5.1 Measuring energy changes

Nature of science:

Fundamental principle—conservation of energy is a fundamental principle of science. (2.6)

Making careful observations—measurable energy transfers between systems and surroundings. (3.1)

Understandings:

- Heat is a form of energy.
- Temperature is a measure of the average kinetic energy of the particles.
- Total energy is conserved in chemical reactions.
- Chemical reactions that involve transfer of heat between the system and the surroundings are described as endothermic or exothermic.
- The enthalpy change (ΔH) for chemical reactions is indicated in kJ mol^{-1}.
- ΔH values are usually expressed under standard conditions, given by $\Delta H°$, including standard states.

Applications and skills:

- Calculation of the heat change when the temperature of a pure substance is changed using $q = mc\Delta T$.
- A calorimetry experiment for an enthalpy of reaction should be covered and the results evaluated.

Guidance:

- Enthalpy changes of combustion ($\Delta H_c^\ominus$) and formation ($\Delta H_f^\ominus$)should be covered.
- Consider reactions in aqueous solution and combustion reactions.
- Standard state refers to the normal, most pure stable state of a substance measured at 100 kPa. Temperature is not a part of the definition of standard state, but 298 K is commonly given as the temperature of interest.
- The specific heat capacity of water is provided in the data booklet in section 2.
- Students can assume the density and specific heat capacities of aqueous solutions are equal to those of water, but should be aware of this limitation.
- Heat losses to the environment and the heat capacity of the calorimeter in experiments should be considered, but the use of a bomb calorimeter is not required.

International-mindedness:

- The SI unit of temperature is the Kelvin (K), but the Celsius scale (°C), which has the same incremental scaling, is commonly used in most countries. The exception is the USA which continues to use the Fahrenheit scale (°F) for all non-scientific communication.

Theory of knowledge:

- What criteria do we use in judging discrepancies between experimental and theoretical values? Which ways of knowing do we use when assessing experimental limitations and theoretical assumptions?

Utilization:

- Determining energy content of important substances in food and fuels.

Syllabus and cross-curricular links:
Topic 1.1—conservation of mass, changes of state
Topic 1.2—the mole concept

Aims:

- **Aim 6**: Experiments could include calculating enthalpy changes from given experimental data (energy content of food, enthalpy of melting of ice or the enthalpy change of simple reactions in aqueous solution).
- **Aim 7**: Use of databases to analyse the energy content of food.
- **Aim 7**: Use of data loggers to record temperature changes.

UNIT 5.1 – MEASURING ENERGY CHANGES

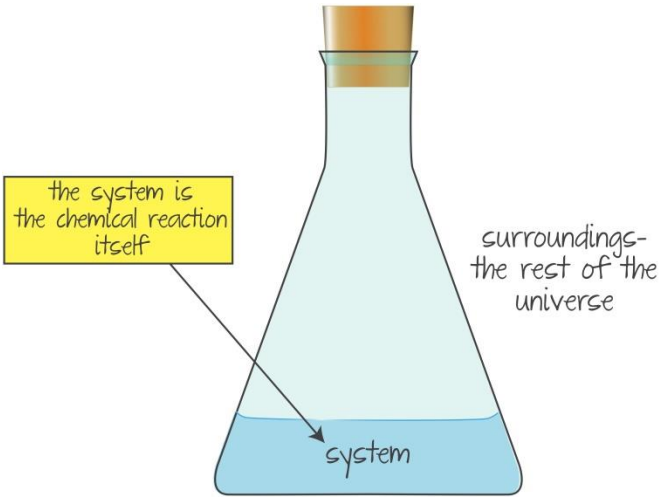
MEASURMENT OF HEAT ENERGY

Energy: The ability to do work

Chemical Energy: Energy released or absorbed during chemical reactions

Heat: A mode of energy transfer which occurs as a result of a temperature difference

SYSTEM AND SURROUNDINGS



Open System: Can exchange energy and matter with the surroundings.

Closed System: Can exchange energy, but not matter, with the surroundings.

The surroundings include:

- The solvent (in this case, water).
- The air around the test tube.
- The test tube itself.
- Anything dipping into the test tube (e.g. a thermometer).

