

Topic 20 – Energy Changes

Subject content

- (a) describe the meaning of enthalpy change in terms of exothermic (ΔH negative) and endothermic (ΔH positive) reactions
- (b) represent energy changes by energy profile diagrams, including reaction enthalpy changes and activation energies
- (c) describe bond breaking as an endothermic process and bond making as an exothermic process
- (d) explain overall enthalpy changes in terms of the energy changes associated with the breaking and making of covalent bonds
- (e) describe hydrogen, derived from water or hydrocarbons, as a potential fuel, reacting with oxygen to generate electricity directly in a fuel cell (details of the construction and operation of a fuel cell are not required).

Definition

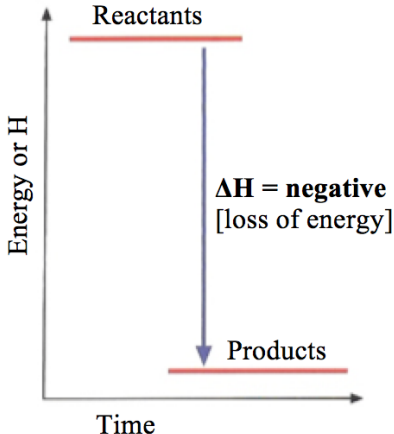
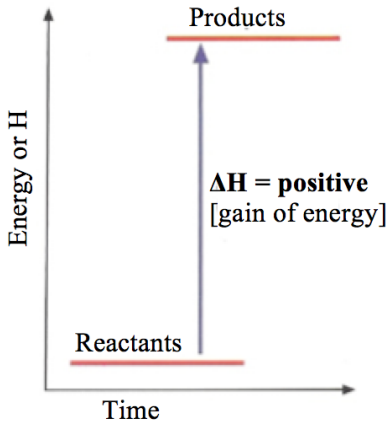
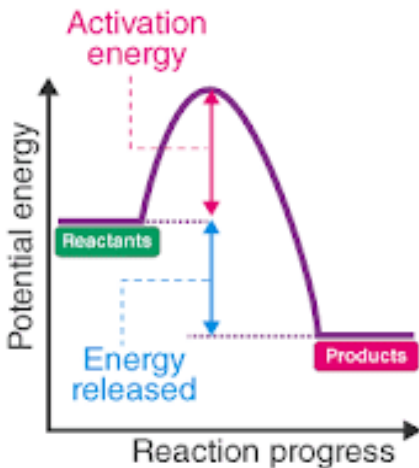
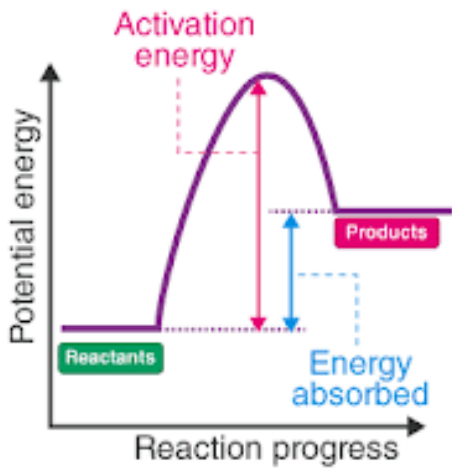
| Term | Definition | SI unit |
|------------------------------------------------|---------------------------------------------------------------------------------------------------------------------------------|----------|
| Exothermic change | Change where heat is released into surroundings | |
| Endothermic change | Change where heat is absorbed from surroundings | |
| Enthalpy change (ΔH) | Net amount of heat energy absorbed / released during chemical reaction (= difference in energy content of reactants & products) | kJ / mol |
| Activation energy (E_a) | Minimum amount of energy possessed by reactant particles for chemical reaction to occur | kJ |

Formulae

| Enthalpy change | Enthalpy change | Activation energy |
|-------------------------------------------------------|--------------------------------------------------------|----------------------------------------------|
| $\Delta H = H_{\text{product}} - H_{\text{reactant}}$ | $\Delta H = H_{\text{absorbed}} - H_{\text{released}}$ | $E_a = H_{\text{max}} - H_{\text{reactant}}$ |

