# Chapter 2 Energy Changes in Physical and Chemical Processes

## Objectives

By the end of this topic, the learner should be able to:

- (a) Define endothermic and exothermic reactions using the ΔH notation.
- (b) Draw energy level and energy cycle diagrams.
- (c) Explain fusion and vaporisation as evidence of inter-particle forces.
- (d) Explain that energy changes in chemical reactions are due to bond breaking and bond formation.
- (e) Define and explain various types of heat changes.
- (f) Carry out some experiments to determine enthalpy changes for some reactions.
- (g) Write correct simple thermochemical equations.
- (h) State Hess's Law and carry out related calculations.
- (i) State and explain the factors that influence the choice of a fuel.
- (j) Explain the environmental effects of fuels.

# Energy

**Energy is the ability to do work.** SI Units: Joules(J). Smaller units used mostly in Chemistry: kilojoules (kJ).

Law of Conservation of Energy: **Energy can neither be created nor destroyed but can only be converted from one form to another.** For example, heat energy to kinetic energy. Many physical and chemical changes are accompanied by energy changes.

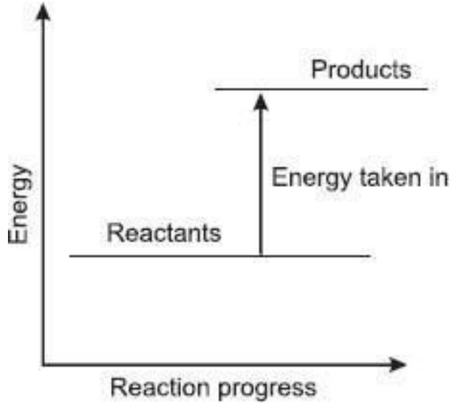
There are many forms of energy e.g. chemical, electrical, heat, kinetic and potential energy.

In this topic, we focus mainly on **heat energy** as whether evolved or taken in during reactions.

# Endothermic Reactions

A process that is accompanied by absorption of heat is called an **endothermic reaction/process.** In an endothermic reaction, the system developed by products and reactants absorbs heat energy from the surrounding. *The apparatus of reaction feels colder.*

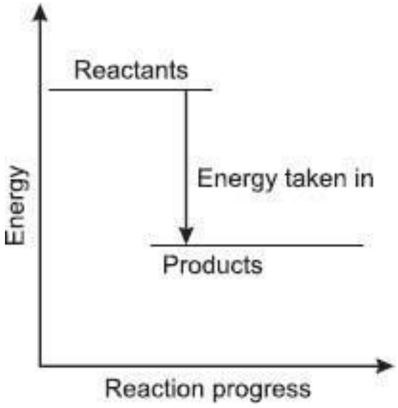
Therefore, the **products have a higher energy compared to the reactants.** This can be illustrated on an **Energy Level Diagram** as shown below:



**Energy level diagrams** are graphical illustrations which show the relative energies of the products and reactants in a reaction.

## Exothermic Reactions

A process that is accompanied by **production of heat**. Heat is released to the environment. *The apparatus of reaction feels warmer or hotter. Diagram:*



## Enthalpy

This is the heat content in a chemical denoted by H. It is not possible to measure this H since the internal energy contains components that are not known.

Thus, we need to react the chemical in order to measure the energy change it would undergo to achieve this since there would be a change in temperature.

The change in energy is branded **Enthalpy change** and is abbreviated $\Delta$ H given by $\Delta H = mc\Delta T$

$m - mass/volume\ of\ solution$

where $c - specific\ heat\ capacity\ of\ water = 4.2kJ/mol\ or\ 4200J/mol$

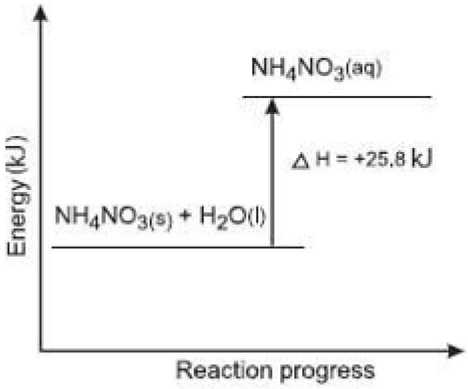
$\Delta T - change\ \text{in}\ temperature$

It can also be calculated using the relation:

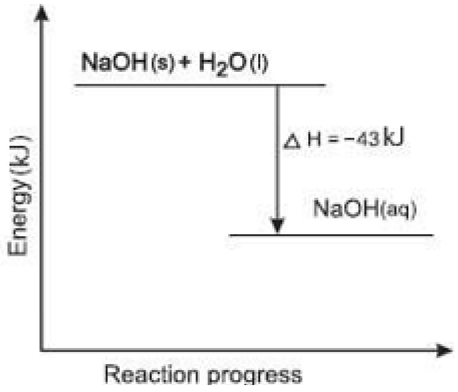
**Enthalpy change (ΔH) = Enthalpy of products – Enthalpy of reactants.**

**The Enthalpy Change must be shown on energy level diagrams.**

**Example for endothermic:**



**Example for Exothermic process**



The energy change associated with a chemical change is usually shown at the end of the balanced chemical equation. For example,

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)\ \Delta H = -728\ kJ\ mol^{-1}$$

$$NaNO_3(s) + (aq) \xrightarrow{H_2O} NaNO_3(aq)\quad \Delta H = +21\ kJ\ mol^{-1}$$

## Bond Breaking and Bond Formation in Physical and Chemical Processes

Eenergy transfer occur during chemical reactions because of the breaking and making of chemical bonds.

**Melting/freezing/fusion/solidification** and **boiling/vaporization/evaporation** are the two physical processes. Melting /freezing point of pure substances is fixed /constant. The boiling point of pure substance depend on **external** atmospheric **pressure**.

Melting/fusion is the physical change of a **solid** to **liquid**. Freezing is the physical change of a **liquid** to **solid**.

Melting/freezing/fusion/solidification are therefore two **opposite** but **same** reversible physical processes. i.e.

$$A(s) \rightleftarrows A(l)$$

Boiling/vaporization/evaporation is the physical change of a **liquid** to **gas/vapour**.

Condensation/liquification is the physical change of **gas/vapour** to **liquid**.

Boiling/vaporization/evaporation and condensation/liquification are therefore two **opposite** but **same** reversible physical processes. i.e.

$$B(l) \rightleftarrows B(g)$$

**Bond breaking is an endothermic process e.g. melting and boiling**

**Bond formation is exothermic process e.g. freezing and condensation**.

**Latent heat of fusion** is the amount of heat energy required to convert a given amount of a solid substance to a liquid at its melting point**.**

**Molar heat of fusion (Molar enthalpy of fusion)** is the amount of heat energy required to convert one mole of a solid substance at its melting point to a liquid.

**Latent heat of vaporisation** is the heat absorbed by a substance when changing from the liquid state to the gaseous state at constant temperature.

**Molar heat of vaporisation (Molar enthalpy of vaporisation)** is the amount of heat energy required to convert one mole of a liquid substance at its boiling point to its gaseous state.

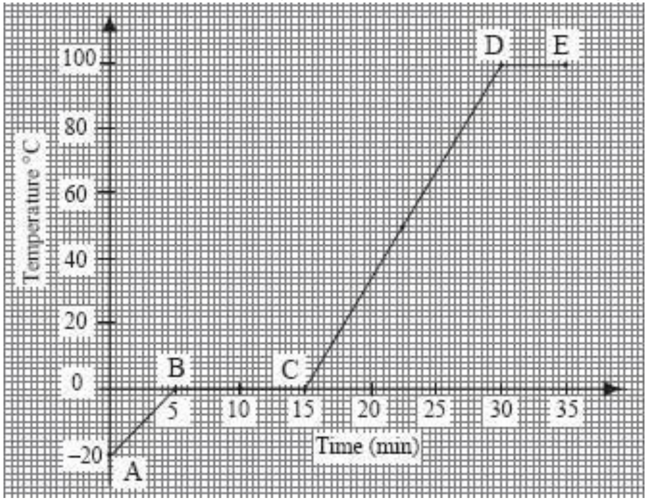


*Fig. 2.4: Heating curve of pure ice*

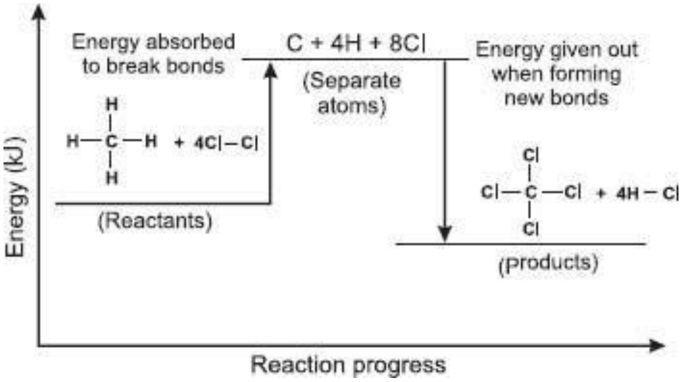
The molar heat of fusion and molar heat of vaporisation of a substance can be used to estimate the strength of the intermolecular forces holding the particles together in the solid and in the liquid states respectively. If the forces holding the particles together in the solid structure are strong, then the molar heat of fusion is high.

**molar heat of fusion and vaporisation for water and ethanol.**

| Compound | Molar heat of fusion (kJ mol⁻¹) | Molar heat of vaporisation (kJ mol⁻¹) | Melting point(k) |
|---|---|---|---|
| Water | 6.02 | 40.7 | 273 |
| Ethanol | 2.39 | 18.7 | 188 |

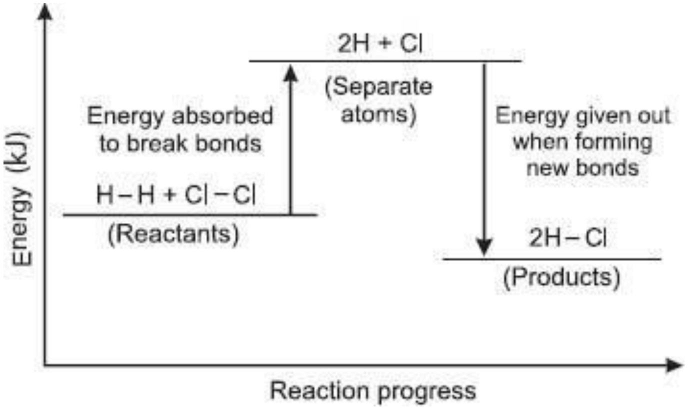
Methane reacts very slowly with chlorine in the dark. However, in the presence of light, a violent reaction occurs.

**For the reaction to occur, the covalent bonds in both methane and chlorine must be broken to obtain separate atoms. Energy is required to break these bonds. Thus, bond breaking is an endothermic process.** Once the bonds are broken, the atoms rearrange themselves and new bonds are formed in the products. In the process of bond formation, energy is given out. **Therefore, bond formation is an exothermic process.** The energy spent in breaking the bonds in the methane and chlorine molecules is less than the energy given out when the products are formed. Overall, the reaction therefore gives out energy. These changes can be illustrated using the energy level diagram shown in the figure below:



The Overall energy change is the sum of total energy for endothermic and total energy for exothermic processes.
Calculations:

The reaction between **hydrogen gas and chlorine gas to form hydrogen chloride** can be similarly illustrated as shown in figure below:



In some reactions, more energy is spent in breaking the bonds in the reactants than is given out when new bonds are formed in the products for such reactions, the overall process is endothermic.

If the bond strengths in reactants and products are known, then the enthalpy change for the reaction can be determined as follows:

### *Example*

Use the following bond energies to determine whether the reaction below is exothermic or endothermic:

$$H_2(g) + Cl_2(g) \longrightarrow 2HCl(g).$$

| **Bond** | **Energy in kJ mol$^{-1}$** |
|---|---|
| H–H | +436 |
| Cl–Cl | +244 |
| H–Cl | +432 |

**Working**

$$\text{H–H} + \text{Cl–Cl} \longrightarrow 2\text{H–Cl}$$

$$\underbrace{(346) + (244)}_{\text{Energy for bond breaking}} \longrightarrow \underbrace{2(432)}_{\text{Energy for bond making}}$$

Heat of reaction = Bond breaking energy + Bond formation energy

= 680 kJ + – 864 kJ

= –184 kJ

Since the overall energy is negative (–) the formation of hydrogen chloride (HCl) gas from hydrogen gas and chlorine gas is an **exothermic reaction**.

**Examples of common enthalpies**

Ammonium nitrate dissolves in water with the absorption of heat. This change is endothermic. When 2 g of ammonium nitrate is dissolved in 100 ml distilled water, a temperature change of about 1.5°C (1.5 K) is noted. The enthalpy change is calculated using the formula:

Enthalpy change = Mass of solution × Specific heat capacity × Temperature change

$\Delta H = MC\Delta T$ where M = Mass

C = Specific capacity

$\Delta T$ = Change in temperature in Kelvin

In this case, it is assumed that the mass of the solution is equal to that of a similar amount for water, i.e., 100 g (0.1 kg). The specific heat capacity of water is used in the calculation. Therefore:

$$\text{Heat absorbed} = 0.1\text{kg} \times 4.2\ \text{kJ Kg}^{-1}\text{K}^{-1} \times 1.5\ \text{K}$$

$$= 0.63\ \text{kJ}$$

The number of moles of ammonium nitrate is calculated as follows:

$$\begin{aligned} \text{Number of moles} &= \frac{\text{Mass}}{\text{Molar mass}} \\ &= \frac{2.0\ \text{g}}{80\ \text{g mol}^{-1}} \\ &= 0.025\ \text{mol} \end{aligned}$$

This means that when 0.025 mol/ of ammonium nitrate dissolves, 0.63 kJ of heat is absorbed. Therefore, the enthalpy changes when 1 mol of ammonium nitrate dissolves is:

$$\begin{aligned} &= \frac{1\ \text{mol} \times 0.63\ \text{kJ}}{0.025\ \text{mol}} \\ &= 25.2\ \text{kJ} \end{aligned}$$

Thus when 1 mol of ammonium nitrate dissolves 25.2 kJ of heat is absorbed. The equation for the reaction including the enthalpy change is:

$$NH_4NO_3(s) \overset{H_2O}{\rightarrow} NH_4NO_3(aq)\ \Delta H = +\ 25.2\ kJmol^{-1}$$

This value: + 25.22 kJ mol$^{-1}$ is called the **molar heat of solution** of ammonium nitrate.

The enthalpy change that occurs when one mole of a substance is dissolved in a solvent to give an infinitely dilute solution is called **molar heat of solution**. A mole of solute should be dissolved in a sufficient amount of solvent to ensure complete dissolution.

