ENERGY CHANGES IN CHEMICAL PROCESSES

Energy is the capacity of overcoming resistance or doing work. It is a property of matter; it can do work and can be obtained from work. Chemical changes are normally accompanied by energy changes. The most common form of energy change in chemical reactions is the heat (enthalpy) change.

Endothermic and Exothermic Reactions

Endothermic reactions.

These are reactions during which heat is absorbed from the surroundings. The temperature of the surroundings decreases. The products are at a higher energy level than the reactants and the enthalpy change is positive ($+ \Delta H$)

Examples of endothermic reactions.

    (i)   Dissolving ammonium nitrate in water.

    (ii)   Dissolving sodium thiosulphate in water.

Exothermic reactions.

 These are reactions during which heat is liberated to the surroundings. The temperature of the surrounding increases. The products are at a lower energy levels than the reactants. And the enthalpy change is negative ($-\Delta H$).

Examples of exothermic reactions

    (i). Adding concentrated sulphuric acid to little water in a test tube.

    (ii). Reacting dilute sulphuric acid with magnesium ribbon.

    (iii). Reacting sodium hydroxide and dilute hydrochloric acid.

    (iv). Burning magnesium in air.

Enthalpy

The energy or heat content of a system is called the enthalpy. It is denoted by the symbol H. The enthalpy change of the system is indicated as ΔH (Δ means *Change of*) and refers to heat evolved or absorbed.

Heat energy is measured in kilojoules (kJ). The heat change in a chemical reaction (ΔH) is expressed as a negative value if heat is evolved and a positive value if heat is absorbed.

    $\Delta H = {}^-92$ kJ means the heat change for the reaction is exothermic.

    $\Delta H = {}^+113$kJ means the heat change for that reaction is endothermic.

HEAT OF REACTION

The heat of reaction for a chemical reaction is the amount of heat evolved or absorbed when molar quantities of substances react together completely.
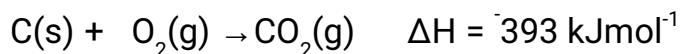
Types of Heats of Reactions

    1. *Heat of combustion*

This is the amount of heat liberated when one mole of a substance is completely burnt in oxygen.

Example 1;

Carbon undergoes complete combustion according to the equation.

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

Calculate the amount of heat evolved when 20 g of carbon is completely burnt.

Solution 1

12 g of carbon liberates 393 kJ

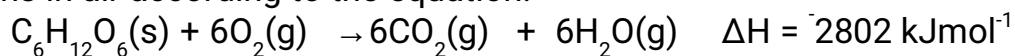1 g of carbon liberates $\dfrac{393}{12}$ kJ

Therefore 20 g of carbon liberates $\dfrac{393 \times 20}{12}$ kJ

$$= {}^{-}655 \text{ kJ}$$

Example 2;

Glucose burns in air according to the equation.

$$C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(g) \quad \Delta H = {}^{-}2802 \text{ kJmol}^{-1}$$

Calculate the amount of heat produced when 45 g of glucose is completely burnt in air.

Solution 2

Molar mass of glucose = $(12 \times 6) + (1 \times 12) + (16 \times 6)$ = 180 g

180 g of glucose liberates 2802 kJ

1 g of glucose liberates $\dfrac{2802}{180}$ kJ

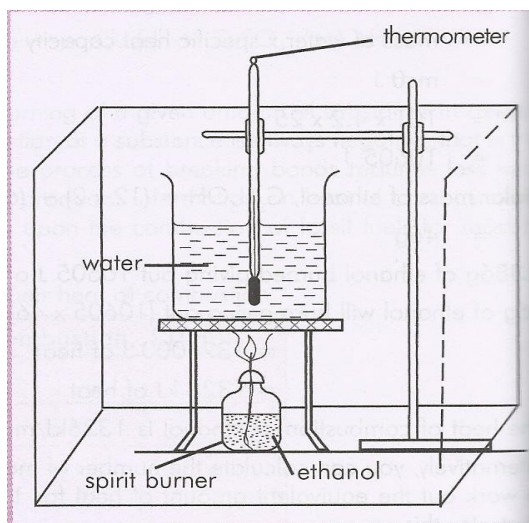Therefore 45 g of glucose liberates $\dfrac{2802 \times 45}{180}$ kJ

$$= {}^{-}700.5 \text{ kJ}$$

Determination of Heat of Combustion of Ethanol

Procedure

A known mass of water (150 cm^3) is put in a metal can. Initial temperature of water is noted, t_1 °C

The spirit lamp is filled with ethanol and weighed. The mass of the lamp and ethanol before burning is noted, m_1 g

The lamp is lit and placed under the metal can. To shield it from draughts; the whole apparatus is placed in a large tin to prevent heat loss.

