hydrogen** and **hexane**.
Task: Use the graph to determine these values.
(iii) A colleague suggests using the slope and intercept of the graph to predict the enthalpy of combustion of **heptane**.
Task: Carry out this calculation.
(iv) Finally, the group notices the graph has a near-linear trend.
Task: Explain why this shape makes sense in terms of molecular structure and bonding.

6. (a) At a **mining research centre in Karamoja**, a team of young Ugandan chemists is investigating why some ionic compounds obtained from mineral ores dissolve easily in water while others do not.

During their investigation, they realise that two major energy terms—**lattice energy** and **hydration energy**—play a key role in determining this behaviour. One of the interns is struggling to differentiate these two concepts and needs help.

The scientists at the centre later begin analysing the available laboratory data to predict which compounds from the ores might be more soluble in the streams around Mt. Moroto. They suspect that the relationship between lattice energy and hydration energy is critical in determining the solubility patterns observed.

Task: As a chemistry learner who is part of the research team,

(i) explain clearly to the intern the difference between lattice energy and hydration energy. (02 scores)

(ii) Use your knowledge to briefly explain to the scientists how lattice energy and hydration energy influence the solubility of ionic compounds in water. (03 scores)

(b) **A local aluminium-processing project in Jinja** aims to produce aluminium fluoride (AlF_3), an important compound used in aluminium purification. Before scaling up production, the project team needs to understand the thermodynamics involved and the following data was collected by the project chemists:

- Standard enthalpy of formation of aluminium fluoride = **–1301 kJ mol^{-1}**.
- Standard enthalpy of atomization of aluminium = **+314 kJ mol^{-1}**.
- Standard enthalpy of bond dissociation of fluorine gas = **+158 kJ mol^{-1}**.
- First ionization energy of aluminium = **+577 kJ mol^{-1}**.
- Second ionization energy of aluminium = **+1820 kJ mol^{-1}**.
- Third ionization energy of aluminium = **+2740 kJ mol^{-1}**.
- First electron affinity of fluorine = **–348 kJ mol^{-1}**.

One of the new technicians is unsure about what the term *standard enthalpy of formation* means.

To evaluate the energy requirements of producing aluminium fluoride, the team wants to represent the steps involved from the raw elements to form solid aluminium fluoride, AlF_3.

The project also wishes to predict whether aluminium fluoride will dissolve significantly in water during industrial handling. They obtain the following additional hydration energy values from data books:

- Hydration energy of Al^{3+} = **–4690 kJ mol^{-1}**.
- Hydration energy of F^- = **–364 kJ mol^{-1}**.

While preparing a safety and environmental manual for the project, the team realizes that they need to understand factors that influence hydration energies of ions, as these determine how chemicals will behave when discharged into water bodies like Lake Victoria.

Task: As a chemistry student, help the project team to;

(i) Enable the technician understand what is meant by the term standard enthalpy of formation using the aluminium fluoride project as context. (01 score)

(ii) Represent the steps involved from raw materials by drawing a well-labelled Born–Haber energy level diagram showing the formation of aluminium fluoride from its constituent elements. (06 scores)

(iii) Use the diagram above and the data obtained to calculate the lattice energy of aluminium fluoride. (02 scores)

(iv) Use these values to find out the enthalpy of solution of aluminium fluoride and from the result, comment on whether AlF_3 is expected to be highly soluble in water with a reason. (04 scores)

(v) State and explain **any two** factors that affect the hydration energy of ions. (04 scores)

7. A group of Senior Six learners from **Kampala Secondary School** visits a chemical manufacturing plant that produces solvents used in cleaning electrical equipment. One of the solvents produced at the plant is **tetrachloromethane (carbon tetrachloride, CCl_4)**.

During the visit, the plant chemist explains that understanding the **energy changes during bond formation** helps chemists design efficient production processes and minimise energy costs. The chemist further explains that chemists often use a **Born–Haber cycle** to analyse the different energy steps involved when atoms combine to form a compound.

The following thermochemical data were provided to the learners:

- Standard enthalpy of formation of carbon tetrachloride = **–97.5 kJ mol^{-1}**

- Standard enthalpy of atomisation of carbon = **718 kJ mol^{-1}**
- Standard enthalpy of atomisation of chlorine = **122 kJ mol^{-1}**

Task: As one of the learners in the group;

(i) Construct a **Born–Haber cycle** that illustrates the steps involved in the formation of tetrachloromethane from its elements. Clearly indicate the **energy changes for each step** in the cycle. (02 scores)

(ii) Use the data provided above and your Born–Haber cycle to calculate the **average bond energy of the carbon-to-chlorine (C–Cl) bond** in tetrachloromethane. (02 scores)

8. At a petroleum research laboratory in **Jinja**, the chemists are studying hydrocarbons obtained from crude oil refining. Some of these hydrocarbons are used as **raw materials in the production of synthetic rubber and plastics**.

