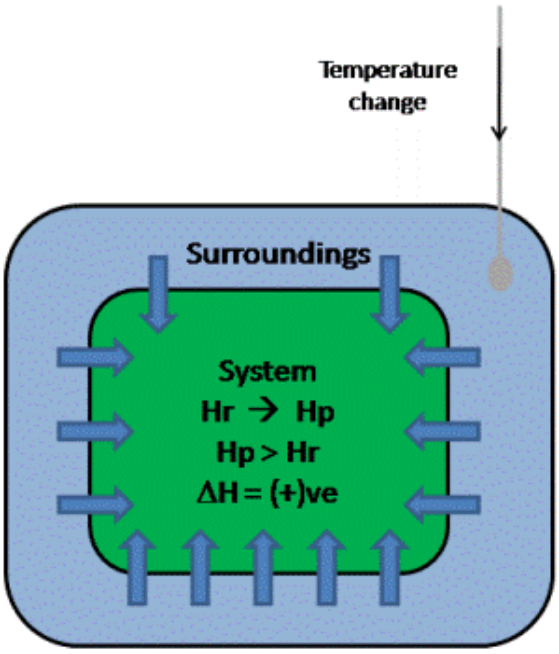
1.4 Enthalpy

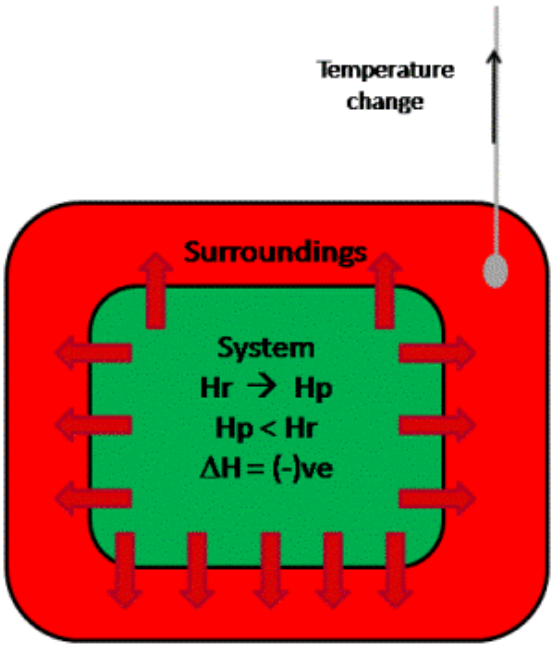
What is chemical energy?

- Chemical energy is a form of potential energy which is stored in chemical bonds.
- Chemical bonds are the attractive forces that bind atoms together.
- As a reaction takes place, bonds break in the reactants and new bonds are formed in the products.
- The difference in energies of the bonds between the atoms of reactants and products is called chemical energy.

Enthalpy, H

- Enthalpy, H, is the heat energy that is stored in a chemical **system**, as **reactants or products**.
- It is impossible to measure the enthalpy content of a system directly but we can measure the differences in enthalpy contents.
- The difference in enthalpy contents is the energy either given out or absorbed.





- The energy exchanged with the **surroundings** can be given out as light, sound, electrical but is more usually heat.
- This change in enthalpy content is called **enthalpy change, ΔH**.

Conservation of energy (First Law of thermodynamics)

Energy cannot be created or destroyed just transferred between the system and the surroundings

- The surroundings, water, tube, air, either gains the energy given out by the system (getting hot)
- Or the surroundings, water, tube, air loses energy to the system (getting cold) - like ethanol evaporating from your hand
- This means that:

Heat loss in a chemical system	=	**Heat gain to the surroundings**	**Temperature increases**
Heat gain in a chemical system	=	**Heat loss to the surroundings**	**Temperature decreases**

- What your thermometer is measuring is the **surroundings**. That is - the water, tube, air the energy is transferred from / to.

Enthalpy change, ΔH:

Enthalpy Change, ΔH:
Is the change in heat energy at constant pressure

- This is the difference between the enthalpy contents of the products and the reactants at constant pressure:

$$\Delta H = H_{products} - H_{reactants}$$

- It is rare for enthalpy contents of products and reactants to be exactly equal.
- This means that there is almost always a transfer of energy between the **system** and the **surroundings**.