The metal tin is heated until the temperature of the water. The final temperature of the

water is noted, t_2 °C

The lamp is putout immediately and reweighed to find the new mass of lamp and ethanol after burning, m_2 g

Treatment of results:

Mass of ethanol = $(m_1 - m_2)$g

Mass of water = 150 g

Temperature rise = $(t_2 - t_1)$°C

Specific heat capacityof water = 4.2 Jkg^{-1}K^{-1}

Heat gained by water = mass X specific heat capacity X temeparature rise

$= 150 \times 4.2 \times (t_2 - t_1)$

Molar mass of ethanol, C2H5OH = (12X2) + (1X5) + 16 + 1 = 46

$(m_1 - m_2)$g of ethanol liberates $630(t_2 - t_1)$

1 g of ethanol liberates $\dfrac{630(t_2 - t_1)}{(m_1 - m_2)}$ kJ

Therefore 46 g of ethanol liberates $\dfrac{630(t_2 - t_1) \times 46}{(m_1 - m_2)}$ kJmol^{-1}

Note: The heat of combustion in this experiment would be lower than the theoretical value because;

Heat is lost to the environment during the experiment.

Some heat produced warms the metal can and air but does not pass into the water.

Example 1;

When 0.382 g of ethanol was burnt, the heat liberated raised the temperature of 100 g of water from 16.5°C to 43.5°C. Calculate the heat of combustion of ethanol
(specific heat capacity of water =Jg^{-1}k^{-1}, C =12, H =1, O =16).

Solution 1;

Heat gained by water = MCθ

$= 100 \times 4.2 \times (43.5 - 16.5)$

$= 11340$ J

$= \dfrac{11340}{1000}$ kJ

$= 11.34$ kJ

Molar mass of ethanol = (12X2) + (1X5) + 16 + 1 = 46

0.382 g of ethanol liberate 11.34 kJ

Therefore 46 g of ethanol liberate $\dfrac{-11.34 \times 46}{0.382}$ kJmol^{-1}

Example 2;

When 0.16 g of methanol was burnt, the heat generated raised the temperature of 200 g of water from 21 °C to 25 °C. Calculate the heat of combustion of methanol (specific heat

capacity of water= 4.2 $Jg^{-1}k^{-1}$, C=12, O=16, H =1).

Solution 2;

Heat gained by water = MCθ

= 200 X 4.2 X (25-21)

= 3360 J = 3.36 kJ

Molar mass of ethanol = 12 + (1X3) + 16 + 1 = 32
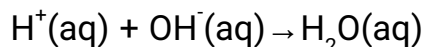
0.16 g of methanol liberate 3.36 kJ

1 g of methanol liberates $\dfrac{3.36}{0.16}$ kJ

Therefore 32 g of methanol liberates $\dfrac{3.36 \text{ X } 32}{0.16}$ $kJmol^{-1}$

= 672 $kJmol^{-1}$

2. Heat of Neutralization

This is the amount of heat evolved when one mole of aqueous hydrogen ions reacts with one mole of aqueous hydroxide ions to form one mole of water.

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(aq)$$

Determination of Heat of Neutralization Between Sodium Hydroxide and Hydrochloric Acid

Procedure

A fixed volume of 2 M hydrochloric acid (100 cm^3) is measured into a plastic beaker. The initial temperature of the acid is noted.

A fixed volume of 2 M sodium hydroxide (100 cm^3) is measured into another plastic beaker and its initial temperature noted.

Hydrochloric acid is poured into a beaker containing sodium hydroxide, stirring gently with a thermometer. The highest temperature of the mixture is noted.