A chemical equation which shows the enthalpy change during reaction is called a **thermochemical equation**. A thermochemical equation is illustrated on an energy level diagram as shown below:

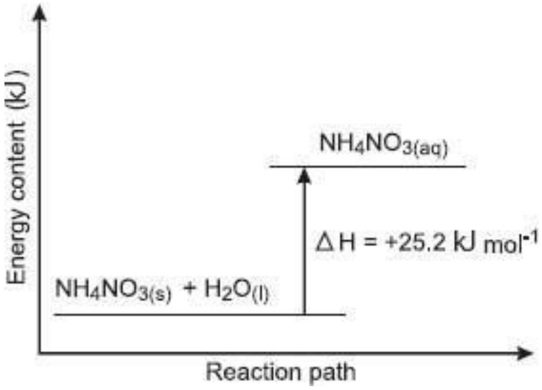


*Fig. 2.7: Energy level diagram for dissolving ammonium nitrate*

Sodium hydroxide dissolves in water with the evolution of heat. A temperature rise of about 5°C is recorded when 2 g of sodium hydroxide dissolves in 100 ml distilled water.

Heat evolved = Mass × Specific heat capacity × Temperature change

= 0.1 Kg × 4.2 kJ $Kg^{1} K^{1}$ × 5 K

= 2.1 kJ

Moles of sodium hydroxide used is given by:

$$\text{No. of moles} = \frac{\text{Mass}}{\text{Molar mass}}$$

$$= \frac{2.0\ \text{g}}{40\ \text{g mol}^{-1}}$$

$$= 0.05\ \text{mol}$$

Therefore, 0.05 mol liberates 2.1 kJ of heat.

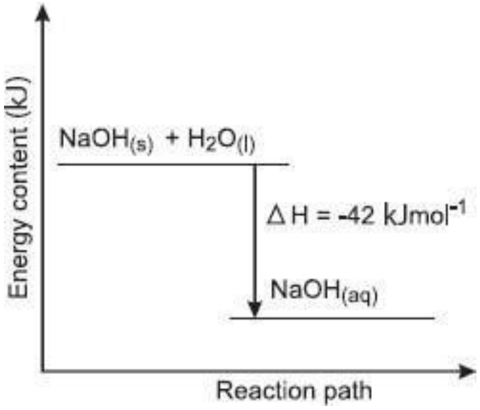
$$1\ \text{mol liberates:}\ \frac{1.0 \times 2.1}{0.05}\ \text{kJ of heat} = 42\ \text{kJ}$$

The thermochemical equation for the reaction is:

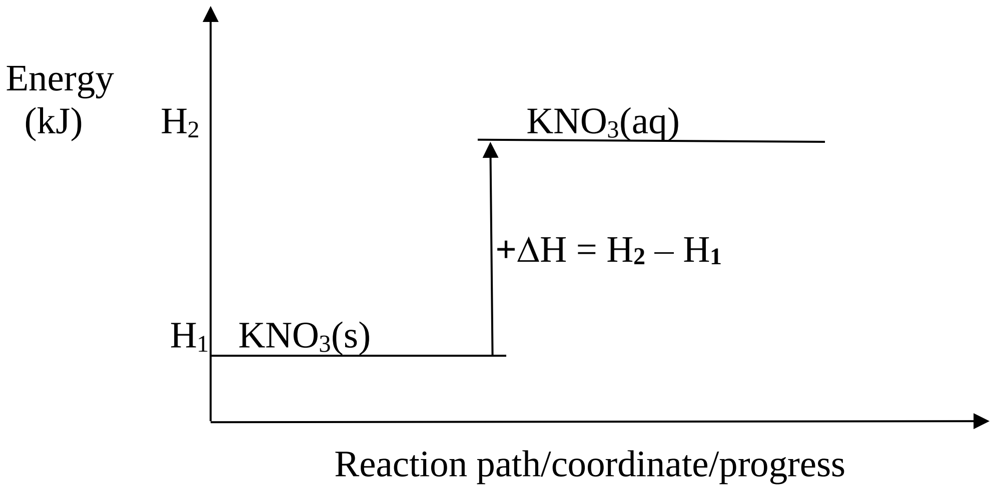
$$NaOH(s) \longrightarrow NaOH(aq)\ \Delta H = -42\ \text{kJ mol}^{-1}$$

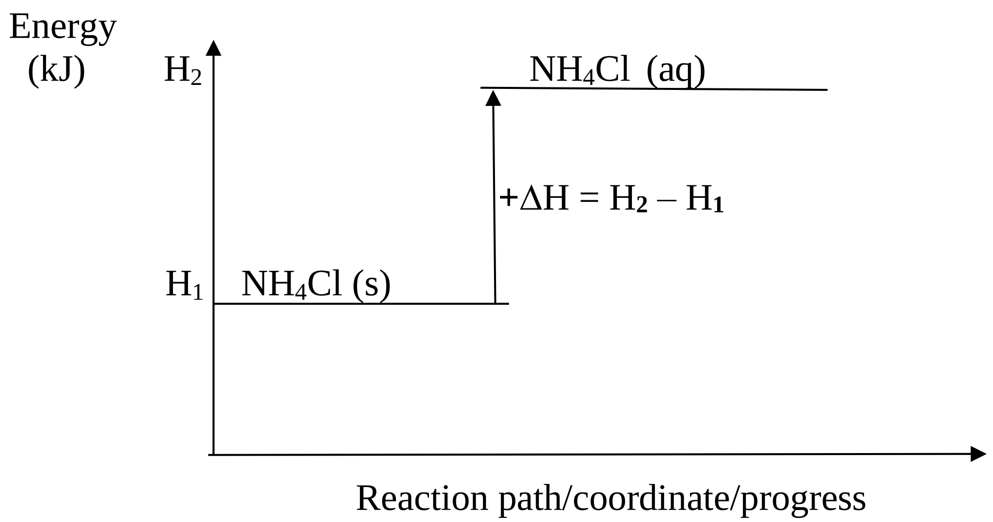
This (i.e., – 42 kJ $mol^{-1}$) is the molar heat of solution of sodium hydroxide.

The thermochemical equation can be illustrated on an energy level diagram.

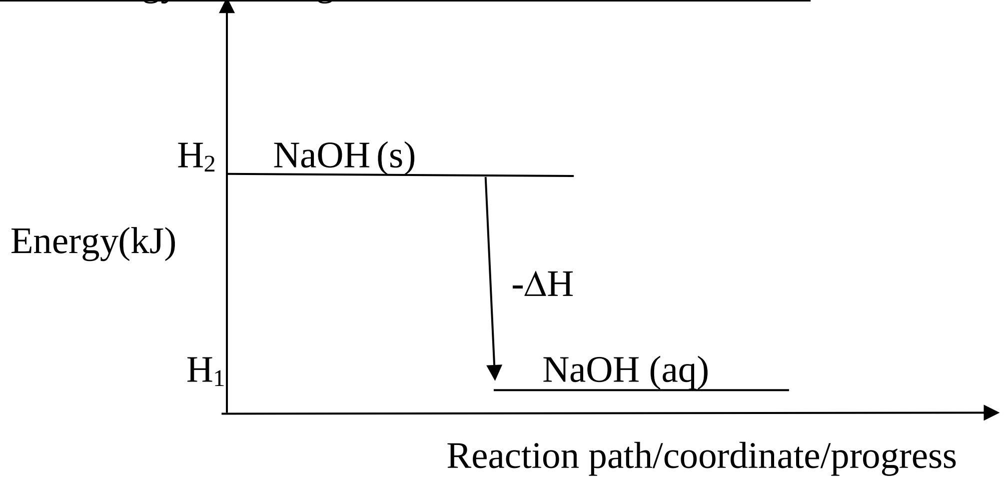


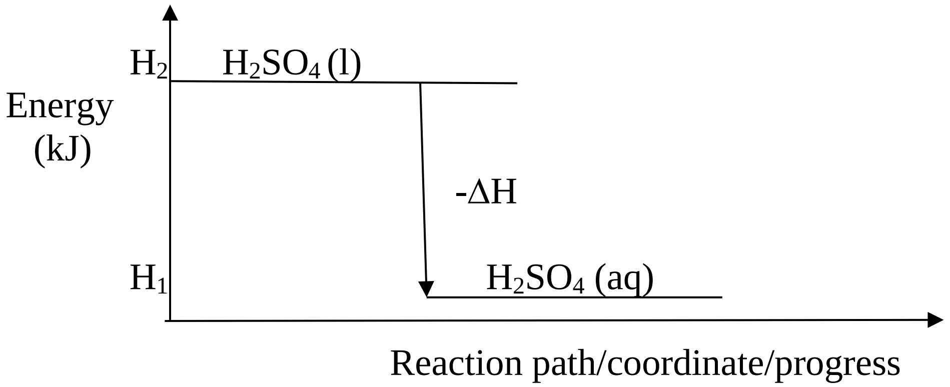
## Sketch energy level diagrams for endothermic dissolution





Sketch energy level diagrams for exothermic dissolution





### *Experiment 2.4: What is the heat of solution of concentrated sulphuric(VI) acid?*

Wrap a clean 250 ml plastic beaker with tissue paper. Secure the tissue paper with a rubber band. Measure exactly 98 $cm^3$ of distilled water into the beaker. Note the steady temperature of the water. Measure 2 $cm^3$ of concentrated sulphuric(VI) acid. Hold the beaker containing distilled water in a tilted position and carefully pour the 2 $cm^3$ of concentrated sulphuric(VI) acid into the beaker. See figure 2.9.

**Caution**:

Concentrated sulphuric(VI) acid should always be added to water and never vice

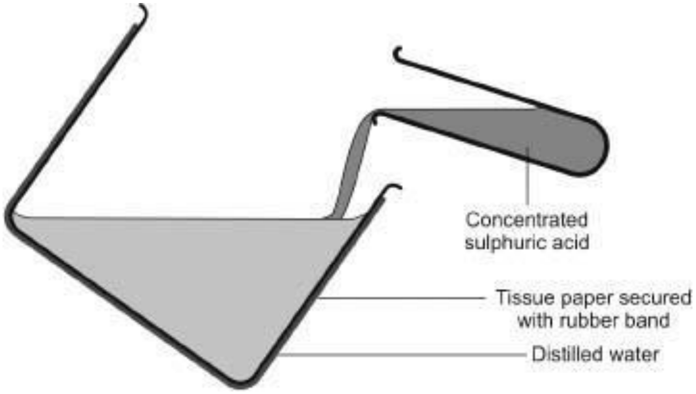


*Fig: 2.9 : Dissolving concentrated sulphuric (VI) acid in water*

Stir the mixture carefully using the thermometer and record the highest temperature of the solution. Record your results in table 2.5.

***Table 2.5: Dissolving concentrated sulphuric(VI) acid***

| | |
|---|---|
| Final temperature of solution (°C) | |
| Initial temperature of water (°C) | |
| Temperature change | |

## Answer the following questions

1. State whether the change is exothermic or endothermic.
2. Calculate the enthalpy change for the reaction.
3. Determine the mass of the concentrated acid used hence the moles of sulphuric(VI) acid. (Concentrated sulphuric(VI) acid is 98% pure and has a specific density of 1.84 (cm$^{-3}$).
4. Calculate the enthalpy change when one mole of concentrated sulphuric(VI) acid dissolves.
5. Write down the thermochemical equation for the reaction.

## Discussion

Concentrated sulphuric(VI) acid dissolves in water with the evolution of heat. When 2cm$^3$ of concentrated sulphuric(VI) acid is dissolved in 98 cm$^3$ of water, a temperature rise of 4.5°C is recorded.

The total volume of the solution = (2 + 98) cm$^3$ = 100 cm$^3$

The mass of solution is given by = Volume × Density

= 100 cm$^3$× 1.0 g cm$^{-3}$

= 100g (0.1g)

Heat evolved = Mass × Specific heat capacity × Temperature change

= 0.1 Kg × 4.2 kJ $Kg^{-1}$ $K^{-1}$ × 4.5 K = 1.89 kJ

The mass of the concentrated acid is calculated from the formula:

Mass = Specific density × Volume

= 1.84g $cm^{-3}$ × 2 $cm^3$ = 3.68 g

Thus the 2 $cm^3$ of concentrated sulphuric(VI) acid has a mass of 3.68 g. But since the acid is 98% pure, the actual mass of sulphuric(VI) acid is calculated by:

Mass of acid = Mass of conc. acid × percentage purity

$$= 3.68 \text{ g} \times \frac{98}{100} = 3.6064 \text{ g}$$

The moles of the acid is given by:

$$= \frac{3.6064 \text{ g}}{98 \text{ g mol}^{-1}}$$

$$= 0.0368 \text{ mol}$$

Therefore, when 0.0368 mol of sulphuric(VI) acid dissolves, 1.89 kJ of heat is evolved.

When 1 mol of the sulphuric(VI) acid dissolves, $\frac{1.0 \times 1.89 \text{ kJ}}{0.0368}$ of heat is evolved.

= 51.358 kJ of heat.

Thus the molar heat of solution of sulphuric(VI) acid

= –51.358 kJ $mol^{-1}$

The negative sign is added to show that the process is exothermic.

The thermochemical equation:

$$H_2SO_4 \text{ (l)} \xrightarrow{H_2O} H_2SO_4\text{(aq)} \qquad \Delta H = -51.358 \text{ kJ mol}^{-1}$$

The thermochemical equation is illustrated on an energy level diagram as shown.

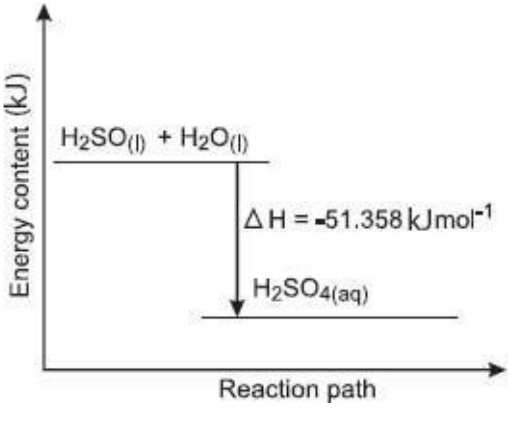


*Fig. 2.10: Energy level diagram for dissolving concentrated sulphuric(VI) acid.*

The experimental values of molar heat of solution obtained in experiments are usually lower than those quoted in data books, since the heat gained by the apparatus and that lost to the surrounding during an experiment is not accounted for in the calculations.

### *Experiment 2.5: What is the enthalpy of combustion of ethanol?*

Pour exactly 100 $cm^3$ of distilled water into a 250 ml glass beaker. Record the temperature of the water. Half fill a small bottle with ethanol and close the bottle with a lid fitted with a clean wick. Weigh the bottle and its contents and arrange the apparatus as shown in figure 2.7.

Light the wick to start heating the water. Stir the water carefully with the thermometer and extinguish the flame when the temperature of the water has risen by 15°C. Record the temperature of the water. Weigh the bottle and its contents again. Record the results obtained.

