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Reactivity 1.1

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Heat and temperature

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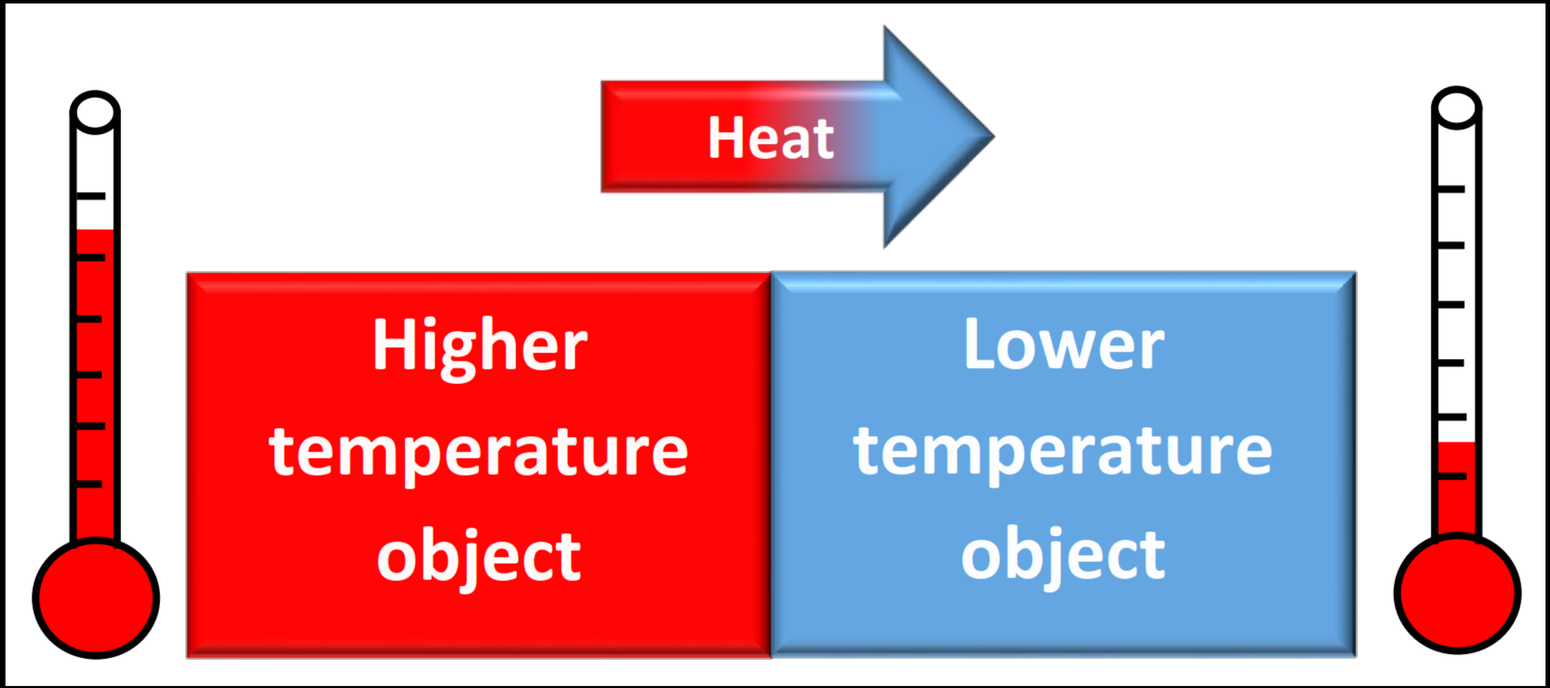
Heat