20.1 Exothermic and Endothermic Changes

Exothermic & endothermic changes

| Aspect | Exothermic | Endothermic |
|--------------------------------------------------|-----------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------|------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------|
| Definition | Heat → surrounding (release) | Heat ← surrounding (absorb) |
| Energy level diagram (arrow: reactant → product) |  <p>Energy or H</p> <p>Reactants</p> <p>Products</p> <p>$\Delta H = \text{negative}$ [loss of energy]</p> <p>Time</p> |  <p>Energy or H</p> <p>Reactants</p> <p>Products</p> <p>$\Delta H = \text{positive}$ [gain of energy]</p> <p>Time</p> |
| Energy profile diagram |  <p>Potential energy</p> <p>Reactants</p> <p>Products</p> <p>Activation energy</p> <p>Energy released</p> <p>Reaction progress</p> |  <p>Potential energy</p> <p>Reactants</p> <p>Products</p> <p>Activation energy</p> <p>Energy absorbed</p> <p>Reaction progress</p> |
| Characteristic | <ul style="list-style-type: none"> • Release heat to surroundings • Increase in surrounding temp • -ve ΔH • Energy: product < reactant | <ul style="list-style-type: none"> • Absorb heat from surroundings • Decrease in surrounding temp • +ve ΔH • Energy: product > reactant |
| Reason | Energy released to form bonds in product > energy absorbed to break bonds in reactant | Energy released to form bonds in product < energy absorbed to break bonds in reactant |
| Examples | <ul style="list-style-type: none"> • Condensation • Freezing • Dissolving acids • Dissolving anhydrous salts | <ul style="list-style-type: none"> • Evaporation, boiling • Melting • Dissolving ammonium salts |
| | <ul style="list-style-type: none"> • Combustion • Respiration • Neutralisation | <ul style="list-style-type: none"> • Photosynthesis • Thermal decomposition (e.g. calcium carbonate) |

| | | |
|------------------------------------|-----------------------------------------------------------------------------------------------------------------------------------------------------------------------|-------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------|
| | <ul style="list-style-type: none"> Corrosion of metals Reaction b/w acid & metal Displacement reaction Heat pack | <ul style="list-style-type: none"> Action of light on silver bromide on photographic film Reaction of ammonium nitrate + barium hydroxide Electrolysis of water Reaction of acetic acid + bicarbonate |
| Variation of temperature with time | <ul style="list-style-type: none"> Initial: rise → highest temp Complete: drop → room temp | <ul style="list-style-type: none"> Initial: drop → lowest temp Complete: rise → room temp |

* **ARROW**: one direction (specify positive / negative change)

Enthalpy

Enthalpy: energy content of substance (not measurable)

High enthalpy → less stable → release a lot of energy during reaction

Enthalpy change (measurable)

ΔH = total energy of products – reactants

Activation energy

Heat provides energy: change less energetic → more energetic particles to start reaction

Collision theory

- Reactant particles collide with each other → reaction occur
- Rate of chemical reaction proportional to no. of collisions b/w reactant particles
- Requirements
 - Reacting particles collide with each other
 - Reacting particles have proper orientation
 - Reacting particles have sufficient energy (\geq activation energy)

20.2 Bond Breaking and Bond Making

Energy changes

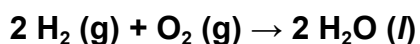
- Bond breaking** → endothermic ($\Delta H > 0$): heat absorbed
- Bond making** → exothermic ($\Delta H < 0$): heat released

Exothermic OR endothermic

| Overall enthalpy change (ΔH) = energy absorbed (bond breaking) – released (bond making) | |
|------------------------------------------------------------------------------------------------------|-----------------------------------------|
| Exothermic | Endothermic |
| absorbed (breaking) < released (making) | absorbed (breaking) > released (making) |

