5.1 ENTHALPY CHANGE

🎯 Learning outcomes

Candidates should be able to:

1. understand that chemical reactions are accompanied by enthalpy changes and these changes can be exothermic (ΔH is negative) or endothermic (ΔH is positive)
2. construct and interpret a reaction pathway diagram, in terms of the enthalpy change of the reaction and of the activation energy
3. define and use the terms:
   (a) standard conditions (this syllabus assumes that these are 298 K and 101 kPa) shown by $\ominus$.
   (b) enthalpy change with particular reference to: reaction, ΔH_r, formation, ΔH_f, combustion, ΔH_c, neutralisation, ΔH_{neut}
4. understand that energy transfers occur during chemical reactions because of the breaking and making of chemical bonds
5. use bond energies (ΔH positive, i.e. bond breaking) to calculate the enthalpy change of reaction, ΔH_r
6. understand that some bond energies are exact and some bond energies are averages
7. calculate enthalpy changes from appropriate experimental results, including the use of the relationships $q = mc\Delta T$ and $\Delta H = -mc\Delta T/n$

Initial assessment

Give some examples of exothermic reactions and endothermic reactions from your daily life.

Introduction

Almost all chemical reactions are accompanied by energy changes as bonds are broken and formed.

Any chemical reaction will involve breaking some bonds and making new ones. Energy is needed to break bonds and is given out when new bonds are formed. It is improbable that these two processes will involve exactly the same amount of energy - and so some energy will either be absorbed or released during a reaction.

Usually, the energy changes involve heat, but they can also involve sound, light or even electrical energy.

Reactions that **give out heat** are called **exothermic** reactions.

Reactions that **take in energy** are known as **endothermic** reactions.

Enthalpy change, ΔH: the heat energy transferred during a chemical reaction

For an exothermic reaction, heat is lost from the reactants. The value of **ΔH** is negative.
For an endothermic reaction, heat is gained by the reactants. The value of **ΔH** is positive.
Enthalpy changes are measured in kJ mol^{-1}.

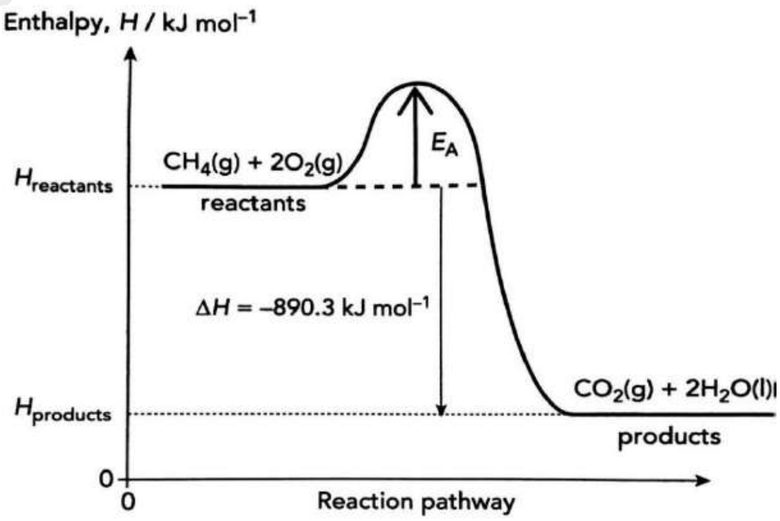


Figure: Energy change in an exothermic reaction

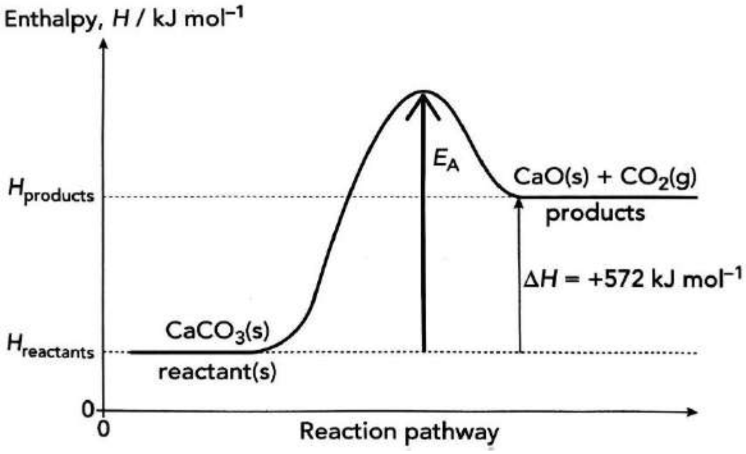


Figure: Energy change in an endothermic reaction

Activation energy, E_A: the minimum energy the colliding particles must possess to break bond to start a chemical reaction.

When drawing reaction pathway diagrams, the activation energy arrow should start at the level of the reactants and end at the highest point on the energy hump.
When drawing the arrow for ΔH make sure that the arrow points downwards for an exothermic reaction and upwards for an endothermic reaction.

Standard conditions

To make any comparison of enthalpy changes a fair comparison, we must use the same conditions, called **standard conditions**. These are:
a pressure of 101 kPa (1.01×10^5 Pa, approximately normal atmospheric pressure)
a temperature of 298 K (25 °C)
each substance involved in the reaction is in its normal physical state (solid, liquid or gas) at 101 kPa and 298 K.
The symbol ⦵(a superscript Plimsoll symbol) is used to indicate that the enthalpy change refers to a reaction carried out under standard conditions.