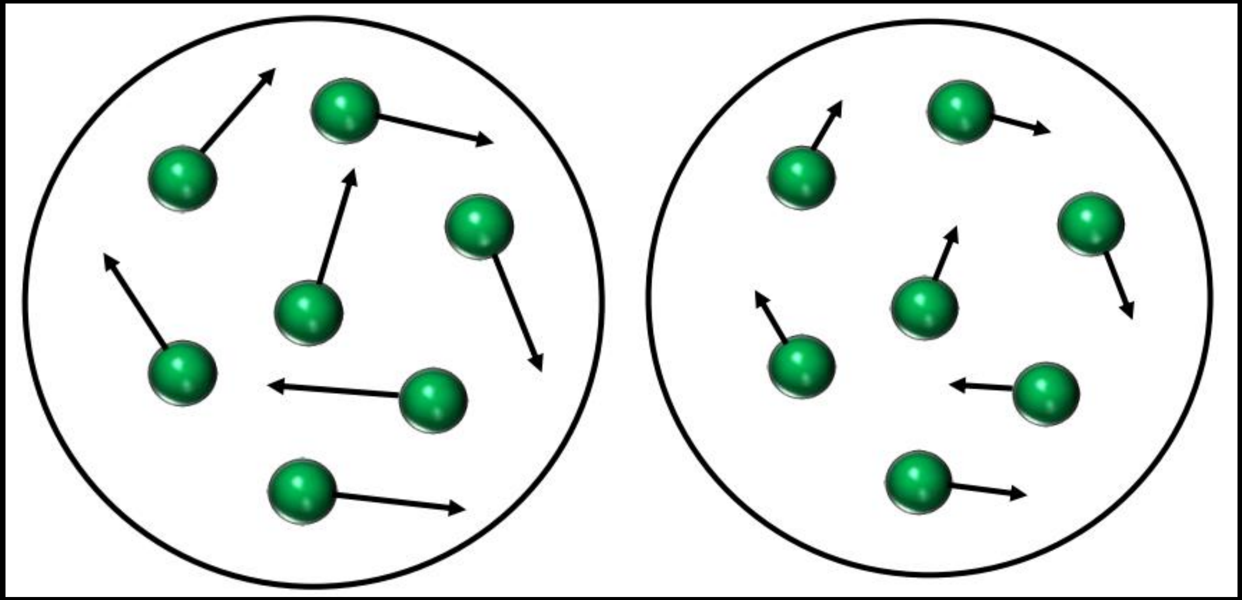
Heat is a form of energy. Heat always transfers or flows from a higher temperature object to a lower temperature object.



Absolute temperature

The absolute temperature in Kelvin (K) is directly proportional to the average kinetic energy of the particles in a substance.

Gas at 300 K particles have higher average kinetic energy



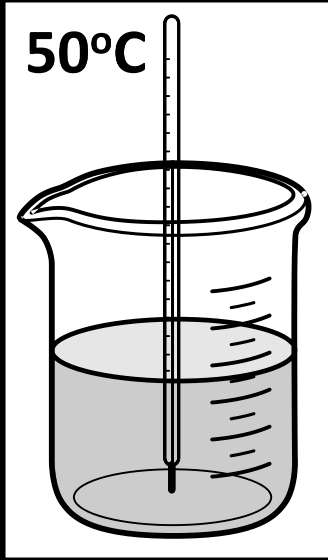
Gas at 100 K particles have lower average kinetic energy

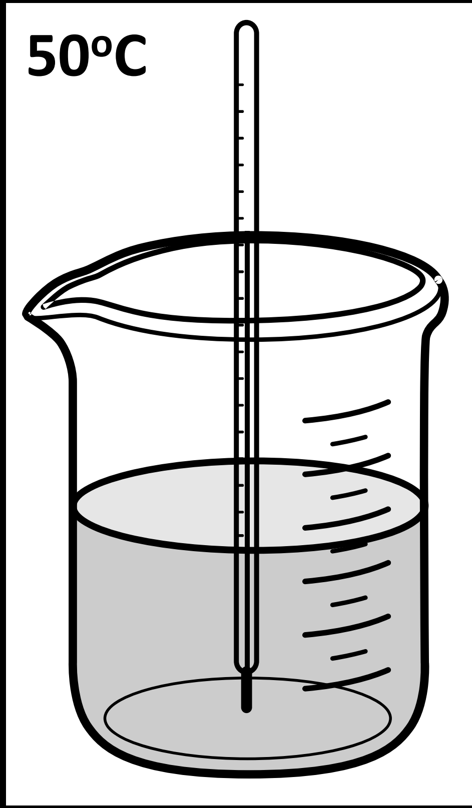
Heat vs temperature

Heat is a measure of the total energy of a substance and therefore depends on the amount of substance present.

Temperature is a measure of the average kinetic energy of the particles in a substance.

Both beakers of water have the same average kinetic energy but the bigger beaker of water has more heat.