One compound being studied is **but-2-ene**, which undergoes several chemical reactions during industrial processing. To understand its reactivity, the chemists analyse the **bond energies** within the molecule.

The following information is available:

- Standard enthalpy of formation of **but-2-ene** = **–211 kJ mol^{-1}**
- Atomisation energy of carbon = **718 kJ mol^{-1}**
- Bond energies:
  - C–C = **336 kJ mol^{-1}**
  - C–H = **412 kJ mol^{-1}**
  - H–H = **436 kJ mol^{-1}**

Task: As a chemistry learner;

(i) Using the information provided, help the research team **calculate the bond dissociation energy of the carbon-to-carbon double bond (C=C)** in but-2-ene. (03 scores)

(ii) Another hydrocarbon of interest in the laboratory is **buta-1,3-diene**, which is widely used in the manufacture of synthetic rubber. Using appropriate bond energy considerations, determine the **enthalpy of formation of buta-1,3-diene**. (02 scores)

9. **(a) Energy changes in fuel production.**

At a **biofuel research laboratory in Uganda**, a group of chemistry learners are studying the energy changes involved in the formation and combustion of organic fuels such as **propan-2-ol**, which can be used as an alternative fuel in small-scale energy systems.

During the investigation, the learners are reminded that the enthalpy change of a reaction can be determined indirectly using known enthalpy values of related reactions.

The following thermochemical data were obtained during the study:

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

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

$C_3H_7OH(l) + 4\frac{1}{2}O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$ $\Delta H = -2010$ kJ mol^{-1}

Task: As a chemistry student;

(i) State **Hess' law of constant heat summation**.

(ii) Using the information above, **calculate the standard enthalpy of formation of propan-2-ol.**

(b) Thermochemistry of explosives.

In a **forensic chemistry laboratory**, scientists are analyzing **nitroglycerine ($C_3H_5N_3O_9$)**, a compound used in explosives and medical vasodilators. The researchers are interested in determining the energy changes involved in its formation and decomposition.

They collected thermochemical data for the substances involved in the reactions.

$^1/_2N_2(g) + O_{2(g)} \longrightarrow NO_{2(g)};$ $\Delta H_f = +34$ kJmol^{-1}

$C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)};$ $\Delta H_f = -394$ kJmol^{-1}

$H_{2(g)} + \frac{1}{2}O_{2(g)} \longrightarrow H_2O_{(g)}; \quad \Delta H_f = -286\ kJmol^{-1}$

$C_3H_5N_3O_{9(l)} + {}^{11}/_4\ O_{2(g)} \longrightarrow 3CO_{2(g)} + {}^{5}/_2\ H_2O_{(l)} + 3NO_{2(g)}; \quad \Delta H = -1540\ kJmol^{-1}$

Task: As a chemistry learner;

(i) Using the provided thermochemical data, **calculate the standard enthalpy of formation of nitroglycerine**.

(ii) **Calculate the enthalpy change for the decomposition of nitroglycerine** into its products.

(iii) Under normal environmental conditions, nitroglycerine tends to **decompose rather than combust**.

Suggest a reason why **decomposition occurs more frequently than combustion**.

(c) Industrial production of hydrofluoric acid.

At a **chemical processing plant**, hydrofluoric acid is manufactured using the reaction between **calcium fluoride (fluorspar)** and **concentrated sulphuric acid**. The plant manager instructs technicians to grind the calcium fluoride before adding it to the acid.

The following enthalpies of formation are known:

Substance	**ΔHf (kJ mol⁻¹)**
$CaF_2(s)$	−1214
$H_2SO_4(l)$	−815
$CaSO_4(s)$	−1435
$HF(l)$	−269

Task: As a chemistry student;

(i) Explain why **calcium fluoride is ground into a fine powder** before mixing it with sulphuric acid.

(ii) Write the **balanced equation** for the reaction that occurs.

(iii) Calculate the **enthalpy change for the reaction**.

(iv) Suggest whether **warming the reaction mixture** would enhance or retard the reaction and explain your answer.

(d) Lattice energy of magnesium oxide.

During a **materials science study**, learners are investigating why **magnesium oxide (MgO)** is a very stable compound and has a very high melting point. They use thermochemical data to determine its lattice energy using the **Born-Haber cycle**.

The following data were provided:

$Mg(s) + \frac{1}{2}O_2(g) \rightarrow MgO(s) \qquad \Delta H = -642\ kJ\ mol^{-1}$

$Mg(s) \rightarrow Mg(g) \qquad \Delta H = 148\ kJ\ mol^{-1}$

$Mg(g) \rightarrow Mg^+(g) + e^- \qquad \Delta H = 738\ kJ\ mol^{-1}$

$Mg^+(g) \rightarrow Mg^{2+}(g) + e^- \qquad \Delta H = 1452\ kJ\ mol^{-1}$

$\frac{1}{2}O_2(g) \rightarrow O(g) \qquad \Delta H = 152\ kJ\ mol^{-1}$

$O^-(g) + e^- \rightarrow O^{2-}(g) \qquad \Delta H = 798\ kJ\ mol^{-1}$

Task: As a chemistry student;

(i) Represent the information above using an **energy level diagram (Born–Haber cycle)**.