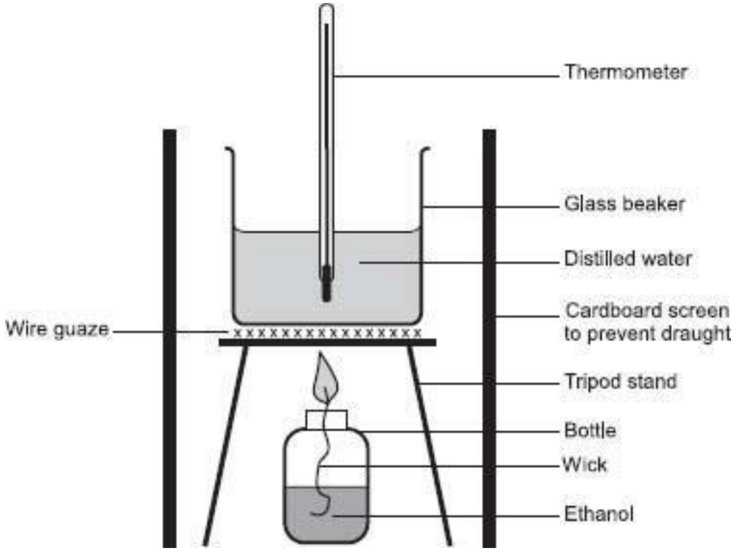


*Fig 2.11: Combustion of ethanol*

Initial mass of bottle with its contents $W_1$ =________________________g

Final mass of bottle with its contents, $W_2$ =________________________g

Mass of ethanol burned, $W_1 - W_2$ =____________________________g

Final temperature of water, $t_2$ =_____________________°C

Initial temperature of water, $t_1$ =_____________________________°C

Change in temperature $t_2$, – $t_1$ =______________________°C

### Answer the following questions

1. Determine the number of moles of the ethanol burned.
2. Determine the amount of heat given out.
3. Using your answers to questions 1 and 2, calculate the molar enthalpy of combustion of ethanol.

4. The theoretical enthalpy of combustion of ethanol is –1368 kJ $mol^{-1}$, why does the value calculated from experimental results differ from this?
5. Write the thermochemical equation for the reaction.

## Discussion

The following results were obtained in a similar experiment.

Initial mass of bottle + contents = 28.25 g

Final mass of bottle + contents = 28.02 g

Mass of ethanol burnt = 0.23 g

Final temperature of water = 43.0°C

Initial temperature of water = 28.0°C

Rise in temperature of water = 15°C (15 K)

Formula mass of ethanol, $CH_3CH_2OH$ = (12 × 2) + 16 + 6 = 46 g

$$\text{Moles of ethanol burned} = \frac{0.23\ \text{g}}{46\ \text{g mol}^{-1}}$$

$$= 0.005\ \text{mol}$$

The heat evolved by the burning ethanol is absorbed by the water. It is this heat that raises the temperature of water. Therefore, the heat evolved is determined thus:

Heat evolved = Mass of solution × Specific heat capacity × Temperature change

$= 0.1\ \text{kg} \times 4.2\ \text{kJ g}^{-1}\text{K}^{-1} \times 15\ \text{k} = 6.3\ \text{kJ}$

Therefore, 0.005 moles of ethanol liberate 6.3 kJ of heat.

Therefore 1 mol of ethanol will liberate:

$$\frac{6.3\ \text{kJ} \times 1.0\ \text{mol}}{0.005\ \text{mol}}$$

= 1260 Kg of heat

The thermochemical equation is:

$$CH_3CH_2OH(l) + 3O_2(g) \xrightarrow{\quad} 2CO_2(g) + 3H_2O(l)\ \Delta H = -1260\ \text{kJ mol}^{-1}$$

**Molar heat of combustion** *is the enthalpy change that occurs when one mole of a substance is completely burned in oxygen.*

The thermochemical equation is illustrated on an energy level diagram as shown.

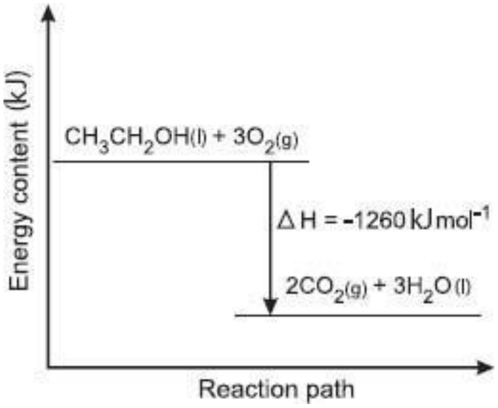


*Fig. 2.12: Energy level diagram for the combustion of ethanol*

The negative sign is added to show that the process is exothermic. Experimentally determined heats of combustion are usually lower than the theoretical values because the heat lost to the surrounding is not accounted for in the calculations (experimental error).

### *Experiment 2.6: What is the enthalpy change when zinc reacts with copper(II) sulphate solution?*

Wrap a 250 $cm^3$ plastic beaker with tissue paper. Measure 100 $cm^3$ of 0.5 M copper (II) sulphate into the beaker. Note the steady temperature of the solution. Carefully, transfer 4.0 g of zinc powder into the plastic beaker and stir carefully with the thermometer as shown in figure 2.8. Record the highest temperature attained by the solution. The experiment is repeated using magnesium or iron instead of zinc.

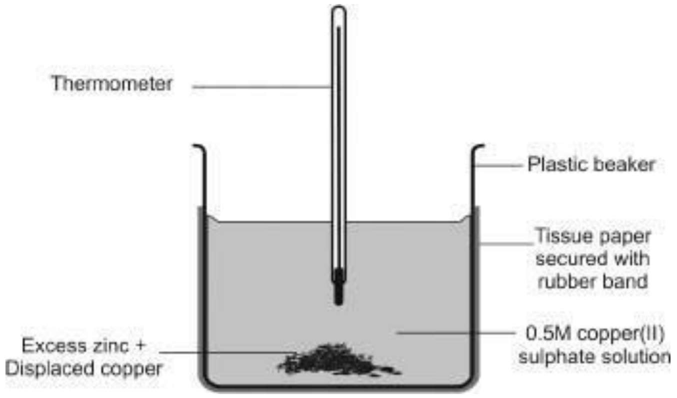


*Fig. 2.13 Displacing copper(II) ions with zinc.*

## Answer the following questions

1. State whether the reaction is exothermic or endothermic.
2. What other observations were made? Explain.

3. Write the equation for the reaction that takes place.
4. Calculate the number of moles of zinc used.
5. Calculate the number of moles of copper.
6. Why is it necessary to use excess zinc powder, magnesium or iron filings in this reaction?
7. Calculate the molar heat of displacement of copper(II) ions with zinc. Given that:

(a) Density of solution is 1 g $cm^{-3}$

(b) The volume of the solution remains unchanged after the reaction.

(c) The specific heat capacity of the solution is 4.2 kJ $g^{-1}$ $K^{-1}$ (Cu= 63.5, S=32.0, O =16.0, Zn = 65.0)

8. Comment on your results given that the molar heat of displacement of copper (II) ions and zinc is 216 kJ $mol^{-1}$

## Discussion

When zinc is added to copper(II) sulphate solution, an exothermic reaction occurs. The blue colour of the solution fades as a brown solid is deposited. During the reaction, the blue copper(II) ions in the solution are reduced to copper metal which is the brown solid deposited at the bottom of the beaker.

The zinc atoms are oxidised to zinc ions which are colourless according to the following equation:

$$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$$

The expected change in temperature for this experiment is 24.5°C (24.5 K).

$$\text{Moles of zinc used} = \frac{4\ g}{65\ g\ mol^{-1}}$$

$$= 0.062\ mol$$

$$\text{Moles of copper (II) ions} = \frac{0.5\ mol \times 100\ cm^3}{1000\ cm^3}$$

$$= 0.05\ mol$$

From the equation the reaction ratio between metal and copper(II) ions is 1:1.

Excess zinc is used in this experiment in order to ensure that all the copper(II) ions are changed to copper metal.

Heat evolved = Mass × Specific heat capacity × Temperature change

= 0.1 kg × 4.2 kJ $Kg^{1}$ $K^{1}$×24.5 K

= 10.29 kJ

Thus, when 0.05 mol of copper(II) ions are displaced from solution by zinc, 10.29 kJ of heat is evolved.

When 1 mol of Copper (II)ions are displaced, $\frac{10.29 \times 1}{0.05}$ = 205.8 kJ of heat is evolved.

The molar heat of displacement of copper by zinc therefore, is –205.8 kJ mol$^{-1}$

Thermochemical equation for the reaction is:

$$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+} + Cu(s)\ \Delta H = -205.8\ kJ\ mol^{-1}$$

**Molar heat of displacement** *is the enthalpy change that occurs when one mole of a subtance is diplaced from a solution of its ions.*

The experimental value obtained in this experiment is low because the heat lost to the surrounding and that absorbed by the apparatus is not accounted for in the calculation.

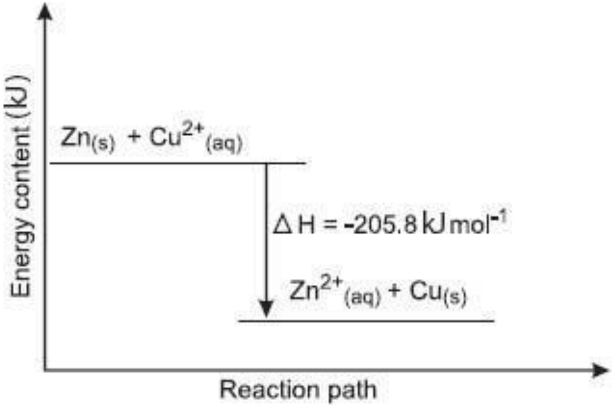


*Fig. 2.14 : Energy level diagram for displacement of copper(II) ions from solution*

The value obtained when magnesium is used is higher than that of zinc and iron, while that of zinc is higher than that of iron. The further apart the two metals are in the reactivity series, the higher the molar heat of displacement.

### *Experiment 2.7: What is the heat of neutralisation when hydrochloric acid reacts with sodium hydroxide*

Wrap a clean 250 ml plastic beaker with tissue paper. Secure the tissue paper with a rubber band. Transfer exactly 50 cm$^3$ of 2 M hydrochloric acid solution into the beaker. Note the steady temperature $T_1$ of the hydrochloric acid solution. Using a clean dry measuring cylinder, measure exactly 50 cm$^3$ of 2 M sodium hydroxide solution and note its steady temperature $T_2$. Add the 50 cm$^3$ of 2 M hydrochloric acid. Carefully, stir the contents of the beaker with the thermometer while adding the sodium hydroxide solution. Note the highest temperature $T_4$ attained by the resulting solution. Record the results.

Temperature of the acid $t_1$ =____°C

Temperature of sodium hydroxide $t_2$ =_____°C

Highest temperature of mixture $t_4$ =_____°C

Average temperature of the two solutions, $t_3 = \frac{t_1 + t_2}{2}$ =_°C

Temperature change, $\Delta T = t_4 - t_3 =$ ____°C

Repeat the experiment using:

1 M sulphuric(VI) acid with 2 M sodium hydroxide.

2 M ethanoic acid with 2 M sodium hydroxide.

2 M hydrochloric acid with 2 M ammonia solution.

**Answer the following questions**

1. For each experiment:

(a) Calculate the molar heat of neutralisation.

(b) Write the ionic equation. Include the enthalpy change.

### Discussion

Strong acids and strong alkalis dissociate completely in water into their ions. During neutralisation, hydrogen ions react with hydroxide ions to form water molecules. Heat energy is liberated in the process. A sample calculation on enthalpy is shown below.

Temperature of hydrochloric acid, $t_1 = 22.5°C$

Temperature of sodium hydroxide solution $t_2 = 22.5°C$

Average temperature of the acid and alkali,

$$t_3 = \frac{t_1 + t_2}{2} = \frac{22.5 + 22.5}{2} = 22.5\ °C$$

The highest temperature of the mixture $t_4$

$= 35.5\ °C$

Temperature change $\Delta T = t_4 - t_3$

$= 35.5 - 22.5$

$= 13.0\ °C$

The specific heat capacity of the solution is 4.2 kJ $Kg^{-1}\ K^{-1}$. If in the experiment, 50 $cm^3$ of 2 M hydrochloric acid are neutralised by 50 $cm^3$ of 2 M sodium hydroxide, then the final volume of the resulting solution

$= 50 + 50 = 100\ cm^3$

If the density of the resulting solution is taken to be 1 $g/cm^3$ then:

The mass of the solution = 100 × 1 = 100 g

Heat evolved $= MC\Delta T$

$$= \frac{100 \times 4200\ J \times 13.0}{1000}$$

$= 5460\ J = 5.46\ kJ$

Mol of hydrochloric acid used $= \frac{2 \times 50}{1000} = 0.1$ mol. Therefore,

mol of $H^+$ ions = 0.1 mol since one mole of HCl produces one mole of $H^+$ ions.

Mol of sodium hydroxide used $= \frac{2 \times 50}{1000} = 0.1$mol

Mol of $OH^-$ ions = 0.1 mol.

The equation of the reaction is

$$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$$

Thus 1 mol of $H^+$ reacts with 1 mol of $OH^-$ to form 1 mol of water.

In the reaction between the acid and the base, 0.1 mol of water are formed.

When 0.1 mol of water are formed 5.46 kJ of heat energy is evolved. Therefore, when 1.0 mol of water is formed

$$= \frac{5.46 \times 1}{0.1}$$

= 54.6 kJ of heat energy is evolved.

This is the molar heat of neutralisation of hydrochloric acid by sodium hydroxide.

The enthalpy change can be incorporated in the neutralisation equation as follows:

$$H^+(aq) + OH^-(aq) \longrightarrow H_2O\ (l)\ \Delta H_{neut} = -\ 54.6\ kJ\ mol^{-1}$$

The enthalpy change that occurs when an acid and a base react to produce one mole of water is called **molar heat of neutralisation**.

Strong acids and strong bases react to liberate about 57.2 kJ when they form one mol of water during neutralisation.