Enthalpy Change

A standard enthalpy change **refers to the energy transferred at 298 K and standard pressure (usually 1 atmosphere or 101 kPa).**

Standard enthalpy change of reaction, ΔH_r
The standard enthalpy change of reaction is the enthalpy change **when the amounts of reactants shown in the stoichiometric equation react to give products** under standard conditions. The reactants and products must be in their standard states.

Standard enthalpy change of formation, ΔH_f
The standard enthalpy change of formation is the enthalpy change **when one mole of a compound is formed from its elements** under standard conditions. The reactants and products must be in their standard states.
By definition, the standard enthalpy change of formation of any element in its standard state is zero.

Standard enthalpy change of combustion, ΔH_c

The standard enthalpy change of combustion is the enthalpy change **when one mole of a substance is burnt in excess oxygen** under standard conditions. The reactants and products must be in their standard states.

Standard enthalpy change of neutralisation, ΔH_{neut}

The standard enthalpy change of neutralisation is the enthalpy change **when one mole of water is formed by the reaction of an acid with an alkali** under standard conditions.

Use of bond energies to calculate the enthalpy change of reaction, ΔH_r

(a) Bond energies in diatomic molecules (these are exact values)

Homonuclear

Bond	Energy/kJ mol^{-1}
H-H	436
N≡N	944
O=O	496
P≡P	485
S=S	425
F-F	158
Cl-*Cl*	242
Br-Br	193
I-I	151

Heteronuclear

Bond	Energy/kJ mol^{-1}
H-F	562
H-*Cl*	431
H-Br	366
H-I	299
C≡O	1077

(b) Bond energies in polyatomic molecules (these are average values)

Homoatomic

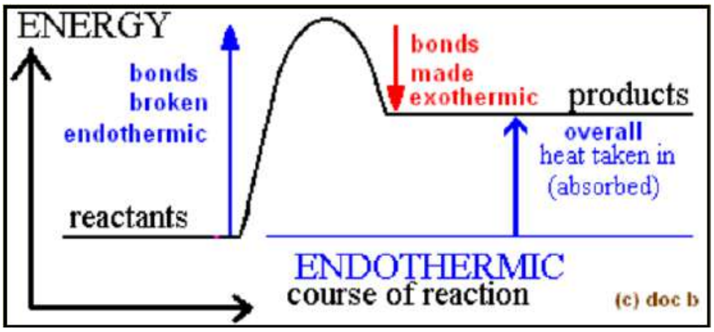
Bond	Energy /kJ mol^{-1}
C-C	350
C=C	610
C≡C	840
C C (benzene)	520
N-N	160
O-O	150
Si-Si	225
P-P	200
S-S	265

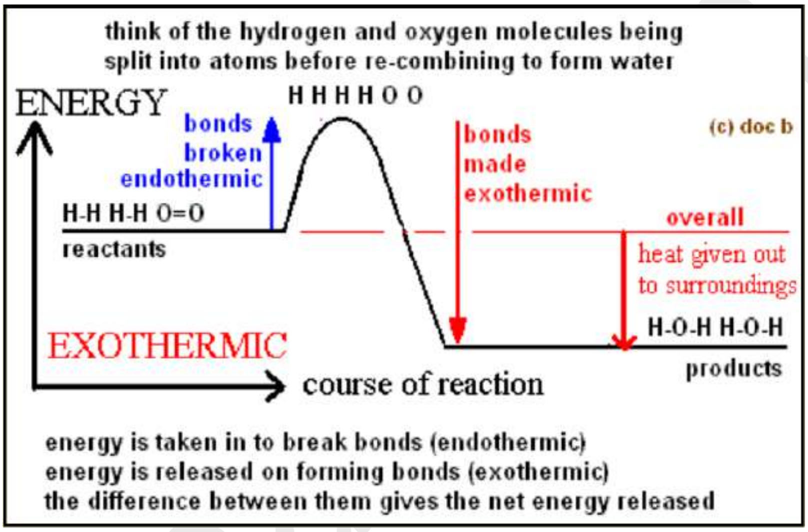
Bond	Energy/kJ mol^{-1}
C-H	410
C-*Cl*	340
C-Br	280
C-I	240
C-N	305
C=N	610
C≡N	890
C-O	360
C=O	740
C=O in CO2	805
N-H	390
N-*Cl*	310
O-H	460
Si-*Cl*	360
Si-H	320
Si-O (in SiO2(s))	460
Si=O (in SiO2(g))	640
P-H	320
P-*Cl*	330
P-O	340
P=O	540
S-H	340
S-*Cl*	250
S-O	360
S=O	500

Enthalpy changes are due to the breaking and forming of bonds. Breaking bonds requires energy. The energy is needed to overcome the attractive forces joining the atoms together. Energy is released when new bonds are formed. Bond breaking is endothermic and bond forming is exothermic.

In a chemical reaction:

- if the energy needed to break bonds is **less** than the energy released when new bonds are formed, the reaction will release energy and is exothermic.
- if the energy needed to break bonds is **more** than the energy released when new bonds are formed, the reaction will absorb energy and is endothermic.





GCSE Chemistry Revision "Bond Energy Calculations" - YouTube
GCSE Chemistry Revision "Bond Energy Calculations 2" - YouTube

Measuring enthalpy changes

The energy released in calorimetry experiments for a known number of moles of a specified reactant is given by **q = mcΔT**.

q is the heat transferred, in J
m is the mass of water, in g
c is the specific heat capacity, in J g^{-1} $°C^{-1}$
ΔT is the temperature change, in °C

When we multiply this up to 1 mole of a stated product or reactant, the equation we use is **ΔH = -mcΔT**
The negative sign shows that the reaction is exothermic.

Specific heat capacity, c: the energy needed to raise the temperature of 1 g of substance by 1°C (by 1 K)

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