Hydrogen + oxygen → water

Chemical equation:



Particles present:

| Hydrogen molecule (2 mol) | Oxygen molecule (1 mol) | Water molecule (2 mol) |
|------------------------------------|-----------------------------------|------------------------------------|
| | | |
| $2 \times 1 = 2 \text{ H-H bonds}$ | $1 \times 1 = 1 \text{ O=O bond}$ | $2 \times 2 = 4 \text{ O-H bonds}$ |

Energy changes

| <u>Energy absorbed – bond breaking</u> <u>(endothermic)</u> | <u>Energy released – bond making</u> <u>(exothermic)</u> |
|-----------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------|------------------------------------------------------------------------------------------------------------------|
| Energy to break O=O bond = 496 kJ Energy to break H-H bond = 436 kJ Total energy for bond breaking = $496 + 2 (436) = 1368 \text{ kJ}$ | Energy to form O-H bond = 463 kJ Total energy for bond making = $2 (2 (463)) = 1852 \text{ kJ}$ |
| <p align="center"><u>Overall heat change (enthalpy change)</u></p> <p align="center"> $\Delta H = \text{energy absorbed} - \text{released}$ $= 1368 - 1852 = -484 \text{ kJ/mol}$ $\therefore \Delta H < 0$ \therefore The reaction is exothermic. </p> | |
| <u>Energy level diagram</u> | <u>Energy profile diagram</u> |
| | |

* bond enthalpies will be given in exam paper

Bond energy: energy absorbed to break / released to make bonds

The stronger a bond is,
the more energy is required to break the bond,
the higher its bond energy

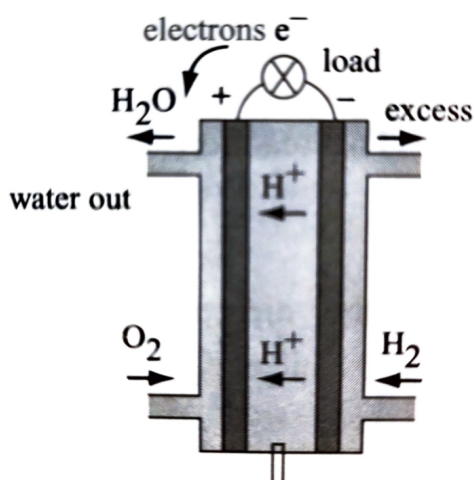
20.3 Combustion of Fuels

Fuel as source of energy

Combustion of fuels: exothermic → release large amount of energy

- **Fuel:** substance that burns easily in presence of oxygen → produce energy
 - Methane: $\text{CH}_4(\text{g}) + 2 \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{g}); \Delta H = -818 \text{ kJ/mol}$
 - Propane: $\text{C}_3\text{H}_8(\text{g}) + 5 \text{O}_2(\text{g}) \rightarrow 3 \text{CO}_2(\text{g}) + 4 \text{H}_2\text{O}(\text{g}); \Delta H = -2052 \text{ kJ/mol}$
- **Fuel cell:** chemical cell to convert chemical energy (in fuels) → electrical energy (useful)

Hydrogen fuel cell (Refer to Topic 19)



| Anode (oxidation) | Cathode (reduction) |
|---------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------|------------------------------------------------------------------------------------------------------------|
| hydrogen | oxygen |
| $2 \text{H}_2(\text{g}) \rightarrow 4 \text{H}^+(\text{aq}) + 4 \text{e}^-$ | $\text{O}_2(\text{g}) + 2 \text{H}^+(\text{aq}) + 4 \text{e}^- \rightarrow 2 \text{H}_2\text{O}(\text{l})$ |
| $2 \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{H}_2\text{O}(\text{l}); \Delta H = -242 \text{ kJ/mol}$ <ul style="list-style-type: none"> ● Generate electricity + water + heat ● Hydrogen reacts with oxygen by undergoing <u>combustion</u> <ul style="list-style-type: none"> ○ Combustion is exothermic – produce heat energy ○ End product of combustion is water | |

