Chapter 14

# Energy Changes in Chemical Reactions - Grade 11

All chemical reactions involve energy changes. In some reactions we barely notice an energy change, but in others we are able to see these energy changes by either an increase or a decrease in the overall energy of the system.

14.1 What causes the energy changes in chemical reactions?

When a chemical reaction occurs, bonds in the reactants break, while new bonds form in the product. The following example may help to explain this.

Hydrogen reacts with oxygen to form water, according to the following equation:

2 *H*2 + *O*2 → 2 *H*2*O*

In this reaction, the single bond between the two hydrogen atoms in the H2 molecule will break, as will the double bond in the O2 molecule. New bonds will form between the two hydrogen atoms and the single oxygen atom in the water molecule that is formed as the product.

For bonds to break, energy must be absorbed. When new bonds form, energy is released. The energy that is needed to break a bond is called the bond enthalpy or bond energy. Bond enthalpies are measured in units of kJ.mol−1.

Definition: Bond enthalpy is a measure of bond strength in a chemical bond. It is the amount of energy (in kJ.mol−1) that is needed to break the chemical bond between two atoms, and the amount of energy that is released when a new bond is formed between two atoms.

## 14.2 Exothermic and endothermic reactions

In some reactions, the energy that must be absorbed to break the bonds in the reactants, is less than the total energy that is released when new bonds are formed. This means that in the overall reaction, energy is released as either heat or light. This type of reaction is called an exothermic reaction. Another way of describing an exothermic reaction is that it is one in which the energy of the product is less than the energy of the reactants, because energy has been released during the reaction. We can represent this using the following general formula:

Reactants → Product + Energy

Definition: An exothermic reaction is one that releases energy in the form of heat or light.

In other reactions, the energy that must be absorbed to break the bonds in the reactants, is more than the total energy that is released when new bonds are formed. This means that in the overall reaction, energy must be absorbed from the surroundings. This type of reaction is known as an endothermic reaction. Another way of describing an endothermic reaction is that it is one in which the energy of the product is greater than the energy of the reactants, because energy has been absorbed during the reaction. This can be represented by the following formula:

Reactants + Energy → Product

Definition: An endothermic reaction is one that absorbs energy in the form of heat.

The difference in energy (E) between the reactants and the products is known as the heat of the reaction. It is also sometimes referred to as the enthalpy change of the system.

## 14.3 Change in Enthalpy (∆H)

Definition: Enthalpy is the heat content of a chemical system, and is given the symbol ’H’.

The Change in Enthalpy for a reaction is sometimes called the Heat of reaction (as in “heat of vaporisation, Heat of solvation, etc.). There are three ways to calculate the change in enthalpy, two will be covered in this course.

* In an exothermic reaction, ∆*H* is less than zero because the energy of the reactants is greater than the energy of the product. For example,

C(s) + O2(g) → CO2(g) ∆H = -393 kJ OR C(s) + O2(g) → CO2(g) + 393 kJ

* In an endothermic reaction, ∆*H* is greater than zero because the energy of the reactants is less than the energy of the product. For example,

C(s) + H2O(g) → H2(g) + CO(g) ∆H = +131 kJ or C(s) + H2O(g) + 131 kJ → CO(g) + H2(g)

Some of the information relating to exothermic and endothermic reactions is summarised in table 14.1.

|  |  |  |  |
| --- | --- | --- | --- |
|  Type of reaction | Exothermic | Endothermic |  |
|  Energy absorbed or re-leased | Released | Absorbed |  |
|  Surroundings - hotter or cooler? | hotter | cooler |  |
|  Sign of ∆H | Negative | Positive |  |

Table 14.1: A comparison of exothermic and endothermic reactions

The **units** for ∆H are kJ. In other words, the ∆H value gives the amount of energy that is absorbed or released per in the reaction as it is written. When the change in enthalpy is quoted for a particular substance (usually a product), it is often quoted as kJ/mole, meaning the kJ released or absorbed per 1 mole of that substance. Sometimes this change in enthalpy per mole is not the same value as the change in enthalpy with the equation because there is more than one moles of the substance in the equation. For example – the heat of formation of water is -286kJ/mole, and this is shown in the equation as

*2 H*2 + O2 → 2 *H2O* ∆H = -572 kJ

**Exercise: Endothermic and exothermic reactions**

1. In each of the following reactions, say whether the reaction is endothermic or exothermic, and give a reason for your answer.
	* 1. H2 + I2 → 2HI + 21kJ
		2. CH4 + 2O2 → CO2 + 2H2O ∆ H = -802 kJ
		3. The following reaction takes place in a flask: Ba(OH)2.8H2O + 2NH4NO3 → Ba(NO3)2 + 2NH3 + 10H2O. Within a few minutes, the temperature of the flask drops by approximately 20C.
		4. Na + Cl2 → 2NaCl ∆H = -411 kJ
		5. C + O2 → CO2
2. For each of the following descriptions, say whether the process is endothermic or exothermic and give a reason for your answer.
	* 1. evaporation
		2. the combustion reaction in a car engine
		3. bomb explosions
		4. melting ice
		5. digestion of food
		6. condensation

## 14.6 Activation energy and the activated complex

Reactions will not take place until the system has some minimum amount of energy added to it. This amount energy is called the activation energy.