Enthalpy: The heat content of a system, measured in kJ/mol
ΔH: The change in enthalpy of a system (exothermic or endothermic)

When heat is released from the system to the surroundings (when the temperature rises): **Exothermic**, ΔH is negative

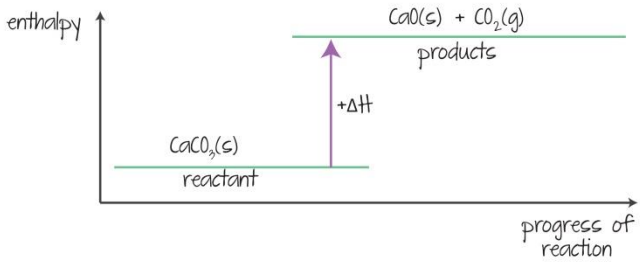
When heat is added to a system from the surroundings (when the temperature falls): **Endothermic**, ΔH is positive

TEMPERATURE

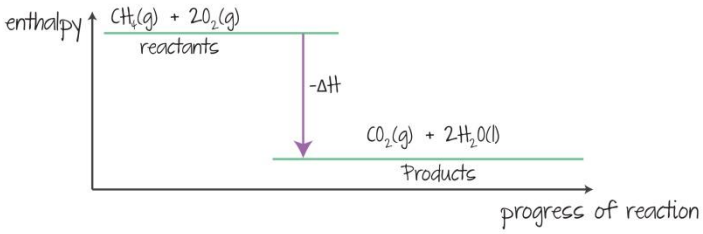
The averages kinetic energy of the particles of a substance.

EXOTHERMIC AND ENDOTHERMIC REACTIONS

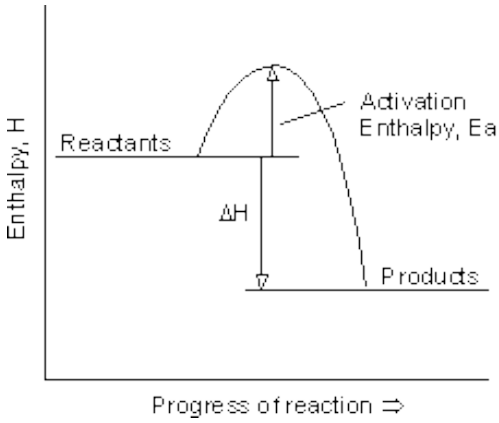
Enthalpy changes can be shown using an **enthalpy level diagram** or **energy profile**.



*When enthalpy of products > enthalpy of reactants = **Endothermic***



*When enthalpy of products < enthalpy of reactants = **Exothermic***



*The value of enthalpy changes will vary with conditions (especially concentration of chemicals) hence **standard enthalpy changes** ΔH^{o} are measured under the following standard conditions"*

- Pressure = 100kPa
- Temperature = 25ºC (298 K)
- Concentration = 1 mol/dm^3
- All substances are in their standard state

EXPERIMENTAL MEASUREMENT OF HEAT CHANGES

In order to find out how much energy a reaction gives off or absorbs, we use the **specific heat capacity** equation.

Specific Heat Capacity: The amount of heat required to raise the temperate of a unit mass of substance one degree/

$$\boxed{Q = m\,c\,\Delta T}$$

Q = Energy (J)
m = mass (g)
c = specific heat capacity (J/g ºC)
ΔT = Change is temperature

Assumptions:

- The reaction is assumed to occur sufficiently rapidly to reach maximum temperature
- There is no heat transfer between the solution, thermometer, the surrounding and the calorimeter
- The solution if sufficiently diluted so it matches water's specific heat capacity

HEAT OF COMBUSTION

The **standard enthalpy of combustion (ΔHc^{θ})** for a substance is the heat energy released when one mole of a pure substance is completely burnt in excess oxygen under standard conditions.

Enthalpy of formation:

$$\Delta H^{\theta} reaction = \sum(\Delta H_f{}^{\theta}\ products) - \sum(\Delta H_f{}^{\theta}\ \ reactants)$$