Treatment of results

Initial temperature of hydrochloric acid = 21.0 ^{0}C.

initial temperature of sodium hydroxide= 22.0 ^{0}C

Average temperature of solutions=21.5 ^{0}C

Final temperature of the mixture = 34.5 ^{0}C

Specific heat capacity of solution = 4.2 $Jg^{-1}k^{-1}$

Heat evolved = MCθ

= 200 X 4.2 X (34.5-21.5) J

= 10920 J = 10.92 kJ

Moles of acid or alkali = $\dfrac{2 \text{ X } 100}{1000}$ moles = 0.2 moles

0.2 moles hydrochloric acid liberates 10.92 kJ

Therefore 1 mole of hydrochloric acid liberate $\dfrac{10.92 \text{ X } 1}{0.2}$ $kJmol^{-1}$

= 54.6 $kJmol^{-1}$

Example 1;

50 cm^3 of 2 M hydrochloric acid reacts with 50 cm^3 of 2 M sodium hydroxide, both at 22 ^{0}C

in a plastic beaker. The mixture was stirred and its maximum temperature was 35 ^{0}C (specific heat capacity of solution=4.2 Jg^{-1}k^{-1}, density of solution=1 gcm^{-1}). Calculate the enthalpy of neutralization of sodium hydroxide.

Solution 1
Heat gained by the solution = MCθ

$$= 100 \times 4.2 \times (35\text{-}22)J = 5.46 \text{ kJ}$$

Moles of sodium hydroxide $\dfrac{50 \times 2}{1000}$ 0.1 moles

0.1 moles of sodium hydroxide liberates 5.46

Therefore 1 mole of sodium hydroxide liberates $\dfrac{5.46 \times 1}{0.1}$ kJmol^{-1}

$$= 54.6 \text{ kJmol}^{-1}$$

Example 2;
When 50.0 cm^3 of 1.0 M sulphuric acid was added to 50.0 cm^3 of 2 M sodium hydroxide, the temperature of the resultant mixture rose by 13.6 ^{0}C
(specific heat capacity of solution= 4.2 Jg^{-1}k^{-1}, density of solution=1gcm^{-1}).
Calculate the enthalpy of neutralization of sodium hydroxide.

Solution 2;
Heat gained by solution = MCθ

$$= (100 \times 4.2 \times 13.6)J = 5.712 \text{ kJ}$$

Moles of sodium hydroxide $\dfrac{50 \times 2}{1000}$ 0.1 moles

0.1 moles of sodium hydroxide liberates 5.712

Therefore 1 mole of sodium hydroxide $\dfrac{5.712 \times 1}{01}$ kJmol^{-1}

$$= 57.12 \text{ kJmol}^{-1}$$

3. Heat of Solution
Enthalpy of solution is the amount of heat evolved or absorbed when one mole of a substance is completely dissolved in water so that further dissolution causes no further temperature changes.

Determination of Heat of Solution of Ammonium Nitrate
Procedure
A known volume of water (100 cm^3) is placed in a beaker and its temperature noted.
A known mass of ammonium nitrate (4.0 g) is added to the water in the beaker.
Stir the mixture well with a stirrer and record the lowest temperature of the solution reached.

Typical results
Initial temperature of water= 18.5 ^{0}C
Final temperature of water = 16.0 ^{0}C
Temperature fall   = 2.5 ^{0}C
Mass of ammonium nitrate= 4.0 g
Mass of water   =100 g
Specific capacity of water =4.2 Jg^{-1}k^{-1}

$$\text{Heat absorbed = MCθ} = 100 \times 4.2 \times 2.5 = 1050 \text{ J}$$
$$= 1.05 \text{ kJ}$$

Molar mass of ammonium nitrate = 80 g

4.0 g of ammonium nitrate liberares 1.05 kJ

1 g of ammonium nitrate liberate $\dfrac{1.05}{4.0}$ kJ

Therefore 80 g of ammonium nitrate liberates $\dfrac{1.05 \times 80}{4.0}$ kJmol^{-1}

$$= {}^{+}21.0 \text{ kJmol}^{-1}$$

Example 1;

0.5 g of ammonium nitrate was dissolved in 50 cm^3 of water. The temperature dropped from 24 ^{0}C to 18 ^{0}C. Calculate the enthalpy of solution of ammonium nitrate (specific heat capacity of water=4.2 Jg^{-1k}-1, density of water=1 gcm^{-3}, N=14, H=1, O=16).