It is possible to draw an energy diagram to show the energy changes that take place during a particular reaction. Let’s consider an example:

*H*2(*g*) + *F*2(*g*) → 2*HF*(*g*)

Figure 14.1: The energy changes that take place during an exothermic reaction

[

](

activated complex)

2

HF

products

H

2

+

F

2

reactants

activation

energy

∆H=

−

268

k.J.mol

−

1

Time

Potentialenergy

Bonds broken

Bonds formed

The reaction between *H*2(*g*) and *F*2(*g*) (figure 14.1) needs energy in order to proceed, to break the necessary bonds in the reactants, and this is the activation energy. Once the reaction has started, a transition state is reached where the two reactants are in an in-between state. The energy that is needed to reach this state is equal to the activation energy for the reaction. The compound that is formed in this transition state is called the activated complex. The transition state lasts for only a very short time, after which either the original bonds reform (and the reaction does not proceed), or new bonds form and a new product forms. In this example, less energy was needed to break the bonds than the energy released when the new bonds formed. The reaction is exothermic and ∆H is negative.

In endothermic reactions, more energy is need to break the bonds in the reactants than is released when new bonds form to make the products.

[

XYZ

]

X+YZ

products

XY+Z

reactants

∆H

*>*

0

Time

Potentialenergy

activation

energy

An energy diagram is shown below (figure 14.2) for the endothermic reaction *XY* +*Z* → *X* +*Y Z*. In this example, the activated complex has the formula XYZ.

Figure 14.2: The energy changes that take place during an endothermic reaction

**Exercise: Summary Exercise**

1. For each of the following, say whether the statement is true or false. If it is false, give a reason for your answer.

* 1. Energy is released in all chemical reactions.
	2. The condensation of water vapour is an example of an endothermic reaction.
	3. In an exothermic reaction ∆H is less than zero. (d) All non-spontaneous reactions are endothermic.
1. For each of the following, choose the one correct answer.
	1. For the following reaction:

A + B ⇔ AB ∆H = -129 kJ.mol−1

i. The energy of the reactants is less than the energy of the product.

ii. The energy of the product is less than the energy of the reactants.

* + 1. The reaction is non-spontaneous.
		2. The overall energy of the system increases during the reaction.

**Activity :: Demonstration : Endothermic and exothermic reactions 1 Apparatus and materials:**

You will need citric acid, sodium bicarbonate, a glass beaker, the lid of an icecream container, thermometer, glass stirring rod and a pair of scissors. Note that citric acid is found in citrus fruits such as lemons. Sodium bicarbonate is actually bicarbonate of soda (baking soda), the baking ingredient that helps cakes to rise. Method:

1. Cut a piece of plastic from the ice-cream container lid that will be big enoughto cover the top of the beaker. Cut a small hole in the centre of this piece of plastic and place the thermometer through it.
2. Pour some citric acid (H3C6H5O7) into the glass beaker, cover the beaker with its ’lid’ and record the temperature of the solution.
3. Stir in the sodium bicarbonate (NaHCO3), then cover the beaker again.
4. Immediately record the temperature, and then take a temperature reading everytwo minutes after that. Record your results in a table like the one below.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Time (mins) | 0 | 2 | 4 | 6 |
| Temperature (0C) |  |  |  |  |

The equation for the reaction that takes place is:

*H*3*C*6*H*5*O*7(*aq*) + 3*NaHCO*3(*s*) → 3*CO*2(*g*) + 3*H*2*O*(*l*) + *NaC*6*H*5*O*7(*aq*)

Results:

* Plot your temperature results on a graph of temperature against time. What happens to the temperature during this reaction?
* Is this an exothermic or an endothermic reaction?
* Why was it important to keep the beaker covered with a lid?
* Do you think a glass beaker is the best thing to use for this experiment? Explain your answer.
* Suggest another container that could have been used and give reasons for your choice. It might help you to look back to chapter ?? for some ideas!

**Activity :: Demonstration : Endothermic and exothermic reactions 2 Apparatus and materials:**

Vinegar, steel wool, thermometer, glass beaker and plastic lid (from previous demonstration). Method:

1. Put the thermometer through the plastic lid, cover the beaker and record thetemperature in the empty beaker. You will need to leave the thermometer in the beaker for about 5 minutes in order to get an accurate reading.
2. Take the thermometer out of the jar.
3. Soak a piece of steel wool in vinegar for about a minute. The vinegar removes the protective coating from the steel wool so that the metal is exposed to oxygen.
4. After the steel wool has been in the vinegar, remove it and squeeze out anyvinegar that is still on the wool. Wrap the steel wool around the thermometer and place it (still wrapped round the thermometer) back into the jar. The jar is automatically sealed when you do this because the thermometer is through the top of the lid.
5. Leave the steel wool in the beaker for about 5 minutes and then record thetemperature. Record your observations.

Results:

You should notice that the temperature increases when the steel wool is wrapped around the thermometer.

Conclusion:

The reaction between oxygen and the exposed metal in the steel wool, is exothermic, which means that energy is released and the temperature increases.

**Activity :: Investigation : Endothermic and exothermic reactions Apparatus and materials:**

Approximately 2 g each of calcium chloride (CaCl2), sodium hydroxide (NaOH), potassium nitrate (KNO3) and barium chloride (BaCl2); concentrated sulfuric acid (H2SO4); 5 test tubes; thermometer. Method:

1. Dissolve about 1 g of each of the following substances in 5-10 cm3 of water in a test tube: CaCl2, NaOH, KNO3 and BaCl2.
2. Observe whether the reaction is endothermic or exothermic, either by feelingwhether the side of the test tube gets hot or cold, or using a thermometer.
3. Dilute 3 cm3 of concentrated H2SO4 in 10 cm3 of water in the fifth test tube and observe whether the temperature changes.
4. Wait a few minutes and then add NaOH to the H2SO4. Observe any energy changes.
5. Record which of the above reactions are endothermic and which are exothermic.

Results:

* When BaCl2 and KNO3 dissolve in water, they take in heat from the surroundings. The dissolution of these salts is endothermic.
* When CaCl2 and NaOH dissolve in water, heat is released. The process is exothermic.
* The reaction of H2SO4 and NaOH is also exothermic.