(ii) Explain why the reaction $O^-(g) + e^- \rightarrow O^{2-}(g)$ has a **very large positive enthalpy value**.

(iii) Calculate the **lattice energy of magnesium oxide**.

(iv) Based on your answer in (iii), predict whether **magnesium oxide is likely to be soluble in water**.

(v) State **two factors that affect the magnitude of lattice energy** and explain how each factor influences it.

(e) Hydration enthalpy of copper(II) sulphate.

At a **school chemistry laboratory**, learners carried out an experiment to determine the **enthalpy of hydration of copper(II) sulphate** using an indirect calorimetric method. The following results were obtained:

- When **5.0 g of anhydrous $CuSO_4$** was dissolved in **75 g of water**, the temperature **rose by 4.7°C**.
- When **5.0 g of hydrated $CuSO_4 \cdot 5H_2O$** was dissolved in **75 g of water**, the temperature **fell by 3.9°C**.

Task: As a chemistry learner;

(i) Explain why **anhydrous copper(II) sulphate dissolves exothermically**, whereas **hydrated copper(II) sulphate dissolves endothermically**.

(ii) Calculate the **enthalpy of hydration of anhydrous copper(II) sulphate to form $CuSO_4 \cdot 5H_2O$** and hence determine its **lattice energy**.

(iii) Suggest **possible sources of error** in this calorimetry experiment.

(iv) State **two factors that influence the dissolution of ionic solutes in water**.

(f) Bond energies in organic reactions.

At a **chemical technology workshop**, students are studying the use of **bond energies** to estimate enthalpy changes in organic reactions. They are given the following **average bond energies** (kJ mol^{-1}):

Bond	**Energy** (kJ mol^{-1})
C–H	463
C≡C	837
C=O	745
O–H	463
C–C	348

The students investigate the hydration of **ethyne** to form **ethanal** under suitable catalytic conditions.

Task: As a chemistry student;

(i) Explain why **bond energy values are usually given as average values**.

(ii) Using the bond energies provided, **calculate the enthalpy change for the reaction between ethyne and water to form ethanal**.

(iii) State **known applications of bond energy terms in chemistry**.

10. Students at a secondary school in Uganda are participating in a **chemistry energy investigation project** aimed at understanding how fuels release energy during combustion. The teacher asks them to explore how the burning of different substances can heat water and how chemists measure the heat changes involved in reactions.

During the lesson, the class discusses important thermochemistry concepts and later carries out an experiment using a **bomb calorimeter** to determine the energy released when **naphthalene ($C_{10}H_8$)** burns. The results obtained from the experiment are later used to analyze temperature changes and calculate the heat released.

Tasks:

(a) During the introductory discussion, the teacher asks learners to explain key thermochemical terms used in energy studies.

(i) Explain what is meant by **standard enthalpy of combustion**. *(01 score)*

(ii) Explain what is meant by **standard enthalpy of formation**. *(01 score)*

(b) In the school laboratory, the students are provided with a **bomb calorimeter**, a known mass of **naphthalene**, and a known quantity of water. They are asked to determine the heat released when naphthalene burns.

Describe how the students could carry out the experiment to determine the **enthalpy of combustion of naphthalene ($C_{10}H_8$)** using a **bomb calorimeter**. *(A diagram is not required.)* *(05 scores)*

(c) During the experiment, **64 g of naphthalene** is completely burned. The heat produced is absorbed by **2.5 kg of water** in the calorimeter. The students measure and record the temperature of the water at regular time intervals as shown below.

Time (min)	**0**	**1.0**	**2.0**	**3.0**	**4.0**	**4.5**	**5.0**	**5.5**	**6.0**	**6.5**	**7.0**	**7.5**	**8.0**	**8.5**
Temperature (°C)	25.0	25.0	25.0	65.5	68.0	69.0	68.0	65.5	64.0	62.5	60.5	59.0	57.5	56.0

(i) Using the data above, plot a **graph of temperature against time**. *(04 scores)*

(ii) Using the graph obtained, determine the **enthalpy of combustion of naphthalene**. *(04 scores)*

(d) Later in the lesson, the teacher introduces students to the use of **Born–Haber cycles** to determine enthalpy changes in ionic compounds.

(i) Using appropriate energy steps, construct a **Born–Haber cycle** for the **standard enthalpy of formation of phosphorus pentachloride (PCl_5)**. *(02 scores)*

(ii) Using the information below, calculate the **standard enthalpy of formation of phosphorus pentachloride**.

- Standard enthalpy of atomisation of phosphorus = **+315 kJ mol^{-1}**
- Standard enthalpy of atomisation of chlorine = **+121 kJ mol^{-1}**
- Average bond energy of **P–Cl** bond = **−276.6 kJ mol^{-1}** *(03 scores)*