Advantages & disadvantages

| Advantage | Disadvantage |
|---------------------------------------------------------------------------------------------------------------------------------------------------------------------|------------------------------------------------------------------------------------------------------------------------------|
| 1. Energy efficient (twice amt of energy compared to other fuels) 2. Does not cause pollution (end product is water) 3. Renewable source of energy (hydrogen) | 1. Expensive to obtain hydrogen 2. Hydrogen flammable + explosive 3. Small mass of hydrogen stored in large quantities |

Typical questions**Multiple choice questions**

1 Which of the following phase changes is endothermic?

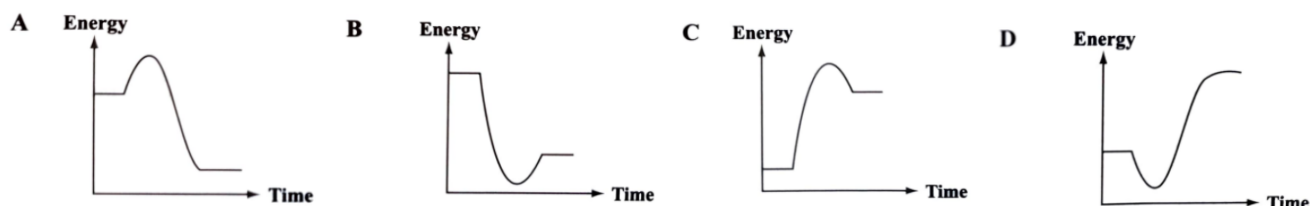
A $\text{H}_2\text{O} (l) \rightarrow \text{H}_2\text{O} (g)$ (Evaporation / Boiling)

B $\text{I}_2 (g) \rightarrow \text{I}_2 (s)$ (Sublimation)

C $\text{Hg} (l) \rightarrow \text{Hg} (s)$ (Freezing)

D $\text{H}_2\text{S} (g) \rightarrow \text{H}_2\text{S} (l)$ (Condensation)

2 Given the reaction $\text{S} (s) + \text{O}_2 (g) \rightarrow \text{SO}_2 (g) + \text{energy}$, which energy profile diagram best represents the enthalpy changes for this reaction?



3 Which of the following processes is exothermic?

A Decomposition by heat of zinc carbonate into zinc oxide and carbon dioxide

B Formation of glucose and oxygen during photosynthesis (Absorb light energy)

C Conversion of glucose to carbon dioxide and water during respiration (Release energy)

D Sublimation of dry ice (Absorb heat to break bonds)

4 Which of the following does not involve an endothermic change?

A An iron nail rusting in air

B A few drops of ethanol drying up (Evaporation)

C A potted plant making food in sunlight (Photosynthesis)

D A candle melting (Melting)

5 The reactions between hydrogen and oxygen to form water occurs in three stages. In which stages are the signs of the enthalpy changes correctly shown?

1. $2 \text{H}_2 (g) + \text{O}_2 (g) \rightarrow 4 \text{H} (g) + 2 \text{O} (g)$; $\Delta H = \text{positive}$ (Bond breaking \rightarrow endothermic)

2. $4 \text{H} (g) + 2 \text{O} (g) \rightarrow 2 \text{H}_2\text{O} (g)$; $\Delta H = \text{positive}$ (Bond forming \rightarrow exothermic)

3. $\text{H}_2\text{O} (g) \rightarrow \text{H}_2\text{O} (l)$; $\Delta H = \text{negative}$ (Condensation \rightarrow exothermic)

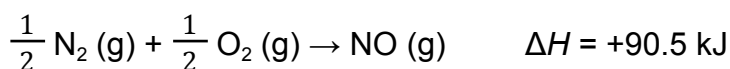
A 1 and 2

B 1 and 3

C 2 and 3

D 1, 2 and 3

- 6 The thermochemical equation for the formation of nitrogen monoxide (NO) in a car engine is shown below.