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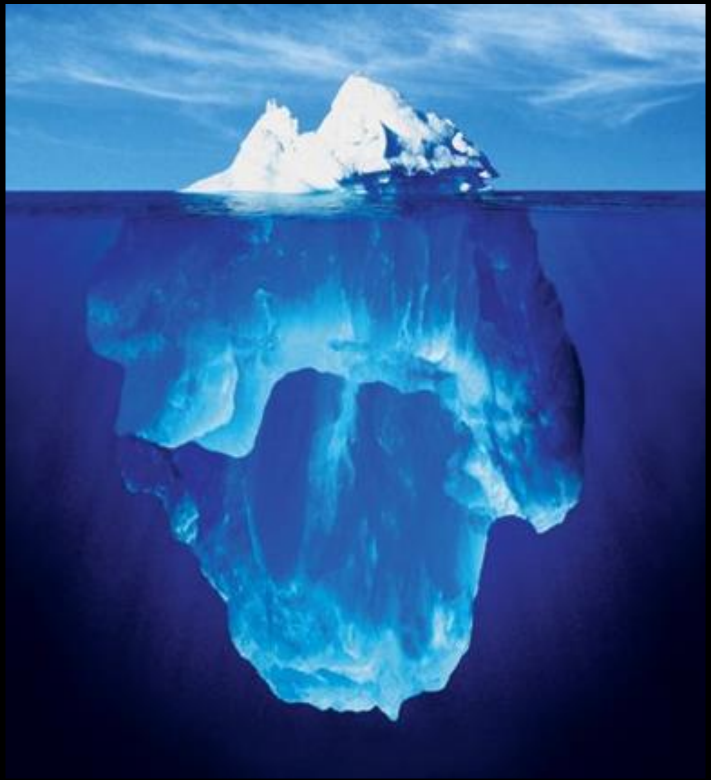
Heat vs temperature



Which has more heat?

Which has the higher temperature?

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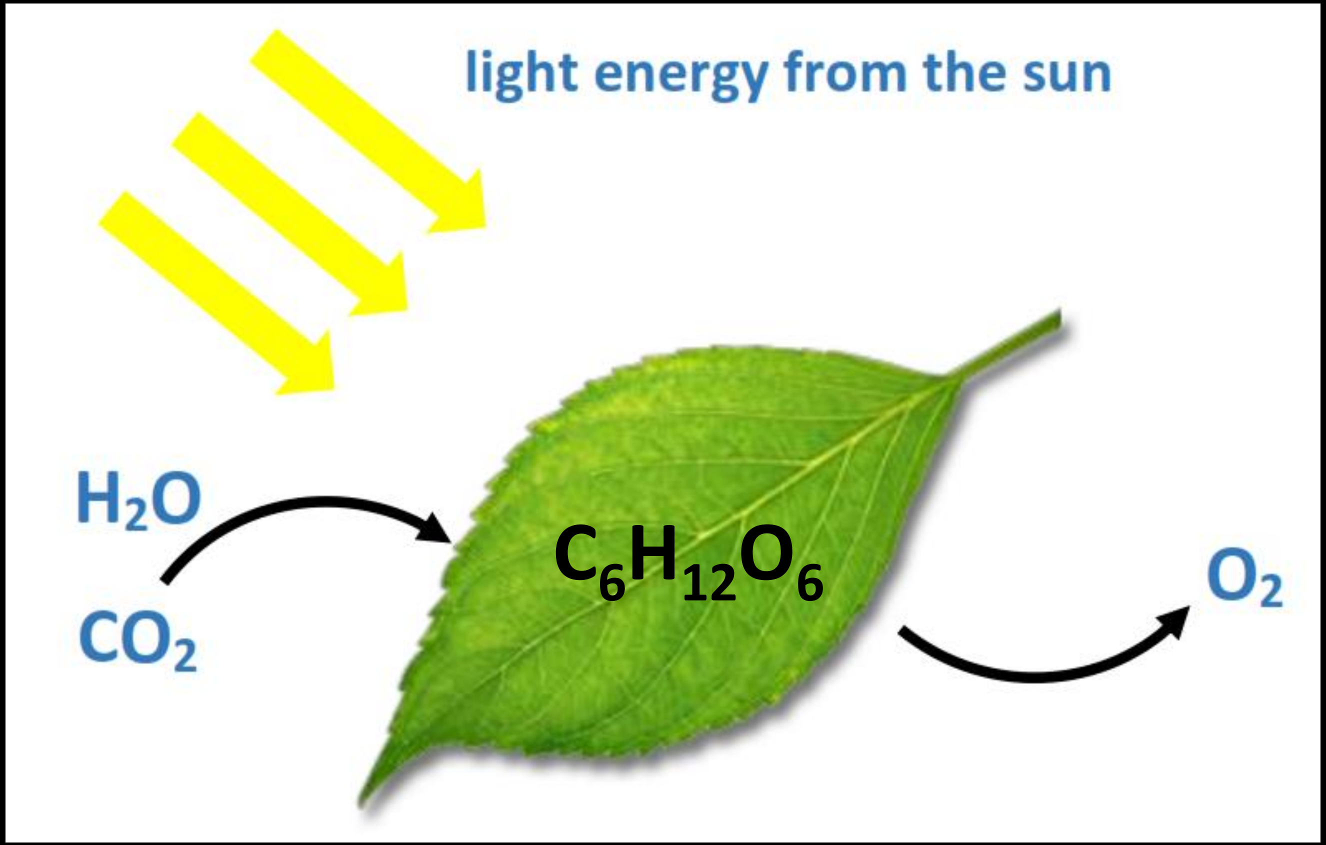
Enthalpy (*H*) and enthalpy change (Δ*H*)

Enthalpy (H) and enthalpy change (ΔH)

Chemical reactions either release heat (exothermic) or absorb heat (endothermic).