Solution 1;

Heat absorbed = MCθ

$= 50 \times 4.2 \times (24-18)$

$= 1.26$ kJ

Molar mass of ammonium nitrate = 80 g

0.5 g of ammonium nitrate liberate 1.26 kJ

1 g of ammonium nitrate liberate $\dfrac{1.26}{0.5}$ kJ

Therefore 80 g of ammonium nitrate liberate $\dfrac{1.26 \times 80}{0.5}$

$$= +201.6 \text{ kJmol}^{-1}$$

4. Heat of Displacement

This is the amount of heat absorbed or evolved when one mole of a substance is displaced from its compound.

Determination of Heat of Displacement of Copper

Procedure

Measure 50 cm^3 of 0.2 M copper(II) sulphate solution into a plastic bottle fitted with a bung carrying a thermometer.

Record the initial temperature of the solution.

Weigh 1 g of iron filings, add the filings to the solution and shake well.

Record the highest temperature attained by the solution.

Results and calculations

Initial temperature of CuSO4 =21.0 ^{0}C

Final temperature   =27.0 ^{0}C

Rise in temperature = 6.0 ^{0}C

Mass of 50 cm^3 of CuSO4 solution = 50 g

Specific heat capacity of solution =4.2 Jg^{-1}k^{-1}

Heat gained by solution = MCθ = 50 × 4.2 × 6.0   = 1260 J = 1.26 kJ

moles of copper(II)sulphate used $\dfrac{0.2 \times 50}{1000}$ = 0.01 moles

0.01 moles of copper(II) sulphate liberate 1.26 kJ

Therefore 1 mole of copper(II)sulphate liberate $\dfrac{1.26}{0.01}$ kJmol^{-1}

$$= \,^-126 \text{ kJmol}^{-1}$$

The thermo chemical equation for the reaction is

$$CuSO_4(aq) + Fe(s) \rightarrow Cu(s) + FeSO_4(aq) \qquad \Delta H = \,^-126 \text{ kJmol}^{-1}$$

Example 1:

(a). When an excess of iron filings is added to 25 cm^3 of 0.4 M copper(II) solution, there was a temperature rise of 15 ^{0}C. Calculate the molar enthalpy of displacement of copper (specific heat capacity of solution=4.2 Jg^{-1}k^{-1}, density of solution=1 gcm^{-3}).
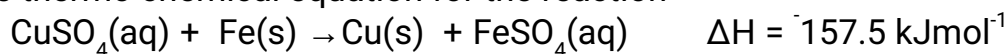
Solution 1;

$$\text{Heat gained by solution} = MC\theta = 25 \times 4.2 \times 15 = 1.575 \text{ kJ}$$

$$\text{moles of copper(II)solution used } \frac{0.4 \times 25}{1000} = 0.01 \text{ moles}$$

0.01 moles of copper(II) solution liberates 1.575 kJ

$$\text{Therefore 1 mole of copper(II) solution liberate } \frac{1.575}{0.01} \text{kJmol}^{-1} = \,^-157.5 \text{ kJmol}^{-1}$$

(b). Write the thermo chemical equation for the reaction

$$CuSO_4(aq) + Fe(s) \rightarrow Cu(s) + FeSO_4(aq) \qquad \Delta H = \,^-157.5 \text{ kJmol}^{-1}$$

Example 2;

To 100 cm^3 of 0.2 M copper(II) sulphate solution in a plastic beaker was added excess zinc dust. There was a temperature change from 20 ^{0}C to 30.4 ^{0}C.

(a). Why was a plastic beaker used?

The plastic beaker is a poor conductor of heat and minimizes heat loss to the surrounding.

(b). State what was observed.

The blue copper(II) sulphate solution turns to colourless and a brown solid is deposited.

(c). Calculate the molar heat of the reaction

(specific heat capacity of solution=4.2 Jg^{-1}k^{-1}, density of solution=1gcm^{-3}).

Answer(=218.4 kJmol$^{1)}$)

(d). Write a thermochemical equation for the reaction.

CuSO4(aq)+Zn(s)⟶Cu(s)+ZnSO4(aq)   ΔH=+218.4 kJmol^{-1}