When one of the reactants or both are weak, the enthalpy of neutralisation is less than in the case when strong acids are used with strong bases, for example, when one mole of hydrochloric acid (strong acid) is neutralised by one mole of ammonium hydroxide (weak base), about 51.4 kJ of heat is evolved for every one mole of water formed.

$$HCl(aq) + NH_4OH(aq) \longrightarrow NH_4Cl(aq) + H_2O(l)\ \Delta H = -\ 51.4\ kJ\ mol^{-1}$$

Similarly if ethanoic acid (weak acid) is neutralised by sodium hydroxide (strong base) only 55.2 kJ of heat is produced for every mole of water formed.

$$C_2H_4CO_2(aq) + NaOH(aq) \longrightarrow C_2H_3COONa(aq) + H_2O(l)\ \Delta H = -\ 55.2\ kJ\ mol^{-1}$$

Weak acids such as ethanoic acid or weak bases such as ammonia solution, are only partially ionised in solution. Therefore, some energy is used up to ionise them before the neutralisation takes place. The enthalpy change obtained when a weak acid reacts with a weak base is even lower. For example, when ethanoic acid reacts with ammonia solution, only 50.4 kJ of heat is evolved.

$$CH_3COOH(aq) + NH_4OH(aq) \longrightarrow CH_3COONH_4(aq) + H_2O(l)\Delta H = -\ 50.4\ kJ\ mol^{-1}$$

## Bond dissociation energies of some (covalent)bonds

| Bond | Bond dissociation energy (kJmole$^{-1}$) | Bond | Bond dissociation energy (kJmole$^{-1}$) |
|---|---|---|---|
| H-H | 431 | I-I | 151 |
| C-C | 436 | C-H | 413 |
| C=C | 612 | O-H | 463 |
| C ≡ C | 836 | C-O | 358 |
| N ≡ N | 945 | H-Cl | 428 |

| | | | |
|---|---|---|---|
| N-H | 391 | H-Br | 366 |
| F-F | 158 | C-Cl | 346 |
| Cl-Cl | 239 | C-Br | 276 |
| Br-Br | 193 | C-I | 338 |
| H-I | 299 | O=O | 497 |
| Si-Si | 226 | C-F | 494 |

## *Practice Problems*

1. Calculate ΔH if a piece of metal with a specific heat of 0.98 $kJkg^{-1}K^{-1}$ and a mass of 2 kg is heated from 22$^{o}$C to 28$^{o}$C.
2. If a calorimeter's ΔH is +2001 Joules, how much heat did the substance inside the cup lose?
3. Calculate the ΔH of the following reaction: $CO_{2\,(g)} + H_2O_{\,(g)} \text{-->} H_2CO_{3\,(g)}$ if the standard values of $\Delta H_f$ are as follows: $CO_{2\,(g)}$: -393.509 KJ /mol, $H_2O_{\,(g)}$ : -241.83 KJ/mol, and $H_2CO_{3(g)}$ : -275.2 KJ/mol.
4. Calculate ΔH if a piece of aluminum with a specific heat of .9 kJ·kg−1·K−1 and a mass of 1.6 kg is heated from 286$^{o}$K to 299$^{o}$K.
5. If the calculated value of ΔH is positive, does that correspond to an endothermic reaction or an exothermic reaction?

## *Solutions*

1. $\Delta H = q = cp_{sp}$ x m x (ΔT) = (.98) x (2) x (+6$^{o}$) = 11.76 kJ
2. Since the heat gained by the calorimeter is equal to the heat lost by the system, then the substance inside must have lost the negative of +2001 J, which is -2001 J.
3. $\Delta H^{\circ} = \sum \Delta v_p \Delta H^{\circ}_f$(products) - $\sum \Delta\ v_r \Delta H^{\circ}_f$(reactants) so this means that you add up the sum of the ΔH's of the products and subtract away the ΔH of the products: (-275.2kJ) - (-393.509kJ + -241.83kJ) = (-275.2) - (-635.339) = +360.139 kJ.
4. $\Delta H = q = cp_{sp}$ x m x (ΔT) = (.9) x (1.6) x (13) = 18.72 kJ
5. Endothermic, since a positive value indicates that the system GAINED heat.

### Standard Conditions for Measuring Enthalpy Changes

The standard conditions are:

i) temperature of 25°C (298 K)

ii) pressure of one atmosphere (101.325 kPa)

Values of enthalpy changes that are measured at these conditions are called **standard enthalpy changes** and given a special symbol; $\Delta H^{\theta}$. The symbol ‘denotes ‘standard’ and it implies a pressure of 101.325 kPa (i.e. one atmospheric pressure) and a temperature of 298 K (i.e., 25°C). A subscript is also added to the symbol to indicate the type of enthalpy change involved. Thus:

- $\Delta H^{\theta}_{c}$ refers to the standard molar enthalpy change of combustion, e.g.,
  $CH_4(g) + 2O_2\ O \longrightarrow CO_2(g) + 2H_2(l)\ \Delta H^{\theta}C_{(CH_4)} = -890\ kJ\ mol^{-1}$
- $\Delta H^{\theta}_{neut}$ Refers to the standard molar enthalpy change of neutralisaiton e.g.
  $HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l)\ \Delta H^{\theta}_{neut} = -58\ kJ\ mol^{-1}$
- $\Delta H^{\theta}_{soln}$ refers to the standard molar enthalpy change of dissolution (solution)
  e.g., $NaNO_3\,(s) + aq \longrightarrow NaNO_3\,(aq)\ \Delta H^{\theta}_{Soln(NaNO_3)} = +\ 21\ kJ\ mol^{-1}$

The value of ΔH in a thermochemical equation refers to molar quantities shown in the equation. For example, the following equations show the standard molar enthalpy change of formation of water.

$$HCl\ (aq) + NaOH\ (aq) \longrightarrow NaCl\ (aq) + H_2O\ (l)$$

$$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)\ \Delta H = -\ 57.2\ kJ\ mol^{-1}$$

Note that only one mole of $H_2O$ is shown in the above equation.

$$H_2SO_4(aq) + NaOH(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$$

The equation for the formation of two moles of would be:

$$2H^+(aq) + 2OH^-(aq) \longrightarrow 2H_2O(l)\ \Delta H = -\ 114.4\ kJ^{-1}$$

The amount of heat evolved when two moles are formed is therefore twice what is evolved when one mole is formed. Note that the units for $\Delta H^\theta$ is the kilo joule per mole ($kJ\ mol^{-1}$).

## Hess's Law

The enthalpy change for some reactions cannot be determined experimentally because the reactions cannot take place under normal conditions. For example, it is not possible to determine the enthalpy of formation of methane ($CH_4$) in the ordinary laboratory because it is not possible to synthesise methane from carbon and hydrogen.

Where direct synthesis of a compound is not possible, the alternative is to determine the enthalpy of formation theoretically using other measurable enthalpies. One method of doing this involves measuring the enthalpy of combustion of the compound and the enthalpies of combustion of its constituent elements. These values are then linked in an energy cycle diagram with the enthalpy change of formation of the compound. Figure 2.15 shows how this can be done for methane.

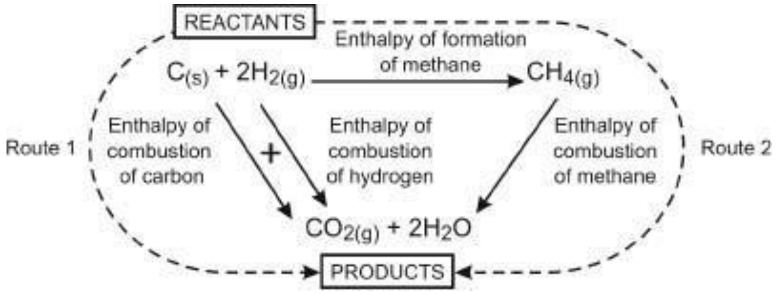


*Fig 2.15 : Energy cycle diagram incorporating enthalpy of formation of methane*

A close look at figure 2.15 reveals that there is more than one way of converting carbon and hydrogen into carbon(IV) oxide and water respectively. Carbon and hydrogen can either be burned directly (route 1) or they can first be combined to form methane which can then be burned (route 2).

Whether route 1 or route 2 is followed, the end product is the same. This means that the energy changes for route 1 are equal to the energy changes for route 2. This must be so because of the law of conservation of energy. From fig. 2.15 are enthalpy of formation of methane can then be calculated from the expression.

$$\boxed{\begin{array}{l}\text{Enthalpy of}\\ \text{formation of}\\ \text{methane}\end{array}} = \boxed{\begin{array}{l}\text{Enthalpy of}\\ \text{combustion of}\\ \text{carbon}\end{array}} + \boxed{\begin{array}{l}\text{Enthalpy of}\\ \text{combustion of}\\ \text{hydrogen} \times 2\end{array}} - \boxed{\begin{array}{l}\text{Enthalpy of}\\ \text{combustion of}\\ \text{methane}\end{array}}$$

$= -393 + (-286 \times 2) - (-890)$

$= -965 + 890$

$= -75 \text{ kJ mol}^{-1}$

This method of indirectly determining the enthalpy of formation of a compound is based on **Hess's Law of constant heat summation**: *The energy changes in converting reactants to products is the same regardless of the route by which the chemical change occurs.*

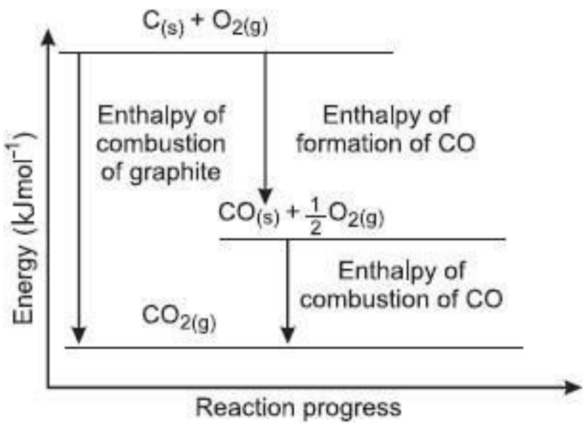


*Fig 2.16: Energy level diagram for formation of carbon(II) oxide*

From the energy cycle and energy level diagram, the enthalpy of formation of carbon(I1) oxide would be given by:

Enthalpy = Enthalpy of combustion of C(s) – Enthalpy of combustion of CO(g)

$\Delta H^{\theta}_{f}\,CO(g) = \Delta H^{\theta}_{c}\,C(s) - \Delta H^{\theta}_{c}\,CO(g)$

$= -393 - -283$

$= -170$ kJ mol$_{-1}$

Application of Hess's law relies on the information provided by thermochemical equations for the reaction in question.

Information provided by the thermochemical equations can be used to draw energy cycle and energy level diagrams linking the **equations** for a reaction as well as determine the enthalpy change for the **reaction**.

## Worked Examples

1. The thermochemical equation for the combustion of carbon (graphite) and carbon(II) oxide are as follows:

$C(s) + O_2(g) \longrightarrow CO_2(g)\ \Delta H = -393$ kJ mol$^{-1}$

$CO(g) + \frac{1}{2} O_2(g) \longrightarrow CO(g)\ \Delta H = -283$ kJ mol$^{-1}$

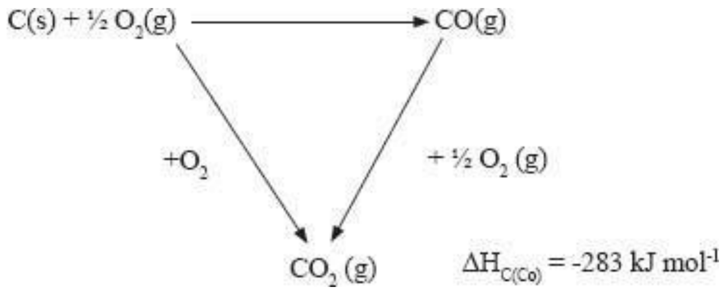
(a) Use the information provided to draw:

(i) An energy cycle diagram.

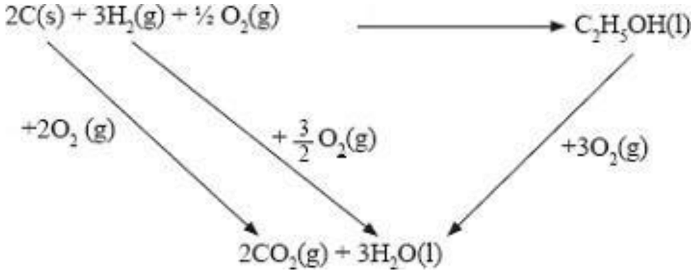
(ii) An energy level diagram for the reaction.

(b) Use the energy cycle and energy level diagrams drawn in (a) to determine the enthalpy of formation of carbon(II) oxide.

These equations can be linked in an energy cycle diagram as follows:



The energy cycle diagram for the formation and combustion of ethanol is :



From the energy level diagram:

$$\text{Enthalpy of formation of ethanol} = 2\times\text{combustion of carbon} + 3\times\text{combustion of hydrogen} - \text{combustion of ethanol}$$

$\Delta H^{\theta}_{f}(C_2H_5OH(l)) = 2\Delta H^{\theta}_{c}(C(s) + 3\Delta H^{\theta}_{c}(H_2(s) – \Delta H_{c}(C_2H_5OH(l)$

$= 2 \times (–393) + 3 \times (–286)–(–1368)$

$= –786 + (–858) – (–1368)$

$= –1644 + 1368$

$= – 276$ kJ mol$^{-1}$

## Relationship Between Heat of Solution, Hydration and Lattice Energy

In experiment 2.1, it was found that dissolving substances in water can be accompanied by the absorption or evolution of heat. When an ionic compound is formed from its constituent gaseous ions, energy is given out. The energy is known as **lattice energy**. It is defined as *the energy change when one mole of an ionic compound is formed from its constituent ions in the gaseous state.*

Dissolving an ionic solid involves first, the interaction of the water molecules with ions in the solid to such an extent that the solid dissociates into monoatomic gaseous ions which scatter in all directions in water. Therefore, energy equivalent to the lattice energy is absorbed. This is the opposite of the process through which the ionic solid was formed. For example, when sodium chloride is dissolved in water, the ionic bonds are first broken to release the gaseous ions.

$$NaCl(s) \xrightarrow{water} Na^+(g) + Cl^-(g) \qquad \Delta H^\circ = +776 \text{ kJ mol}^{-1}$$

The sodium and chloride, ions get surrounded by several water molecules. The water molecules surrounding the sodium ions arrange themselves in such a way that the partially negative oxygen is attracted to the positive sodium ion and the partially positive hydrogen is attracted to the negative chloride ions as shown in figure 2.10.

The ions thus become hydrated.