Enthalpy (*H*) and enthalpy change (Δ*H*)

For a chemical reaction that takes place at constant pressure, the heat absorbed or released is known as enthalpy (*H*).

Enthalpy can be thought of as the heat content of a substance.

The enthalpy of a substance cannot be measured – only a change in enthalpy can be measured.

The term ΔH is the change in enthalpy that occurs when a chemical reaction takes place.

Enthalpy (*H*) and enthalpy change (Δ*H*)

Standard enthalpy change of reaction is the enthalpy change of a reaction carried out under standard conditions (100 kPa) with everything in its standard state.

$$\Delta H^{\ominus}$$

The standard state is the normal, most pure stable state of a substance measured at a pressure of 100 kPa.

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Exothermic and endothermic reactions

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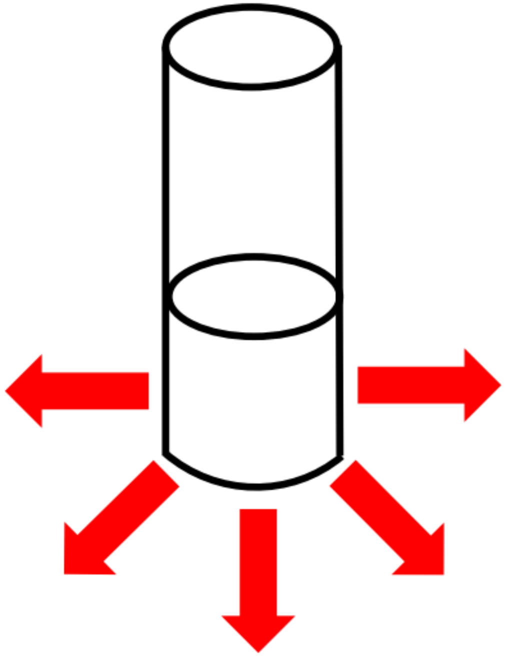
System and surroundings

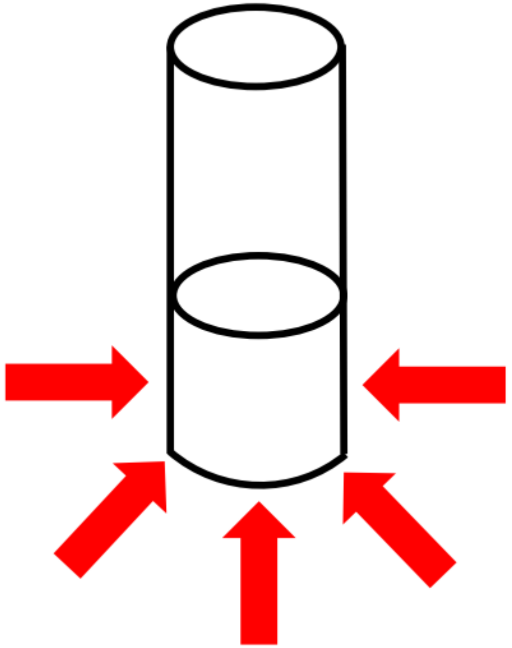
Exothermic



Heat flows from the system to the surroundings

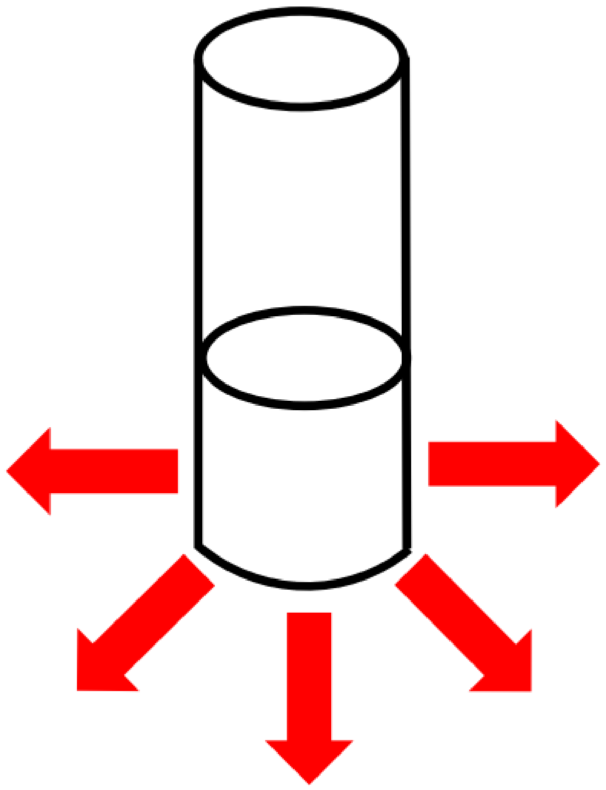
The system is the reactants and products (the chemical reaction). The surroundings consist of everything else outside of the system (including the reaction mixture).