This implies that:

- A** the product has more energy than the reactants
- B** the total energy change for bond making is greater than that for bond breaking
- C** the formation of the N–O bond in an NO molecule is an endothermic process
- D** heat energy is released to the surroundings during the reaction

Structured questions

- 1 Given the following information, calculate the bond energy of the following reactions:

| Bond | Bond energy (kJ/mol) |
|--------------|----------------------|
| C – C | +347 |
| C – H | +413 |
| C – O | +325 |
| H – H | +436 |
| O – H | +463 |
| C = O | +805 |
| O = O | +496 |
| N – H | +391 |
| N \equiv N | +944 |

- (a) The Haber process: $\text{N}_2 (\text{g}) + 3 \text{H}_2 (\text{g}) \rightarrow 2 \text{NH}_3 (\text{g})$

| Bond breaking | Bond forming |
|-------------------------------------------------------------------------------------------------------|---------------------------------------------------------------------------|
| Break 3 H – H bonds, 1 N \equiv N bond Amount of energy required = 3 (436) + 944 = 2253 kJ | Form 6 N – H bonds Amount of energy released = 6 (391) = 2346 kJ |

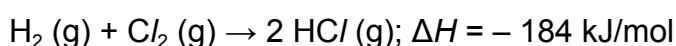
$$\Delta H = 2252 - 2346 = \underline{\underline{-94 \text{ kJ/mol}}}$$

(b) Burning of methane: $\text{CH}_4(\text{g}) + 2 \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{g})$

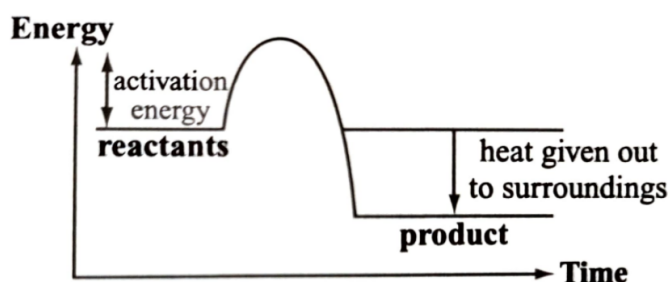
| Bond breaking | Bond forming |
|-----------------------------------------------------------------------------------------------------|----------------------------------------------------------------------------------------------------|
| Break 4 C – H bonds, 2 O = O bonds Amount of energy required = 4 (413) + 2 (496) = 2644 kJ | Form 2 C = O bonds, 4 O – H bonds Amount of energy released = 2 (805) + 4 (463) = 3462 kJ |

$$\Delta H = 2644 - 3462 = \underline{-818 \text{ kJ/mol}}$$

2 Hydrogen and chlorine react to produce hydrogen chloride.



(a) Draw an energy profile diagram for the reaction.



(b) What is the type of energy change occurring when bonds are broken?

Endothermic

(c) What is the type of energy change occurring when bonds are formed?

Exothermic

(d) Which is the larger energy change – bond breaking or bond forming? Explain.

Bond forming.

Forming the HCl bond releases more energy compared to the energy absorbed to break the bonds in H_2 and Cl_2 . The reaction is exothermic and there is a net release of energy.

(e) Find the energy change when 23.5 g of hydrogen gas reacts completely with chlorine gas.

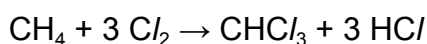
$$\text{No. of moles of H}_2 = \frac{23.5}{1 \times 2} = 11.75 \text{ mol}$$

$$\text{Energy change} = 11.75 \times (-184) = \underline{-2162 \text{ kJ}}$$

- 3 The bond energies of some bonds are given in the table below.

| Bond | H-H | C-Cl | H-Cl | C-C | C-H | C-Cl |
|------------------------|-----|------|------|-----|-----|------|
| Bond energy / (kJ/mol) | 436 | 243 | 432 | 348 | 413 | 330 |

The equation for the reaction between methane (CH_4) and chlorine (Cl_2) to produce chloroform (CHCl_3) and hydrogen chloride (HCl) is shown below.