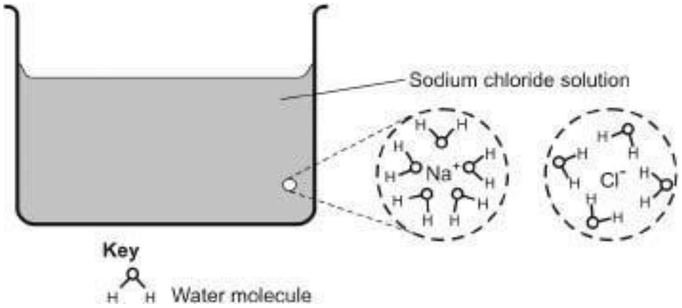


*Fig. 2.17 Hydration of sodium chloride ions*

The hydration process involves the formation of new bonds between the polar water molecules and the ions. The process is always exothermic. The energy given out is known as **hydration energy**.

$$Na^+(g) \xrightarrow{H_2O} N^+(aq) \qquad \Delta H^\theta = -\,390\ kJ\ mol^{-1}$$

$$Cl^-(g) \xrightarrow{H_2O} Cl(aq) \qquad \Delta H^\theta = -384\ kJ\ mol^{-1}$$

**Hydration** *energy is defined as the energy change that occurs when one mole of gaseous ions become hydrated.*

The overall enthalpy change of solution depends on whether the endothermic or the exothermic process is larger. For sodium chloride, the endothermic process is greater than the exothermic process. Therefore, the enthalpy of solution of sodium chloride has a positive value. The following energy cycle diagram can be used to calculate the heat of solution of sodium chloride.

$$\begin{tikzcd}
NaCl(s) \arrow[rr, "\text{Lattice energy}", "(\Delta H_{latt})"'] \arrow[dr, "+\text{ water}", "\text{Heat of solution } (\Delta H_{soln})"'] & & Na^+(g) + Cl^-(g) \arrow[dl, "+\text{ water}"', "\text{Hydration energy } (\Delta H_{hyd})"] \\
& Na^+(aq) + Cl^-(aq) &
\end{tikzcd}$$

From the energy cycle diagram,

Heat of solution = Lattice energy + hydration energy

$$\Delta H_{sol} = \Delta H = \Delta H_{hyd}$$

$$= 781 + (-774)$$

$$= +7 \text{ kJ mol}^{-1}$$

*Figure 2.17 illustrates the same changes using an energy level diagram*

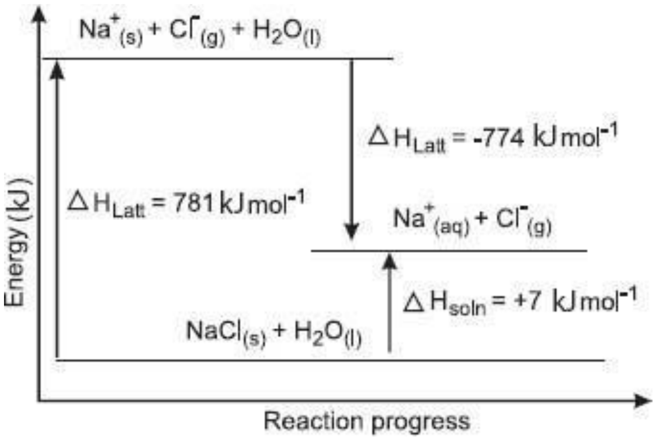


*Fig. 2.17: Energy level diagram for dissolving sodium chloride*

Note that the lattice energy is given positive values in the calculations above, this is because bond breaking is an endothermic process. Note also that the hydration energy for sodium chloride is the sum of the separate hydration energies of sodium ions and chloride ions, i.e.,

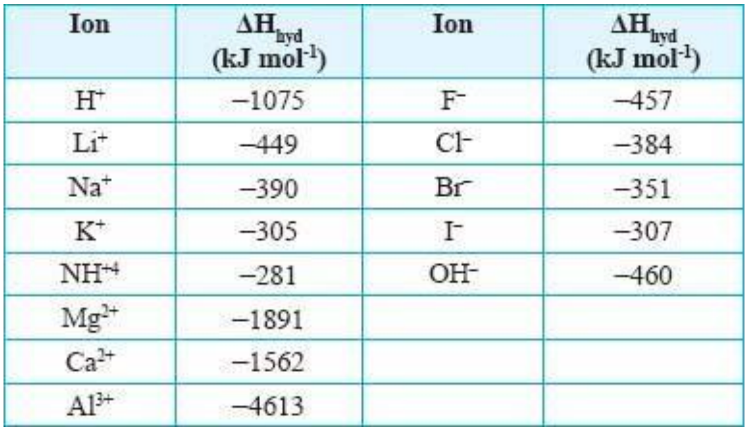
$$\Delta H_{hyd} \, NaCl = \Delta H_{hyd} \, (Na^+) + \Delta H_{hyd} \, (Cl^-)$$

$$= -390 + (-384)$$

$$= -774 \text{ kJ}$$

When ions are hydrated, the amount of energy released depends on the size of the ions and the charge on the ions. Table 2.6 shows the enthalpies of hydration of some ions while table 2.7 shows lattice energy of some ionic solids.

***Table 2.6: Enthalpies of hydration of some common ions***



| Ion | $\Delta H_{hyd}$ (kJ $mol^{-1}$) | Ion | $\Delta H_{hyd}$ (kJ $mol^{-1}$) |
|---|---|---|---|
| $H^+$ | −1075 | $F^-$ | −457 |
| $Li^+$ | −449 | $Cl^-$ | −384 |
| $Na^+$ | −390 | $Br^-$ | −351 |
| $K^+$ | −305 | $I^-$ | −307 |
| $NH^{+4}$ | −281 | $OH^-$ | −460 |
| $Mg^{2+}$ | −1891 | | |
| $Ca^{2+}$ | −1562 | | |
| $Al^{3+}$ | −4613 | | |

*Table 2.7: Lattice energies of some ionic common compounds*

| Compound | ΔH kJ $mol^{-1}$ |
|---|---|
| NaF | –915 |
| NaCl | –781 |
| Na Br | –743 |
| NaI | –699 |
| $MgCl_2$ | –2489 |
| MgO | –3933 |
| AgCl | –890 |

## Exercise

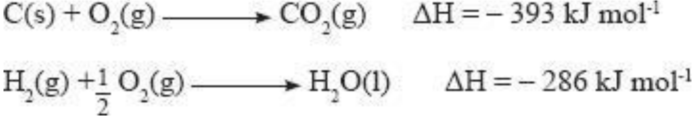
1. (a) Draw an energy cycle diagram for dissolving magnesium chloride in water.

(b) Use the information given in table 2.6 and 2.7 to calculate the enthalpy of solution of magnesium chloride.

(c) Use your answer to part (b) to draw the energy level diagram for dissolving magnesium chloride.

2. Given that:

$C(s) + O_2(g) \longrightarrow CO_2(g)$ $\Delta H = -393$ kJ $mol^{-1}$



$C(s) + O_2(g) \longrightarrow CO_2(g) \quad \Delta H = -393 \text{ kJ mol}^{-1}$

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

Calculate the molar enthalpy of formation of ethyne from its constituent elements.

## Fuels

**A fuel** *is a substance that produces useful energy when it undergoes a chemical or nuclear reaction.* The energy produced in such a reaction can be used either directly as heat, or can be converted into other forms of energy, e.g., electrical energy. Fuels can be solids such as nuclear fuel, coke, coal, charcoals and wood. Other fuels are liquids such as petrol, kerosene and diesel oil, or gases like natural gas, biogas, water gas and petroleum gas.

Heats of combustion are specially important because they provide one way of comparing the heat energy produced when equal quantities of different fuels are burned. Generally not all the heat energy liberated during combustion is utilised. For example, when coal is burned in a power station, only about 50% of the energy liberated is converted into electricity. The other 50% is converted into other forms of energy.

## Heating Values of Fuels

Heating value has the units kJ $g^{-1}$ (kilojoules per gram). It is obtained by dividing the molar enthalpy of combustion by the formula mass of the fuel. For example, ethanol has an enthalpy of combustion of -1360 kJ/mol. On dividing this by the formula mass of ethanol, 46, we get the heating value of ethanol as 30 kJ $g^{-1}$.

The heating value of a fuel is therefore the amount of heat energy given out when a unit mass, or a unit volume of a fuel is completely burned in oxygen. The heating values of some other fuels are given in the table 2.8.

*Table 2.8: Heating values of various fuels*

| Fuels | Heating value in J/g |
|---|---|
| **Solid Fuels** | |
| Charcoal | 33 |
| Coal | 29 |
| Wood | 17 |
| **Liquid fuels** | |
| Ethanol | 30 |
| Fuel oil | 45 |
| Paraffin | 48 |
| **Gaseous Fuels** | |
| Methane (natural gas) | 55 |
| Propane | 50 |
| Butane | 48 |

## How to choose a fuel

The choice of a fuel depends on the purpose it is to be used for. The following points have to be considered in each case:

- Heating value.
- Ease and rate of combustion.
- Availability.
- Ease of transportation.
- Ease of storage.

- Environmental effects.
- Cost.

For example, wood and charcoal are chosen for domestic heating because:

- They are cheap.
- The are readily available.
- They can be easily transported.
- When they burn they do not produce poisonous products.
- They burn slowly.

Methylhydrazine ($CH_3NHNH_3$) is used for rocket propulsion because:

- It burns very rapidly producing large amounts of gases which in turn create a huge thrust as they escape.
- It has a very high heat of combustion (4740 kJ $mol^{-1}$).
- It ignites easily.

## Precautions Necessary When Using Fuels

Improper handling of fuels can result it death and destruction of property. Each type of fuel should be handled with the necessary care and precautions:

- Charcoal stoves should be operated in well ventilated rooms to avoid poisoning by carbon(II) oxide.
- Vehicle engines should not be left running in closed garages to avoid poisoning by carbon(II) oxide.
- Gas cylinders should be stored in well ventilated rooms far from heat sources. Only certified cylinders should be used for gas transportation.
- Fuel storage facilities should be located far away from populated areas.
- People should keep off from fuel spilled from tankers.

## Environmental Effects of Fuels

Fossil fuels such as coal and petroleum contain carbon, nitrogen and sulphur compounds which on burning produce poisonous gases such as sulphur(IV) oxide, sulphur(VI) oxide, carbon(II) oxide and nitrogen(IV) oxide.

The sulphur and nitrogen oxides dissolve in rain-water to produce acid rain. Acid rain:

- Wears limestone buildings and statues.
- Corrodes iron sheets, iron gates and other metallic structures.
- Acidifies lakes leading to death of plants and animals in the lakes.
- Leaches nutrients from plant leaves leading to their death.
- Leaches minerals from the soil leading to poor soil for agricultural activities.

Carbon(IV) oxide and unburned hydrocarbons from fuels contribute to global warming. Global warming causes polar ice to melt causing inundation of low lying coastal lands. Other effects of global warming include unpredictable weather patterns and climatic changes.

*This zero emission car can give 100 miles on a single battery charge*

Measures that are being taken to reduce pollution:

- Designing zero emission vehicles which use solar energy or electrical energy.
- Fitting catalytic converters to vehicle exhaust systems. These catalytic converters convert nitrogen oxides to harmless nitrogen, carbon(II) oxide to carbon(IV) oxides, unburned hydrocarbons to water and carbon(IV) oxide.
- Designing engines which uses unleaded petrol.
- Electronically controlling the quantity of air mixed with fuel to ensure more complete combustion.
- Adding fuels which contain oxygen in their molecules to petrol. For example, methanol and ethanol may be added to petrol. These fuels reduce the quantities of carbon(II) oxide and unburned hydrocarbons emitted.
- Encourage other means of transport such as the use of bicycle and electric trains.

## Project Work

1. (a) Find out the heating values of common fuels available in your area.

(b) Find out about the environmental effects of burning each of the fuels in (a) above.

(c) From your answers in (a) and (b), determine the best choice of fuels for use in your area.

## Summary

1. Chemical reactions involve changes in energy.
2. Reactions which absorb energy from the surrounding are called **endothermic reactions**.
3. Reactions which release energy to the surrounding are called **exothermic reactions**.
4. Enthalpy change is given the symbol ΔH. Endothermic reactions have positive values for ΔH while exothermic reactions have negative values for ΔH.

5. Chemical bonds must first be broken during a reaction so that new, different bonds can be formed. Bond breaking is endothermic while bond formation is exothermic.
6. The amount of heat produced or absorbed during a chemical reaction is the difference between the energy content of the reactants and the products.
7. The standard conditions under which enthalpy changes are measured are 25°C (298 K) and one atmospheric pressure (101.325 kPa) (STP).
8. Hess's Law states that the energy in a chemical change is independent of the route by which the chemical change occurs.

Enthalpy change of reaction = Enthalpies of products – Enthalpies of reactants.

9. The heat of a reaction, $H^{\theta}_{rxn}$ is the amount of heat absorbed or evolved, when molar quantities of reactants in the equation react to give products under standard conditions of temperature and pressure.
10. A fuel is a substance that produces useful energy when it undergoes a chemical or nuclear reaction.

## Revision Exercise

1. In an experiment to compare the heats of combustion of methanol, ($CH_3OH$) and ethanol, ($CH_3CH_2OH$) it was established that the heat evolved during combustion of methanol is about 970 kJ $mol^{-1}$ while that for ethanol is about 1250 kJ $mol^{-1}$.

Write equations, including the heat changes, to show the complete combustion of methanol and ethanol.

2. When 8.0 g of sulphur is completely burned in oxygen in a calorimeter, the heat evolved raises the temperature of 500 $cm^3$ of water by 35°C. Calculate the heat of combustion of sulphur. (Specific heat capacity of water is 4.2 kJ/kg $K^{-1}$, S = 32).
3. Sulphur(IV) oxide reacts completely with one mole of oxygen to form sulphur(VI) oxide and liberating 185 kilojoules of energy.

(a) Write the equation for the reaction, showing the heat change.

(b) What is the heat of combustion when one mole of sulphur(IV) oxide is converted to sulphur(VI) oxide in excess oxygen?

(c) What would be the heat evolved if only 0.2 moles of sulphur(VI) oxide were formed?

4. (a) Define standard heat of formation of a substance.

(b) Ethanol, $CH_3CH_2OH$, cannot be prepared directly from its elements and so its standard heat of formation must be obtained indirectly.

  (i) Write an equation for the formation of ethanol from its constituent elements in their standard states.

  (ii) Draw an energy cycle diagram linking the heats of formation and combustion.

5. Use the table of enthalpies of hydration and lattice energies in the appendix to answer the questions that follow:

(a) Why are the hydration energies of both anions and cations negative?

(b) Calculate the hydration energy for magnesium bromide and aluminium chloride.

(c) Sodium, magnesium and aluminium belong to the same period in the periodic table. Explain the trend in hydration energy values of the cations formed from the atoms.

(d) Fluorine, chlorine, bromine and iodine belong to the same group in the periodic table. Explain the trend in hydration energy values of anions formed from their atoms.

(e) Calculate the enthalpies of solution of:

(i) Potassium iodide.

(ii) Lithium chloride.

(f) Calculate the lattice energy of:

(i) Aluminium chloride.

(ii) Sodium chloride.

(iii) Calcium chloride.

6. Some of the fuels commonly used at home are charcoal, wood kerosene, natural gas and biogas. These burn in oxygen to give out heat energy. Comment on the suitability of each of these substances as fuels.

7. (a) State Hess's Law.

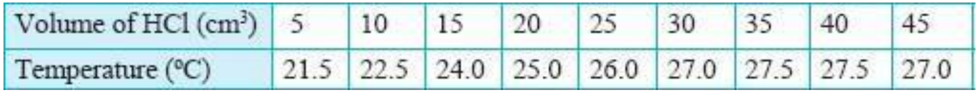
(b) Calculate the heat of formation of propane from the following data:

$$C(s) + O_2(g) \longrightarrow CO_2(g)\ \Delta H = -393\ kJ\ mol^{-1}$$

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

$$C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(l)\ \Delta H = -2209\ kJ\ mol^{-1}$$

8. In an experiment, 50 cm$^3$ of 1.0 M sodium hydroxide solution was placed in a suitable apparatus and 5.0 cm$^3$ portions of hydrochloric acid were added. The resulting mixture was stirred with a thermometer and the temperature taken after each addition. Both solutions were initially at 20°C.



| Volume of HCl (cm$^3$) | 5 | 10 | 15 | 20 | 25 | 30 | 35 | 40 | 45 |
|---|---|---|---|---|---|---|---|---|---|
| Temperature (°C) | 21.5 | 22.5 | 24.0 | 25.0 | 26.0 | 27.0 | 27.5 | 27.5 | 27.0 |

(a) (i) Plot a graph of temperature against volume of the acid added.