Endothermic



Heat flows from surroundings into the system

Exothermic reactions



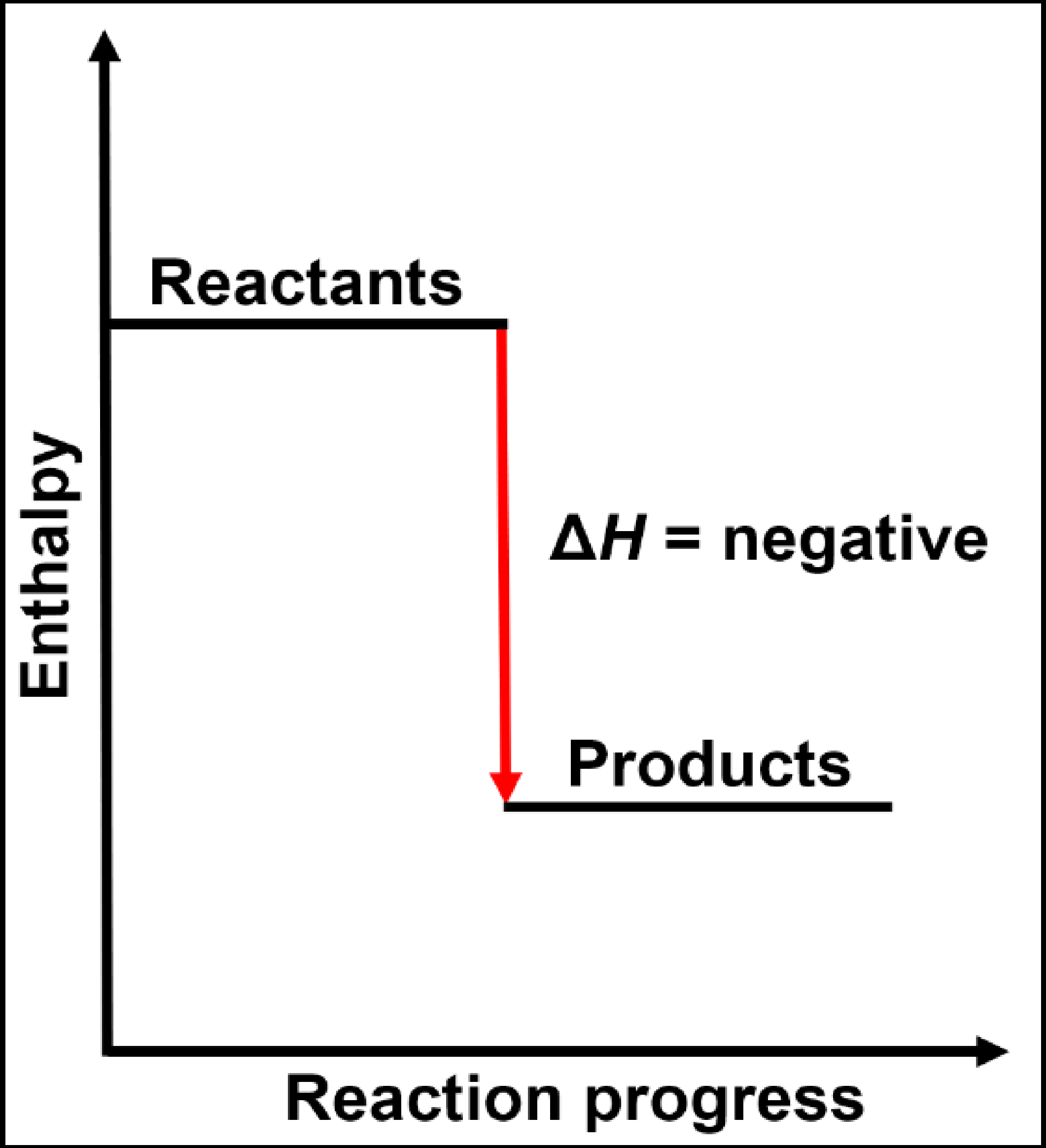
Heat flows from the system to the surroundings

In an exothermic reaction, heat flows from the system to the surroundings.

The temperature of the reaction mixture and the surroundings increases as heat is released.

Examples of exothermic reactions include combustion and neutralisation.

Exothermic reactions



ΔH is negative

Products have lower enthalpy than reactants.

Products are more energetically stable than reactants.