- (a) Calculate the overall enthalpy change of the above reaction.

Total energy taken in to break 3 C-H and 3 C-Cl bonds
 $= 4 (413) + 3 (243) = 2381 \text{ kJ}$

Total energy released during bond forming of 3 C-C and 3 H-Cl bonds
 $= 3 (330) + 3 (432) = 2286 \text{ kJ}$

Overall enthalpy change

$= \text{Total energy absorbed during bond breaking} + \text{Total energy released during bond forming}$
 $= (+2381) + (-2699) = \underline{\underline{-318 \text{ kJ/mol}}}$

(mole ratio in equation is reactants 1:3, products 1:3, multiply accordingly to no. of moles in question)

This is a substitution reaction of alkane.

H in the hydrocarbon is substituted for Cl / Br / F

In this case, only 3 H are substituted for Cl, so 3 C-H bonds are broken and 3 C-Cl bonds are formed.

- (b) Hence, deduce whether the reaction is endothermic or exothermic.

The overall enthalpy change of the reaction has a negative value.

This indicates that the reaction is exothermic.

- 4 When 1.0 mol of ethanol, $\text{CH}_3\text{CH}_2\text{OH}$, is completely burnt in oxygen. 1367 kJ of heat is liberated.

(a) In terms of bond breaking and bond forming, explain why the combustion of ethanol is exothermic.

More energy is released during bond forming in products 2 CO_2 and 3 H_2O than energy is absorbed during bond breaking in reactants $\text{CH}_3\text{CH}_2\text{OH}$ and 3 O_2 .

(b) What is the application of this reaction in the motor industry?

The large amount of energy released is used to generate electricity for engines to operate to power vehicles (FINAL OUTCOME).

(c) Sketch the energy profile diagram for the combustion of ethanol.



(d) Calculate the mass of ethanol required to produce 411 kJ of heat.

$$\text{No. of moles of ethanol} = \frac{411 \text{ kJ}}{1367 \text{ kJ/mol}} = 0.30066 \text{ mol}$$

Mass of ethanol

$$= 0.30066 \text{ mol} \times [2(12) + 6(1) + 1(16)] \text{ g/mol}$$

$$= 0.30066 \times 46 = \underline{\underline{13.8 \text{ g}}} \text{ (3 s.f.)}$$

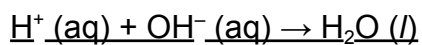
(e) When 1.0 mol of propanol, $\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$, is burnt completely in oxygen, x kJ of heat is liberated. Would you expect the value of x to be greater or smaller than 1367?

The value of x is greater than 1367.

Each molecule of propanol contains more carbon (C) and hydrogen (H) atoms than ethanol. Combustion produces more products (CO_2 and H_2O molecules), so more heat is liberated.

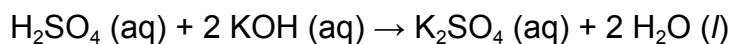
- 5 When 1.0 mol of hydrochloric acid reacts with 1.0 mol of aqueous sodium hydroxide, 57 kJ of heat energy is given out.

(a) Write the ionic equation for the reaction.



- (b) Deduce the heat energy given out when 100 cm³ of 2.0 mol/dm³ sulfuric acid reacts with 200 cm³ of 2.0 mol/dm³ aqueous potassium hydroxide.

(Calculate number of moles of water formed)



No. of moles of H₂SO₄ = 0.100 × 2.0 = 0.200 mol

No. of moles of KOH = 0.200 × 2.0 = 0.400 mol

No. of moles of H₂O = $\frac{2}{1} \times 0.200 = 0.400$ mol

(Same ionic equation)

∴ Heat energy released = 57 kJ/mol × 0.400 mol = **22.8 kJ**