(ii) Use the graph to determine the concentration in moles per litre of the hydrochloric acid.

(iii) Calculate the enthalpy of neutralisation of the alkali with hydrochloric acid. (Specific heat capacity of mixture = 4.2 kJ kg K$^{-1}$ and density of solution is 1.0 g/cm$^3$)

(b) Draw an energy level diagram for the reaction.

(c) Suggest possible source of errors.

9. The molar enthalpy of neutralisation of sulphuric acid with sodium hydroxide is higher than that of ethanoic acid. Explain.

10. Given that:

$C(s) + O_2(g) \longrightarrow CO_2(g)$ $\Delta H = -393\ kJ\ mol^{-1}$
(Graphite)

$C(s) + O_2 \longrightarrow CO_2(g)$ $\Delta H = -395\ kJ\ mol^{-1}$
(Diamond)

Determine $\Delta H$ for $C(s) \longrightarrow C(s)$
(Graphite) Diamond)

11. When 50 $cm^3$ of 2 M of hydrochloric acid was neutralised by an equal volume of 2 M sodium hydroxide solution, both initially at 20 °C, the temperature of the mixture rose to a maximum of 33.7°C. What is the molar enthalpy of neutralisation of sodium hydroxide? Specific heat capacity of each solution is 4.2 $J\ g^{-1}\ K^{-1}$

12. Given that the thermochemical equations for combustion of ethanol, hydrogen and graphite (carbon) are:

$C_2H_5OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(l)\ \Delta H = -1368\ kJ\ mol^{-1}$
$H_2(g) + O_2(g) \longrightarrow H_2O(l)\ \Delta H = -286\ kJ\ mol^{-1}$
$C(s) + O_2(g) \longrightarrow CO_2(g)\ \Delta H = -393\ kJ\ mol^{-1}$

(a) Draw an energy cycle diagram linking the enthalpy of formation of ethanol with its enthalpy of combustion and the enthalpy of combustion of its constituent elements.

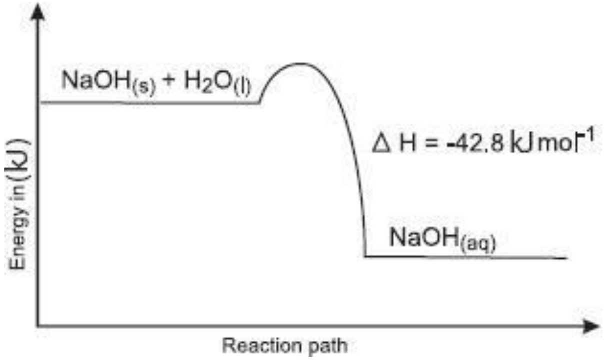
(b) Calculate the enthalpy of formation of ethanol.

13. When 2.65 grams of ammonium nitrate is dissolved in 100 $cm^3$ of water, the temperature of the mixture dropped by 2°C.

(a) Calculate the molar heat of solution of ammonium nitrate (density of solution = 1 g $cm^3$, c= 4.2 kJ, N $^{+}$14, H=1, O =16).

(b) Write a thermochemical equation for the dissolution of ammonium nitrate.

The diagram below shows the dissolution of Sodium hydroxide in water.



(a) (i) State and explain whether the reaction is exothermic or endothermic.

(ii) Given that the molar mass of sodium hydroxide is 40 grams. Calculate the enthalpy change when 10 grams of NaOH is dissolved in water.

14. (i) Define molar heat of combustion of a substance.

(ii) 1.26 grams of element Q was completely burned in air. The heat evolved was used to heat 250 cm$^3$ of water, rising the temperature from 32 to 50°C. Calculate the relative atomic mass of element Q given that molar enthalpy of combustion is –360 kJ mol$^{-1}$ (density of water is 1 g cm$^{-3}$ and specific heat capacity of water = 4.2 kJ $^{-1}$K$^{-1}$)

15. When 6.5 g of Zinc powder are added to 250 cm$^3$ of 0.1 M copper(I) sulphate solution in a plastic container, 5.45 kJ of heat is evolved.

(a) Explain why a plastic container is used instead of a metallic one.

(b) Write an ionic equation for the reaction between zinc powder and copper(II) sulphate solution.

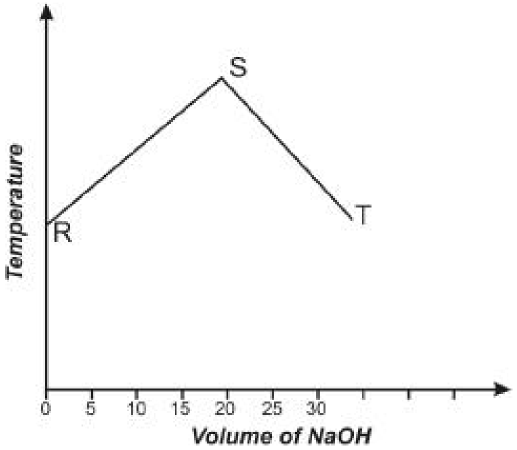
16. What is meant by the term enthalpy of neutralisation?

When 50 cm$^3$ of 1 M $H_2SO_4$ was added to 50 cm$^3$ of 2 M NaOH, the temperature of the resultant mixture rose by 13.6°C

(i) Write an equation for the reaction that occurred.

(ii) Calculate the enthalpy of neutralisation of sodium hydroxide (specific heat capacity of solution = 4.2 Jg$^{-1}$ K$^{-1}$, density of water = 1g cm$^{-3}$).

17. The graph below shows the change in temperature when 0.5 M NaOH solution was added to 20 cm$^3$ of HCl.

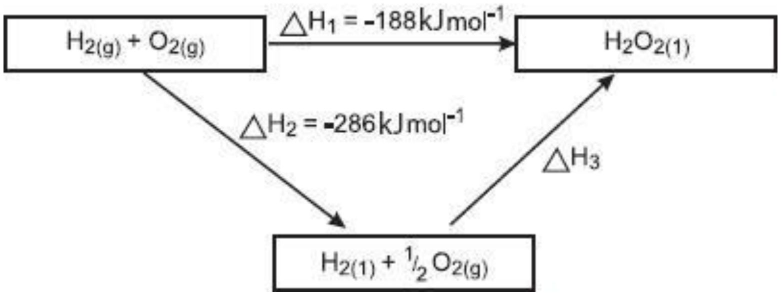


(i) What does point S represent?

(ii) Why does the temperature rise from R to S?

(iii) Determine the molarity of the HCl acid.

18. The diagram below shows an energy cycle.



(a) Give the name of $H^2$

(b) Determine the value of $H^3$

KCSE

## Energy changes in chemical and physical processes

1. 6g of Potassium nitrate solid was added to 120cm$^3$ of water in a plastic beaker. The mixture was stirred gently and the following results were obtained.

Initial temperature = 21.5°C
Final temperature = 17.0 °C

(a) Calculate the enthalpy change for the reaction
(Density =1g/cm$^3$, C= 4.2jg$^{-1}$K$^{-1}$)

(b) Calculate the molar enthalpy change for the dissolution of potassium nitrate
(K=39, N= 14, O =16)

2. (a) The heat of combustion of ethanol, $C_2H_5OH$ is 1370KJ/mole.
(i) What is meant by heat of combustion?
(ii) Calculate the heating value of ethanol
(H = 1.0, C = 12.0, O = 16.0)

3. Use the information below to answer the questions that follow:-

$Ca_{(s)} + ½ O_{2(g)} \longrightarrow CaO_{(s)}$ ΔH =-635KJ/mol
$C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)}$ ΔH= -394KJ/mol
$Ca_{(s)} + C_{(s)} + {}^3/_2O_{2(g)} \longrightarrow CaCO_3$ ΔH = -1207KJ/mol

Calculate the enthalpy change for the reaction:
$Ca_{(s)} + CO_{2(g)} \longrightarrow CaCO_{3(s)}$

4. 0.92g of ethanol were found to burn in excess air producing a temperature rise of 32.5ºC in 200cm$^3$ of water.

C=12.0 H=1.0 O=16.0
Density of water 1g/cm$^3$
Specific heat capacity of water 4.2kj kg$^{-1}$k$^{-1}$

a) Write the equation for combustion of ethanol
b) Determine the molar heat of combustion of ethanol

5. Study the information in the following table and answer the questions that follow. The letters do not represent the actual chemical symbols of the elements.

| **ELEMENT** | **U** | **V** | **W** | **X** | **Y** | **Z** |
|---|---|---|---|---|---|---|

| | | | | | | |
|---|---|---|---|---|---|---|
| NUMBER OF PROTONS | 18 | 20 | 6 | 16 | 19 | 17 |
| NUMBER OF NEUTRONS | 22 | 20 | 8 | 16 | 20 | 20 |

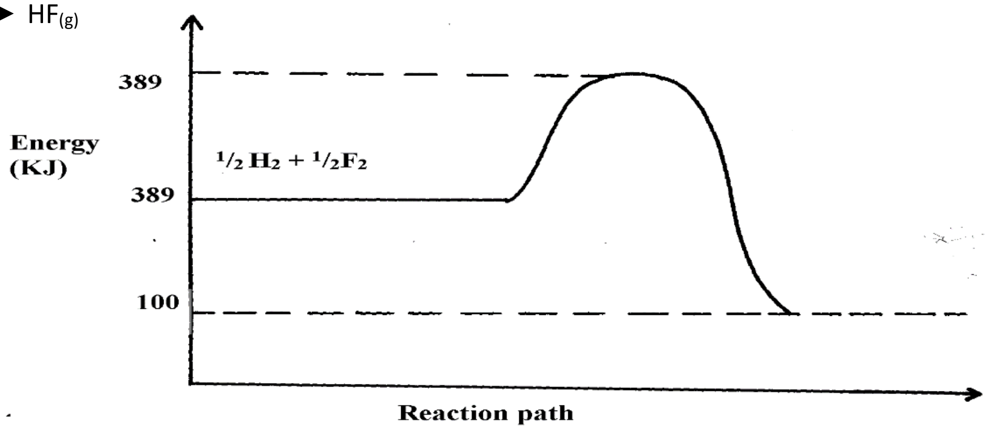
Which of the above elements are:

(i) Likely to be radioactive?

(ii) Able to form a compound with the highest ionic character?

6. The diagram below shows energy levels for the reaction

$\frac{1}{2} H_{2(g)} + \frac{1}{2} F_{2(g)} \longrightarrow HF_{(g)}$



(a) Work out the activation energy for the reaction

(b) Calculate the heat of formation of **HF**

(c) Is the reaction endothermic or exothermic?

7. Using the heats of combustion of the following substances, calculate the heat of formation of ethanol

$C_{(s)} + O_{2\,(g)} \longrightarrow CO_{2\,(g)}; \Delta H = -393KJmol^{-1}$

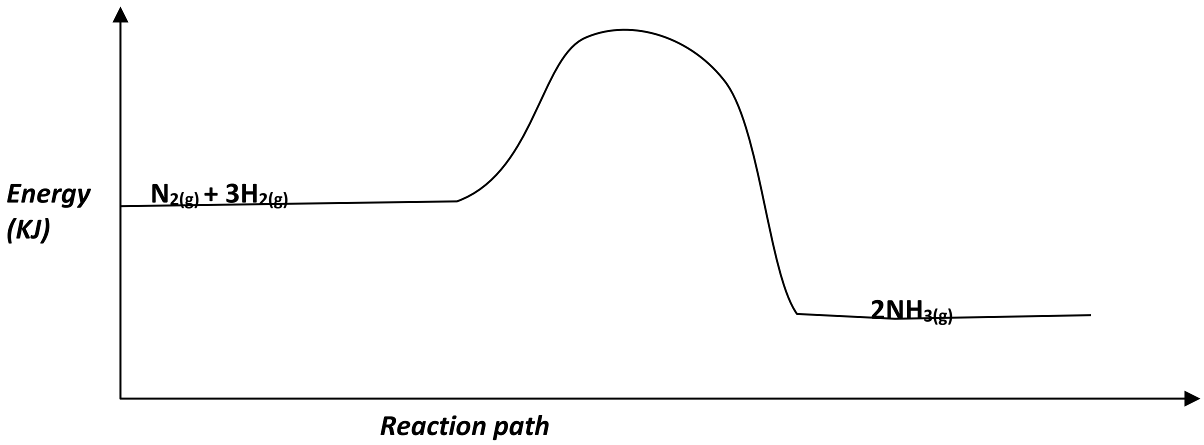
$H_{2\,(g)} + \frac{1}{2} O_{2\,(g)} \longrightarrow H_2O_{(l)}; \Delta H = -286KJmol^{-1}$

$CH_3CH_2OH_{(l)} + O_{2\,(g)} \longrightarrow 2CO_{2\,(g)} + 3H_2O_{(l)}; \Delta H = 1386KJmol^{-1}$

8. Nitrogen and hydrogen react reversibly according to the equation:-

$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}; \Delta H = -92kjmol^{-1}$

The energy level diagram for the above reaction is shown below:-



(a) How would the yield of ammonia be affected by:

(i) A decrease in temperature

(ii) An increase in pressure

(b) How does a catalyst affect reversible reaction already in equilibrium?
(c) On the above diagram, sketch the energy level diagram that would be obtained when iron catalyst is added to the reaction

9. Study the table below and answer the questions that follow

| **Bond type** | **bond energy kJmol$^{-1}$** |
|---|---|
| C-C | 346 |
| C = C | 610 |
| C-H | 413 |
| C-Br | 280 |
| Br-Br | 193 |

a) Calculate the enthalpy change for the following reaction

$C_2H_{4(g)} + Br_{2(g)} \longrightarrow C_2H_4Br_{2(g)}$

b) Name the type of reaction that took place in **(a)** above

***1 mark***

10. Bond energies for some bonds are tabulated below:-

| BOND | BOND ENERGY KJ/mol |
|---|---|
| H – H | 436 |
| C = C | 610 |
| C- H | 410 |
| C - C | 345 |

Use the bond energies to estimate the enthalpy for the reaction

$C_2H_{4(g)} + H_{2(g)} \longrightarrow C_2H_{6(g)}$

11. The able shows the results obtained when 20.2g of potassium nitrate was added in 50cm$^3$ of water.

| Time in (min) | 0.0 | 0.3 | 1.0 | 1.3 | 2.0 | 2.3 | 3.0 | 3.3 | 4.0 |
|---|---|---|---|---|---|---|---|---|---|
| Temperature (°C ) | 25.0 | 25.0 | 25.0 | 25.0 | 17.0 | 17.0 | 20.0 | 20.0 | 20.0 |