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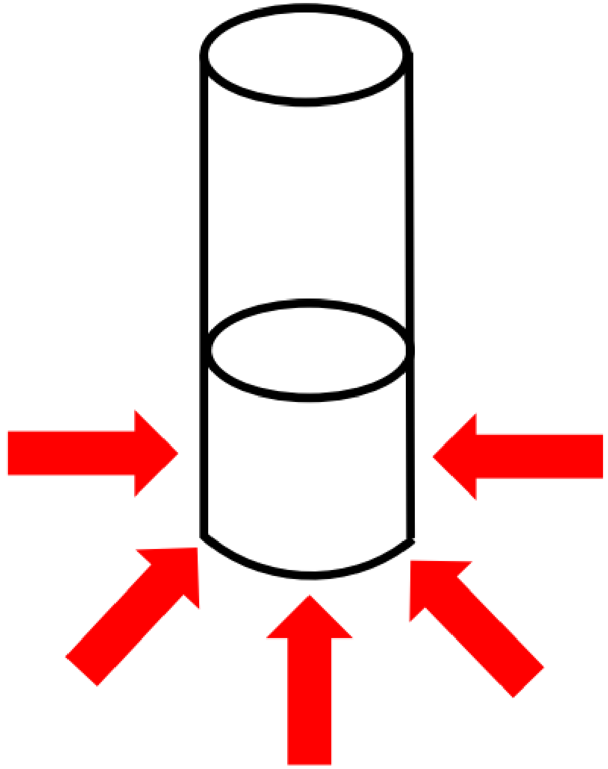
Endothermic reactions



Heat flows from surroundings into the system

In an endothermic reaction, heat flows from the surroundings to the system.

The temperature of the reaction mixture and the surroundings decreases as heat is absorbed.

Examples of endothermic reactions include photosynthesis and thermal decomposition reactions.

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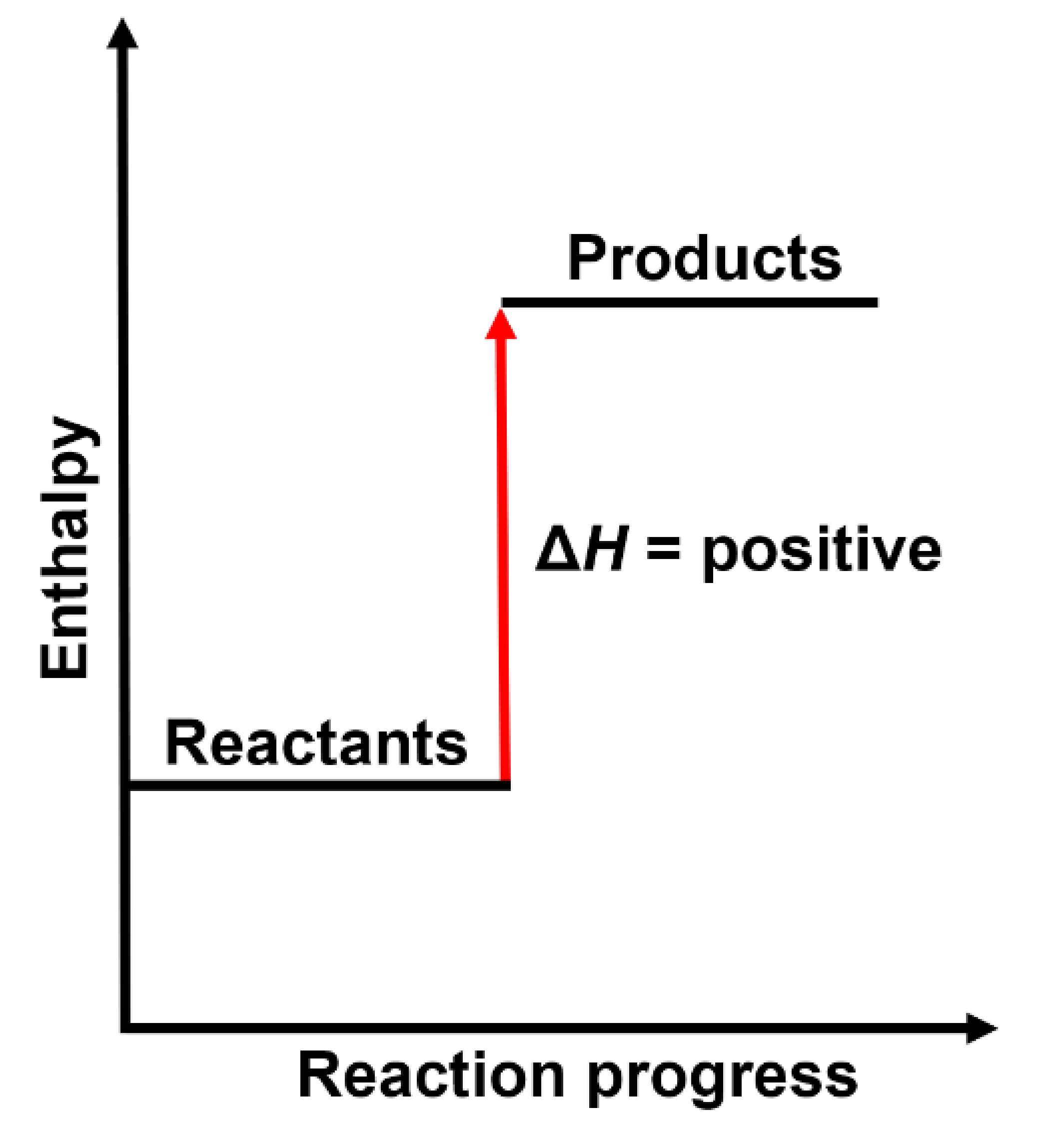
Endothermic reactions



