Calculating Heat Changes

When an exothermic reaction occurs in water, the energy released is transferred to the water and the temperature of the water increases. (Vice versa for endothermic)

This energy is the ΔH and can be calculated if the **mass** of the water and its **temperature change** is known.

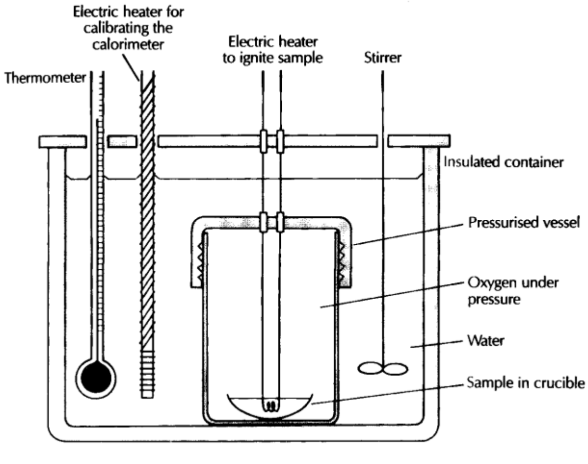
$$\boxed{\Delta H = m.s.\Delta T}$$ (unit is joules)

- **s** = specific heat capacity of water. (4.2 J g^{-1} $°C^{-1}$)
  it takes 4.2J of energy to raise the temperature of 1g of water by 1 ºC
- **m** = mass of water(g)
- **ΔT** is the change in temperature of the water (ºC)
- **ΔH** = negative with a temperature increase (exothermic)
- **ΔH** = positive with a temperature decrease (endothermic)

Example:
How much heat energy is needed to raise the temperature of 200g of water from 44 to 74ºC?

$$\begin{aligned}\Delta H &= m.s.\Delta T\\ &= 200 \times 4.2 \times 30\\ &= 25\ 200\text{J} \quad (25.2\text{kJ})\end{aligned}$$

A calorimeter is used to find the heat of reaction for substances. It is insulated so that all heat energy is transferred within the calorimeter.



Thermo-chemical Equations

Balanced chemical equations that include

- states of all substances
- enthalpy change (ΔH).
  - ΔH is directly proportional to the amount (moles) of substance

Octane is a major component of petrol. Complete combustion of one mole of octane to form carbon dioxide and steam releases 5054kJ. This can be written as a thermochemical equation:

$$C_8H_{18(l)} + 12\tfrac{1}{2}O_{2\,(g)} \rightarrow 8CO_{2\,(g)} + 9H_2O_{(l)} \qquad \Delta_rH = -5054kJmol^{-1}$$

- If you burnt 2 moles of octane, twice the energy would be released (10 108kJ)
- If a reaction occurs in **reverse**, it has the same magnitude of ΔH but the opposite sign:

$H_2O_{(l)} \rightarrow H_2O_{(g)} \qquad \Delta H = +44kJmol^{-1}$

$H_2O_{(g)} \rightarrow H_2O_{(l)} \qquad \Delta H = -44kJmol^{-1}$

Example 1

Calculate the energy released when 250g of petrol (octane) burns completely in a car engine using the above equation.

- **Step 1:** Calculate amount (n) of octane

$n(C_8H_{18})$	=	m/M
	=	250/114
	=	**2.193 mol**

- **Step 2:** Calculate the energy released when this amount burns

Energy	=	$\mathbf{\Delta_rH \cdot n}$
	=	5054 $kJmol^{-1}$ x 2.193 mol
	=	**11 083 kJ** released

Example 2

What mass of methane is burnt to form carbon dioxide and water so as to provide 40 000 kJ of energy?

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O \qquad \Delta_rH = -890kJ\ mol^{-1}$$

- **Step 1:** Calculate amount (n) of methane

$\mathbf{n(CH_4)}$	=	**Energy / Δ_rH**
	=	$\dfrac{40\ 000\ kJ}{890kJ\ mol^{-1}}$
	=	**44.9 mol** required (3s.f)

- **Step 2:** Calculate the mass of methane

$m(CH_4)$	=	n.M
	=	44.94 mol x 16 $gmol^{-1}$
	=	**719 g** (3s.f)

Example 3

12.5 g of NaOH was dissolved in water and excess HCl was added. Using the temperature increase and the heat capacity of water, it was calculated that 18.4 kJ of heat was released

(a) Determine the enthalpy change, $\Delta_r H$, for the following reaction

$$NaOH_{(aq)} \quad + \quad HCl_{(aq)} \quad \rightarrow \quad NaCl_{(aq)} \quad + \quad H_2O_{(l)}$$

- **Step 1:** Calculate amount (n) of NaOH

  n(NaOH) = m/M

  = 12.5/40

  = **0.313 mol (3s.f)**

- **Step 2:** Calculate the enthalpy change when 1 mol reacts

  Δ_rH = **Energy / n**

  = 18.4 kJ / 0.313 mol

  = **- 58.9 kJmol^{-1}** (negative as heat was released)

(b) What mass of NaOH is required to produce 150 kJ of energy

- **Step 1:** Calculate amount (n) NaOH

  n(NaOH) = **Energy / Δ_rH**

  $= \dfrac{150 \text{ kJ}}{58.9 \text{kJ mol}^{-1}}$

  = **2.55 mol** (3s.f)

- **Step 2:** Calculate the mass(m) of NaOH

  m(NaOH) = n.M

  = 2.55 mol x 40 gmol^{-1}

  = **102 g** (3s.f)