(i) Draw the graph of temperature against time
(ii) Using the graph, determine the temperature change
(iii) Calculate the heat change
(iv) Find the molar heat of solution of potassium nitrate

12. When 1.6g of ammonium nitrate were dissolved in 100cm$^3$ of water, the temperature dropped by 6ºC. Calculate its enthalpy change. (Density of water = 1g/cm$^3$, specific heat capacity is 4.2kJ kg$^{-1}$K$^{-1}$)

13. Sodium hydrogen carbonate was strongly heated.

a) Write an equation for the reaction

b) The grid below shows part of the periodic table. Use it to answer the questions that follow. The letters are not the actual symbols.

| | | | | | | | | |
|---|---|---|---|---|---|---|---|---|
| A | | | | | | | | |
| B | | | | C | | | D | E |
| L | F | | G | H | | | J | |
| | | | | | | K | | |

i) Write the equation for the reaction that occurs between elements **L** and **D**

ii) The oxide of **G** reacts with both hydrochloric acid and sodium hydroxide. What is the nature of the oxide of **G**?
iii) Explain why elements **H** has a higher boiling points than element **D**.
iv) State **one** use of element **E**
v) Compare and explain the atomic radius of **B** and **C**
vi) 11.5g of **L was** completely burnt in oxygen .Calculate the volume of gas that was used.
(L = 23, molar gas volume at room temperature is 24dm$^3$)

14. A student has been provided with sodium hydroxide solution of 2M and hydrobromic acid of 4M. He was asked to investigate the equation for the reaction between these two substances and hence determine the molar enthalpy of neutralization. He carried out the reaction and obtained the following results:-

| **Vol. of 4M Hydrobromic acid added to 20cm$^3$ of 2M NaOH** | **Temperature of the mixture (°C)** |
|---|---|
| 4.0 | 26.8 |
| 6.0 | 30.0 |
| 8.0 | 33.2 |
| 10.0 | 36.0 |
| 12.0 | 35.2 |
| 14.0 | 34.4 |
| 20.0 | 30.8 |

(a) Draw a graph of the temperature of the mixture (vertical axis against the volume of the acid added) *
(b) Using the graph estimate the temperature of the mixture when 17cm$^3$ of the acid was added
(c) Both solutions were at room temperature at the start of the experiment. Use your graph to estimate the room temperature (½mk)
(d) What is the significance of the highest temperature of the solution mixture? *
(e) The temperature of the mixture increased during the first additions of the acid. Why did the temperature increase? *
(f) Suggest a reason why the temperature decreased during the latter part of the experiment
(g) Use your graph to determine the volume of 4M Hydrobromic acid which just neutralize 20cm$^3$ of 2M NaOH *
(h) How many moles of Hydrobromic acid are present in your answer in (**g**) above? *
(i) How many moles of NaOH are present in 20cm$^3$ of 2M of NaOH solution? *
(j) Use your answers in (**h**) and **(i**) above to write an equation of the reaction taking place in the experiment. Explain clearly how you have used your answers (1½mks)
(k) Determine the molar enthalpy of neutralization of hydrobromic acid
(1½mks)

15. (a) The following results were obtained in an experiment to determine the enthalpy of solution of sodium hydroxide

Mass of plastic beaker = 8.0g
Mass of plastic beaker + distilled water = 108.15g
Mass of plastic beaker + distilled water + sodium hydroxide = 114.35g

The table below shows the temperature at fixed times after mixing

| **Time/seconds** | 0 | 30 | 60 | 90 | 120 | 150 | 180 | 210 |
|---|---|---|---|---|---|---|---|---|
| **Temperature ( °C)** | 15 | 21 | 29 | 28 | 27 | 26 | 26 | 25 |

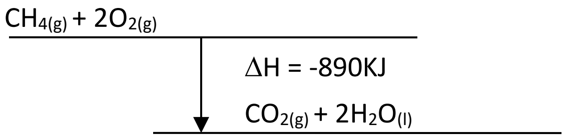
(i) Plot a graph of temperature (y-axis) against time (x-axis)
(ii) From your graph, determine the maximum temperature attained
(iii) Determine the temperature change of the reaction
(iv) Calculate the number of moles of sodium hydroxide used in the experiment

(Na = 11, H = 1, O = 16)

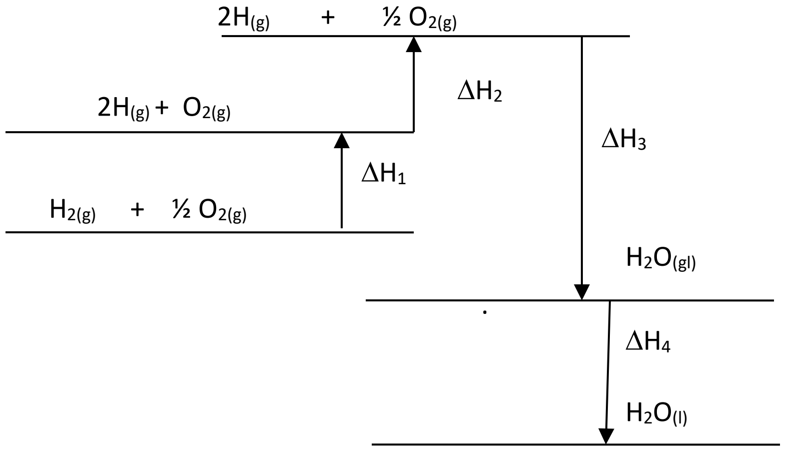
(v) Use your results to determine the molar enthalpy solution of sodium hydroxide. (Density of solution is 1g $cm^{-3}$ , specific heat capacity of solution = 4.18 $KJ^{-1}K^{-1}$)

(b) Below is an energy level diagram of the exothermic reaction

$CH_{4(g)} + 2O_{2(g)} \longrightarrow CO_{2(g)} + 2H_2O_{(l)}$ $\Delta H = -890KJ$



Examine the energy level diagram below and use it to answer the questions that follow



(b) (i) Which ΔH values will have negative sign?

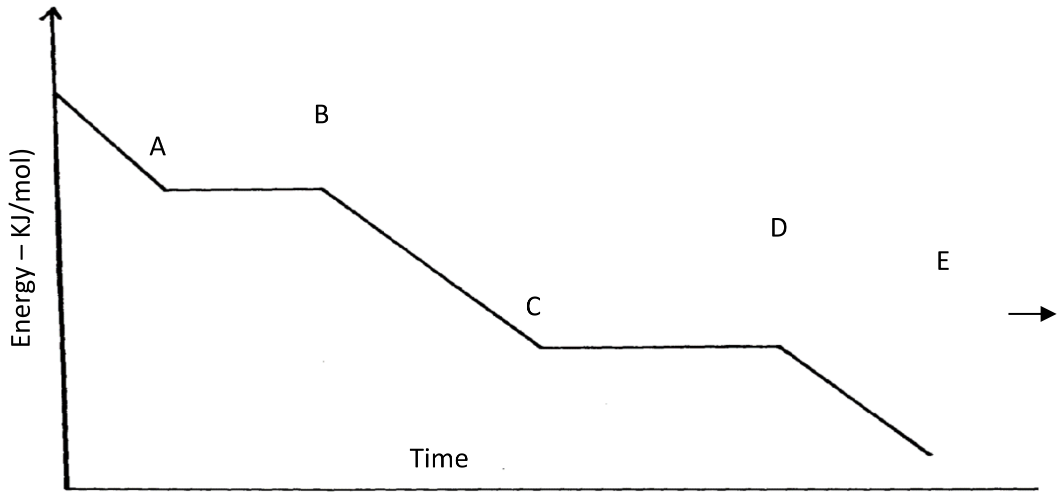
(ii) What physical change is being represented where enthalpy change $\Delta H_4$ is involved? (½mk)

(iii) In terms of $\Delta H_1$, $\Delta H_2$, $\Delta H_3$ and $\Delta H_4$, give the overall enthalpy change for the reaction:-

$H_{2(g)} + \frac{1}{2}O_{2(g)} \longrightarrow H_2O_{(l)}$

(iv) Is the reaction in **(iii)** above exothermic or endothermic?

16. (I) Study the graph below and answer the questions which follow:



(a) Distinguish between molar latent heat of fusion and molar latent heat of vaporization

(b) (i) Explain the changes occurring between points

**BC** ……………………………………… **CD** ………………………

(ii) In an experiment to determine molar enthalpy of neutralization of hydrochloric acid using potassium hydroxide, the data below was obtained. The concentration of potassium hydroxide used was 0.5M

| Volume of 0.5M KOH (cm$^3$) | 0 | 5 | 10 | 15 | 20 | 25 | 30 | 35 |
|---|---|---|---|---|---|---|---|---|
| Total volume of acid + Base | 20 | 25 | 30 | 35 | 40 | 45 | 50 | 55 |
| Temperature (°C) | 24 | 26 | 27 | 28 | 29 | 29 | 28 | 27 |

(i) Plot a graph of temperature (y-axis) against volume of potassium hydroxide used

(ii) From your graph:

(a) Determine the temperature change

(b) Find the volume of potassium hydroxide which completely neutralized 20cm$^3$ of the acid

(iii) Calculate the heat change for the reaction (C = 4.2Jg$^{-1}$K$^{-1}$ density of solution = 1g/dm$^3$)

(iv) Calculate the molar enthalpy of neutralization of hydrochloric acid with potassium hydroxide

17. A typical electrolysis cell uses a current of 40,000 amperes. Calculate the mass (in Kg of aluminium produced in one hour). (Al = 27) (Faraday = 96500Coloumbs )

18. (a) Biogas is a mixture of mainly Carbon (IV) Oxide and methane.
   (i) Give a reason why biogas can be used as a fuel
   (ii) Other than fractional distillation, describe a method that can be used to determine the percentage of methane in biogas

19. Consider the following equilibrium reaction.

$$H_2(g) + Cl_2(g) \rightleftharpoons 2HCl(g) \qquad \Delta H = -74.4KJ$$

a) State and explain the effect of formation of hydrogen chloride if pressure was increased in the equation above

20. Turning of fossil fuels has adverse environmental effects:-
   a) Name **two** pollutants from the burning of petroleum products
   b) Give **one** precaution taken to minimise the pollution by fossil fuels

21. (a) Define molar heat of neutralization
   (b) The rise in temperature when 50cm$^3$ of sodium hydroxide is reacted with two acids is given in the table below:-

| **Acid** | **50cm$^3$ of HCl** | **50cm of Oxalic acid** |
|---|---|---|
| Temp rise (°C) | 7 | 4 |

   (i) Explain the difference in the temperature.

22. Calculate the latent heat of vaporization of water

$$H_2O_{(l)} \longrightarrow H_2O_{(g)}$$

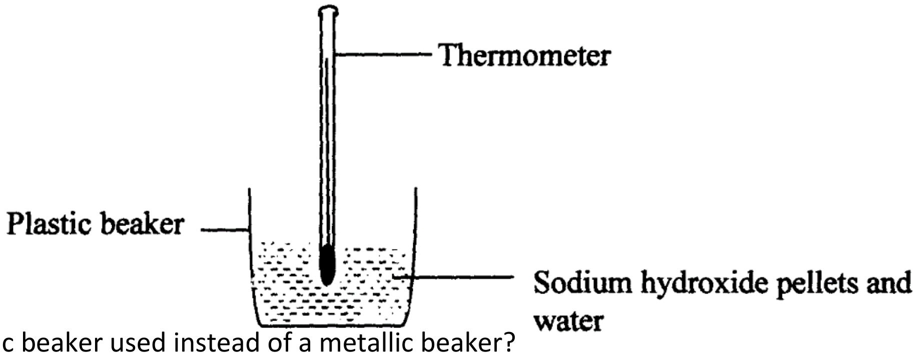
Given the following thermo chemical equations:-

$$H_{2(g)} + \tfrac{1}{2}O_{2(g)} \longrightarrow H_2O_{(g)} \quad \Delta H^\theta = -242KJ/Mol$$

$$H_{2(g)} + \tfrac{1}{2}O_{2(g)} \longrightarrow H_2O_{(l)} \quad \Delta H^\theta = -286KJ/Mol$$

23. (a) Define the term fuel
   (b) State **four** reasons why wood fuel is chosen for domestic cooking

24. The setup bellow was used to investigate the changes that take place when sodium hydroxide pellets dissolve in water.



   a) Why is a plastic beaker used instead of a metallic beaker?
   b) State and explain the observations made in the above reaction

25. (a) What is a fuel?
   (1mark)
   (b)Other than the cost, state **two** other factors to consider when choosing a fuel.

26. The equation below represents changes in the physical state of ions metal:

$$Fe_{(s)} \longrightarrow Fe_{(l)} \qquad DH = +15.4kjmol^{-1}$$

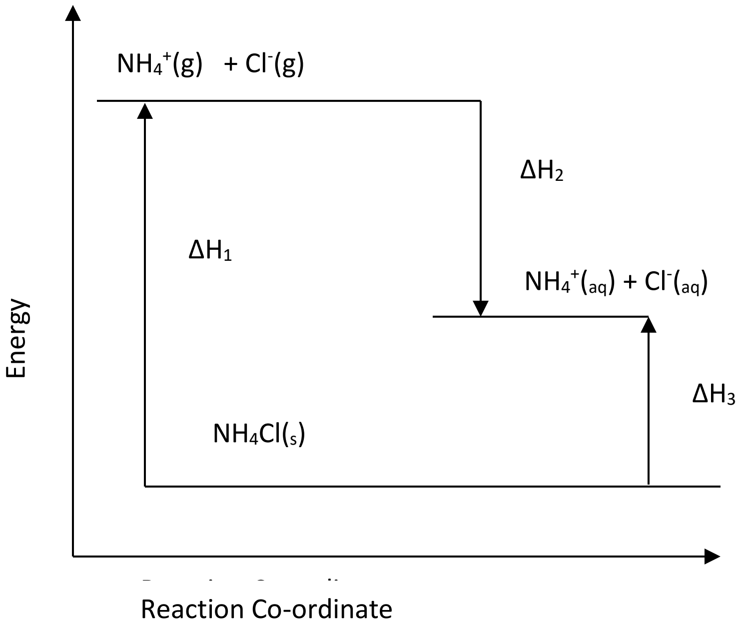
$$Fe_{(l)} \longrightarrow Fe_{(g)} \qquad DH = +354kjmol^{-1}$$

a) Calculate the amount of heat energy required to change 10kg of solid iron to gaseous iron
Fe = 56

b) Iodine can react with chlorine as shown below:-
$I_{2(g)} + Cl_{(g)} \longrightarrow 2Icl_{(s)}$ DH= -68kJ
Determine the molar enthalpy change for this reaction

c) Draw an energy level diagram for the reaction in **(b)** above

27. Study the diagram below and answer the questions that follow:



a) What do $\Delta H_1$ and $\Delta H_2$ represent?
$\Delta H_1$ …………………………………………………………………….
$\Delta H_2$ …………………………………………………………………….
b) Write an expression to show the relationship between $\Delta H_1$, $\Delta H_2$ and $\Delta H_3$.