Δ*H* is positive

Reactants have lower enthalpy than products.

Reactants are more energetically stable than products.

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Summary

	Exothermic reaction	Endothermic reaction
Enthalpy change (ΔH)	negative	positive
Which has lower enthalpy?	products	reactants
Which is more energetically stable?	products	reactants
Heat absorbed or released?	released	absorbed

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Calculating ΔH using q = mcΔT

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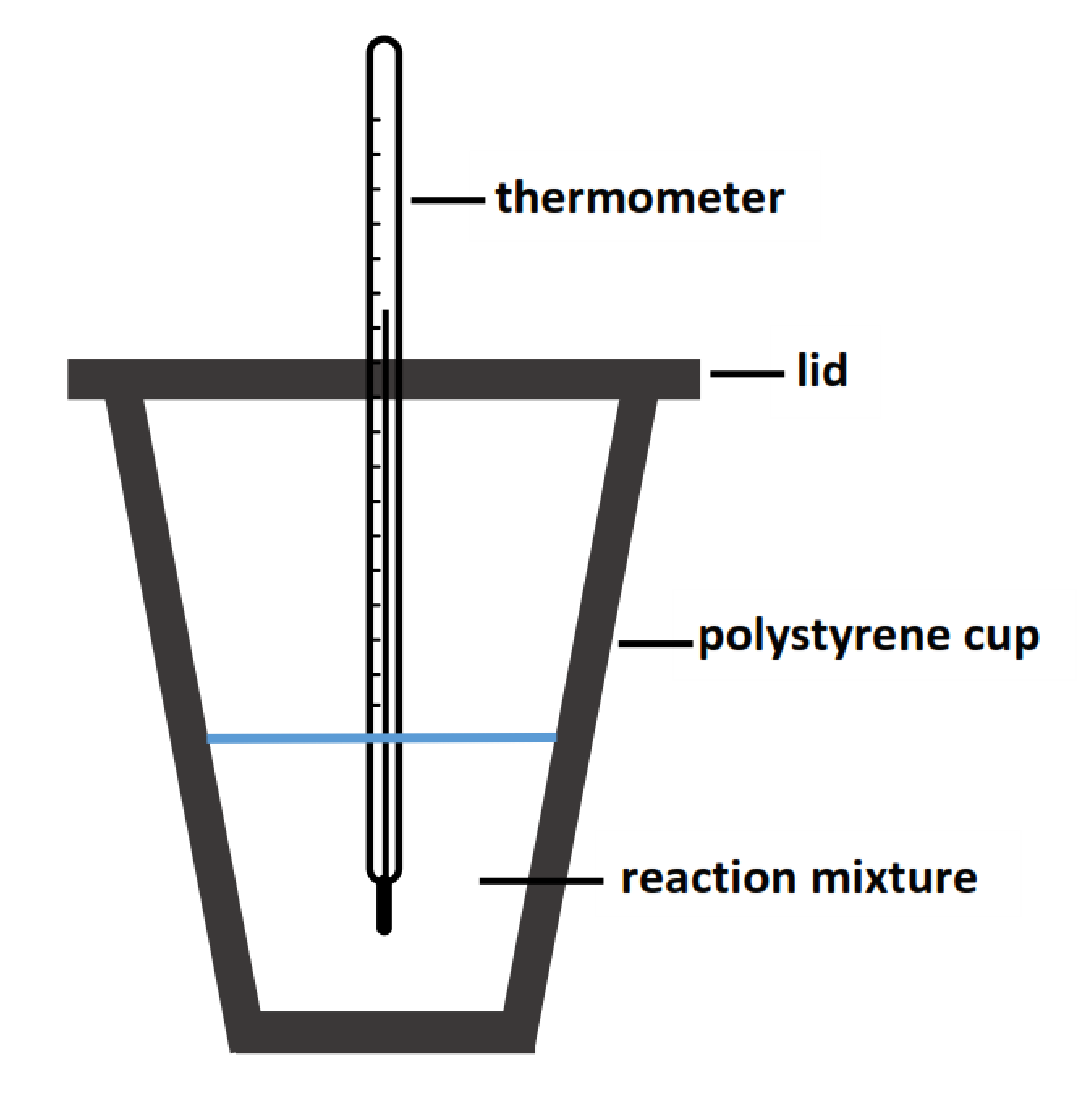
Calculating ΔH using q = mcΔT



Enthalpy changes can be measured using a simple calorimeter.

The heat absorbed or released raises or lowers the temperature of the solution.

$$q = mc\Delta T$$

Calculating ΔH using $q = mc\Delta T$

$$q = mc\Delta T$$

q – heat (J)

m – mass (g)

c – specific heat capacity ($J\ g^{-1}\ {}^{o}C^{-1}$)

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Calculating ΔH using q = mcΔT

Specific heat capacity is the amount of heat required to raise one gram (or kilogram) of a substance by one °C (or K).

Substance	Specific heat capacity ($J\ g^{-1}\ {}^{o}C^{-1}$)
Water	4.18
Copper	0.390
Aluminium	0.900
Iron	0.450

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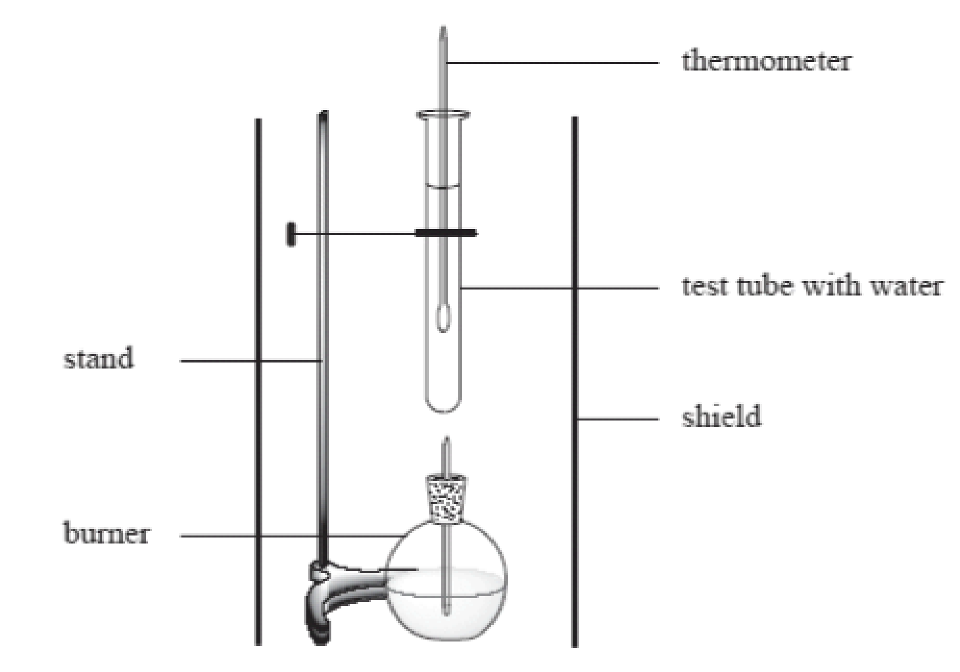
Calculating ΔH using $q = mc\Delta T$



Initial mass of burner and methanol ±0.01g	80.55
Final mass of burner and methanol ±0.01g	80.03
Mass of water in test tube ±0.01g	20.00
Initial temperature of water ±0.5°C	19.5
Final temperature of water ±0.5°C	56.3

m (CH_3OH) 80.55 – 80.03 = 0.52 g

n (CH_3OH) = m ÷ M = 0.52 ÷ 32.05 = 0.016 mol

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Calculating ΔH using q = mcΔT

$$q = mc\Delta T$$

$$q = 20.00 \times 4.18 \times (56.3 - 19.5)$$

$$q = -\ 3080\ J$$

$$q = -\ 3080 \div 0.016 = -\ 192500\ J$$

$$q = -\ 1.93 \times 10^{2}\ kJ\ mol^{-1}$$

Calculating ΔH using q = mcΔT

The literature value for the enthalpy change of combustion (ΔH_c) of methanol is –726 kJ mol^{-1}

$$\% \text{ error} = \frac{(-193)-(-726)}{(-726)} \times 100 = 74\%$$

Systematic errors:

Heat loss to the surroundings; not all heat transferred to the water; incomplete combustion of methanol.