KCSE SCHEME

# Energy changes in chemical and physical processes

***1. (a)*** $\nabla H = \frac{120 \times 4.2 \times 4.5}{1000}$ ***(½mk)***

$= +2.268KJ$ ✓ ***(½mk)***

***(b)*** ***RFM of*** $KNO_3 = 39 + 14 + 48 = 101$

$6g \longrightarrow 2.268KJ$

$101g \longrightarrow \frac{101 X 2.268}{6}$ ✓ ***(½mk)***

$= +38.178KJ mol^{-1}$ ***(½mk***

***2. (i) Heat evolved when one mole of a substance is completely burnt in oxygen***

***(ii) RFM of*** $C_2H_5OH = 46$

***Molar mass*** ✓ ½ ***= 46g***

***Heating value =*** $\frac{1370 KJ}{46g}$

✓ ½ ***= 29.78KJ/g (with units)***

***3.*** $Ca(q) + C(q) + 3/2 O2 (g)$

***4. a)*** $C_2H_6O_{(l)} + 3O_{(g)}$ ___________ $2CO_{2(g)} + 3H_2O$

***b) DH = MCDT***

$$\frac{200}{1000} \times 4.2 \times 32.5 = -27.3Kj$$

*0.92g $C_2H_6O$ ___________ - 27.3Kj*

*46g " ___________ ?*

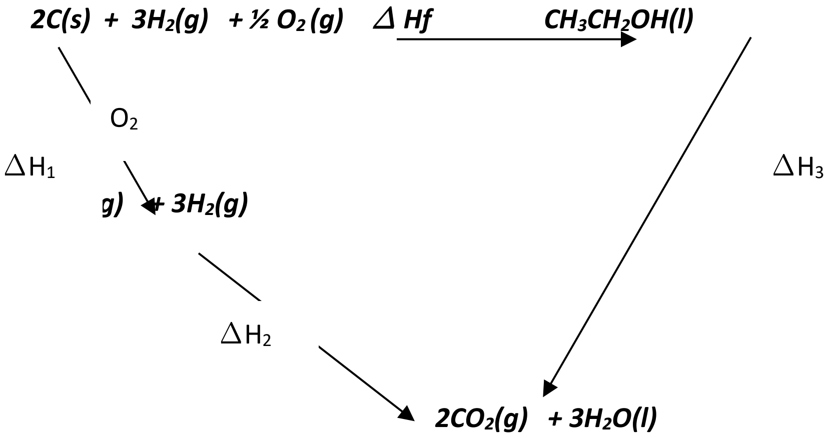
$$\frac{46g \times 27.3Kj}{0.92} = -1365Kj$$

*DHC $C_2HSO_4$ = -1365Kj mol*

**5.** *i) U,V,Y,Z All the 4 or nay 3 exclusively correct penalize ½ mk if wrong answer*
*ii) YZ is/are included any 2 correct ½ mk*

**6.** *(a) 611-389 = +222KJ*
*(b) H = +222 – (611 – 100)* ✓ ½
*= -289KJ* ✓1
*(c) Exothermic reaction* ✓ ½

**7.**



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

$\therefore \Delta Hf = \Delta H_1 + \Delta H_2 - \Delta H_3$ ✓½

$= -393 \times 2 + -286 \times 3 + 1386$ ✓1

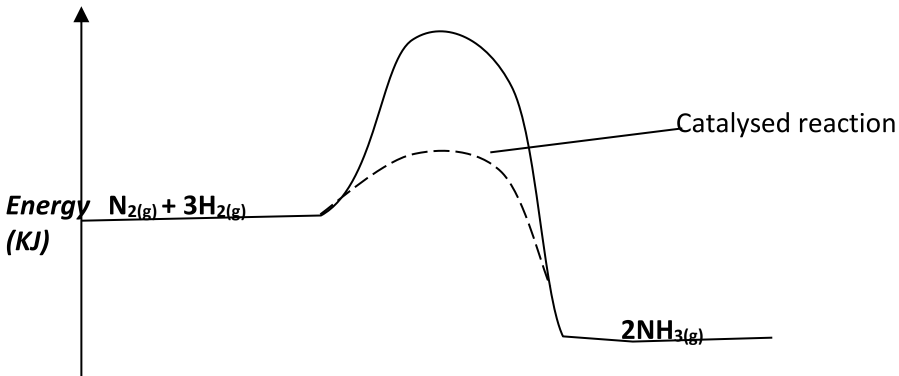
$= -786 - 858 + 1386$

$= -1644 + 1386$ ✓1

$\Delta Hf = -258\ KJmol^{-1}$ ✓½

**8.** *a) i) the yield of $NH_3$ would be lowered ✓½ any supply of heat makes $NH_3$ to decompose to $N_2$ and $H_2$*
*ii)the yield of $NH_3$ would be increased*
*b)a catalyst accelerate the rates of both forward and reverse reactions equally✓½ . Equilibrium position is not affected by a catalyst✓½*

*c)*



9. a) Breaking of 'C = C' = +610 KJ
Breaking of 'Br – Br' = +193
803✓
Formation of 2C – Br = -560
Formation of c-c +243 Kj
-346
- 103KJ✓ 2
marks
b) Addition reaction/ halogenation ✓

10. H H
C = C + H – H H – C – C – H
H H H H

| Bond breaking | Bond formation |
|---|---|
| 4 C-H – 4x410 = 1640 | 6C – H 6x410 |
| C = C – 1 x610 = 610 | = 2460 |
| H – H – 1x436 = 436 | C – C – 3 45 |
| 2686 | 2805 |

H = 2686 – 2805
= -119 Kj/Mol

11. (i) Graph
labeling -*TZM*
plotting – *TZM*
scale – *TZM*
line – *TZM*
total 5mks
(ii) Shown on the graph -*TZM*
(iii) Heat change = MCT
$= \frac{50}{100} x 4.2 x 10.2$
= 2.142kJ
(iv) RFM of $KNO_3$ = 39 + 14 + 48
= 101
$H = 2.142 x \frac{101}{20.2} = -10.71Kjmol^{-1}$

12. MCT $= \frac{100}{1000} X 4.2 X 6 = 2.52$ Kj
Moles of $NH4NO_3 = \frac{1.6}{80} = 0.02$ moles
If 0.02 mol _________ 2.52 Kj
1 mol _______________ $\frac{1 X 2.52}{0.02} = +126KJ/ mol$

13. a) $2 NaHCO_{3 (g)}$ _________ $Na_2CO_{3(g)} + H_2O_{(1)} + CO_{2(g)}$

b) i) $2L_{(g)} + D_{2(g)}$ _________ $2LD_{(g)}$

*ii) Amphoteric oxide*

*iii) Element H has a giant atomic structure with strong covalent bonds throughout its structure while D has simple molecular structure with weak Vander wall forces (2 m)*

*iv) - Used in advertising signs (Advertisements)*
*- Used in florescent tubes (Any two correct use)*

*v) C has a smaller atomic radius than B because it has stronger nuclear charge// more number of protons which attract the outer energy level electrons more firmly (2 mks)*

*vi)* $4L_{(s)} + O_{2(g)}$ __________ $2\,L_2O_{(g)}$

*Moles of L =* $\frac{11.5}{23}$ *= 0.5 moles*

*Moles of O2 =* $\frac{0.5}{4}$ *= 0.125 moles*

*Volume of* $O_2$ *= 0.125 mol X 24 = 3* $dm^3$

$4L_{(s)} + O_{2(g)}$ ___________ $2L2O_{(s)}$

*If 4 x23g* ________________ $24dm^3$

*11.5g of L* _________ $\frac{11.5 \times 24}{4 \times 23} = 3dm^3$

**14.** *(a) Drawn on the graph*

*A = ½ mk*
*S = ½ mk*
*P = ½ mk*
*C = ½ mk*

*b)* $32.5^oC \pm 1$ *Read from the student's correctly plotted graph.*

*c)* $20^oC \pm 0.5$ *Line is extrapolated downwards from the student's correct graph.*

*d) It is end point/ complete neutralization.*

*e) The reaction is exothermic hence as reaction proceeded more heat was produced.*

*f) Reaction was complete hence solution lost heat through radiation to the surrounding.*

*g)* $10.2\ cm^3 \pm 0.1$. *Read from the student's correct graph.*

*h) Moles =* $\frac{M \times V}{1000}$

$= \left(\frac{10.2 \times 4}{1000}\right)$ √½ *= 0.0408 moles* √½

*i) Moles =* $\left(\frac{M \times V}{1000}\right)$

$= \frac{2 \times 20}{1000}$ √½ *= 0.04 moles* √½

*j) HBr : NaOH*
*0.0408 : 0.04*
$\frac{0.0408}{0.04} : \frac{0.04}{0.04}$
*1 : 1*

$HBr_{(aq)} + NaOH \longrightarrow NaBr_{(aq)} + H_2O(l)$

***k)*** $\Delta H = MC\,\Delta t$

$= \dfrac{-30.2g \times 4.2J \times 16.3}{g^0c}$

$= -2067.49J$ √½

***Ans. in (h) = -2067.49 J.***

$\therefore$ ***1 Mole*** $= \dfrac{1 \times 2067.49J}{\text{Ans in “h”}}$ √½ ***e.g.*** $\dfrac{1 \times 2067.49}{0.0408}$

***= -Ans.*** ***e.g*** $50673.82\ J\ mol^{-1}$

***Or*** $50.67382KJ\ mol^{-1}$ √½

***15.*** ***a)(ii) Max. temperature attained :*** $29^0c$

***(iii) Temperature change o the reaction*** $= (29-115)^0c$

$= 14^0c$

***Mass of NaOH used = (114.35 – 108 .15)g***

***= 6.2g***

***R.F.M of NaOH = 40g***

***Moles of NaOH used =*** $\left[\dfrac{6.2}{40}\ moles\right]$

***= 0.155moles***

***(v) Heat released = Mass X Specific Heat capacity X Temperature change***

***Mass of water used = (108.15 – 8)g***

***= 100.15g***

$\therefore$ ***Heat released =*** $\left[\dfrac{100.15 \times 4.18 \times 14}{1000}\ kj\right]$

***=100.15kj***

***0.155 moles NaOH*** $\rightarrow 5.861\ kj$

***1 mole NaOH*** $\rightarrow \left[\dfrac{1x\ 5.861}{0.155}\ kj\ mole^{-1}\right]$

$= -37.8\ kjmol^{-1}$

***(b) i)*** $\Delta H_3$ ***and*** $\Delta H_4$

***ii) Condensation***

***iii)*** $\Delta H = \Delta H_1 + \Delta H_2 + \Delta H_3 + \Delta H_4$

***iv) Exothermic.***

***16.*** ***I – a – Latent heat of fusion is the heat change that occurs when one mole of a solid substance changes into liquid at constant temperature.***

***- Latent heat of vapourization is the heat change that occurs when one mole of liquid substance changes into gas at constant temperature.***

***b – BC – The liquid loses heat as it cools hence decrease in kinetic energy of the particles***

***- CD - The liquid changes to solid as temperature remains constant at freezing point.***

***II. (i) Scale – *TZM****

***Plot – *TZM****

***Line***

(ii) Should be shown on the graph – if not shown penalize ( ½ mk)

(iii) Heat change = m x c x $\Delta T$
Where m = (vol. of acid (20cm$^3$) + volume of bas in (b) above) x 1g/cm$^3$
$\Delta T$-as read form the graph
(iv) moles of acid

$$\text{Moles of base} = \frac{0.5 \times \text{volume}}{1000} \text{ in (b) above}$$

Mole ratio acid: Base = 1:1
Moles of acid $\longrightarrow$ heat change in (iii)above
1mole $\longrightarrow$ ?

$$\text{Molar heat change} = \frac{1 \times \text{heat in (iii)}}{\text{Moles of acid}}$$

17. Q = 40000 x 60 x 60 = 144000000c

$$\text{Mass of Al} = \frac{144000000 \times 27}{3 \times 96500} \quad ✓1$$

$$= 13.43\text{kg} \quad ✓1$$

18. (a) (i) Contains methane which is a fuel or contains methane which can burn
(ii) Pass a known volume of biogas through Sodium hydroxide (Potassium hydroxide) solution to absorb Carbon (IV) Oxide. Measure the volume ✓1 of remaining gas

$$\% = \frac{\text{Volume of methane}}{\text{Volume of Biogas}} \times 100$$

19. a) No effect – Reaction is not accompanied by volume changes/ similar volumes of reactants and products

20. a) – carbon IV Oxide;
- Sulphur IV Oxide;
- Lead;
(b) Availed low sulphur diesel/ availed unleaded petrol

21. (a) Heat change that occurs when one mole of hydrogen combines with one mole of hydroxide ions. //Heat evolved when one mole of water s formed during reaction of $H^+$ and $OH^-$ ions
(b) HCl produces a higher temperature rise than oxalic acid;
HCl is a stronger acid than oxalic acid;

22.

$$\begin{tikzcd}
H_2O_{(l)} \arrow[rr, "\Delta H_2"] & & H_2O_{(g)} \\
& H_{2(g)} + \frac{1}{2} O_2(g) \arrow[ul, "\Delta H_1"] \arrow[ur, "\Delta H_3"'] &
\end{tikzcd}$$

$\Delta H_2 = -\Delta H_1 + \Delta H_3$ ✓½
$= \Delta H_3 - \Delta H_1$ ✓½ ✓1
$= -242 - {}^{-}286$
$= -242 + 286$ ✓1
$= +44$KJ/mol ✓1 (No units of sign = ½mk)

23. (a) Chemical substance that burns to produce useful amount of heat.
(b) (i) Its cheap
(ii) Its readily available (½mk)
(iii) It burns slowly (½mk)

***(iv) Does not produce poisonous gas. (½mk)***

***24. a) Metallic beaker would make most of the heat be lost to the environment***
***b) - Thermometer reading increased***
***- The reaction is exothermic***

***25. a) A substance that produce heat energy when burnt***

***b) 1. Availability***
***2. ease of transport***

***26. a) 1 mole Fe (56) required _____________ 15.4 + 354***
***= 396.5Kj***
***10,000 (10 kg) ____________ ?***

$$\frac{10{,}000g}{56g} \times 369.5\ Kj$$

***= 6596.285Kj***

***b)*** $\frac{-68Kj}{2}$ ***= - 34 Kj √ ½***

***27. a)*** $\Delta H_1$ ***– Lattice energy ✓1***
$\Delta H_2$ ***– Hydrogen energy ✓1***
***b)*** $\Delta H_3 = \Delta H_2 + \Delta H_1$ ***✓1***