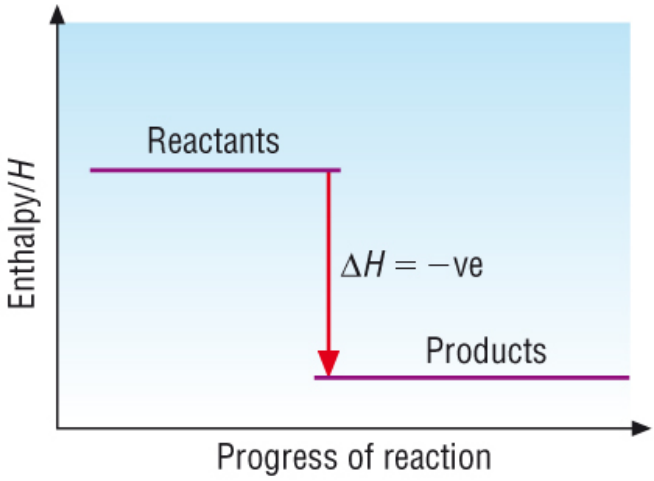
Exothermic reactions:

- This means that the enthalpy content of the **products** is **smaller** than the **reactants**.
- The excess energy is **transferred from the system to the surroundings - hot**

$H_{products} < H_{reactants}$

$\Delta H = H_{products} - H_{reactants}$
ΔH = Small - Large
ΔH = Negative value

Energy is transferred from the system to the surroundings



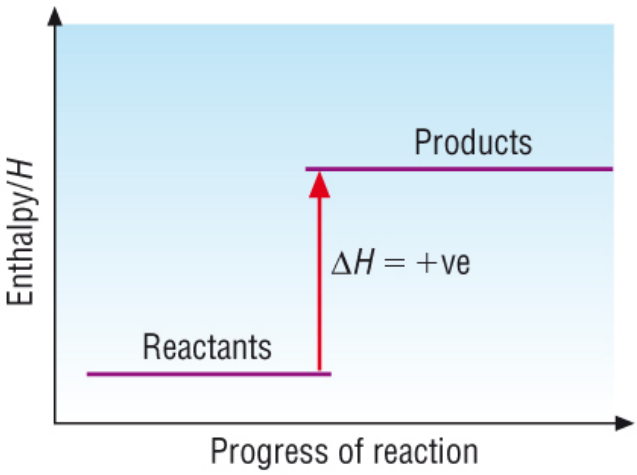
Endothermic reactions:

- This means that the enthalpy content of the **products** is **greater** than the **reactants**.
- The excess energy is **transferred to system from the surroundings - cold**

$H_{products} > H_{reactants}$

$\Delta H = H_{products} - H_{reactants}$
ΔH = Large - Small
ΔH = Positive value

Energy is transferred to the system from surroundings



Exothermic and endothermic reactions

1) Combustion of fuels:

- The most common example is the oxidation of methane forming carbon dioxide and water:

$$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)} \quad \Delta H = -890 \text{ Kj mol}^{-1}$$

- The negative sign means that the reaction is **exothermic**.
- This means that $H_{products} < H_{reactants}$
- The units tell you that **890 Kj** of energy is given out per **mole** of methane.
- Changes in enthalpy are given in **molar quantities**:

2) Respiration:

- An important exothermic reaction providing energy for all living things:

$$C_6H_{12}O_{6(aq)} + 6O_{2(g)} \rightarrow 6CO_{2(g)} + 6H_2O_{(l)} \quad \Delta H = -2801 \text{ Kj mol}^{-1}$$

- The negative sign means that the reaction is **exothermic**.
- This means that $H_{products} < H_{reactants}$
- The units tell you that **2801 Kj** of energy is given out per **mole** of sugar.
- Changes in enthalpy are given in **molar quantities**:

Endothermic reactions:

1) Photosynthesis:

- An important endothermic reaction without which there would be no life.

$$6CO_{2(g)} + 6H_2O_{(l)} \rightarrow C_6H_{12}O_{6(aq)} + 6O_{2(g)} \quad \Delta H = +2801 \text{ Kj mol}^{-1}$$

- The reverse reaction of respiration.
- The positive sign means that the reaction is **endothermic**.
- This means that $H_{products} > H_{reactants}$
- The units tell you that **2801 Kj** of energy is taken in per **mole** of sugar made.
- Changes in enthalpy are given in **molar quantities**:

2) Thermal decomposition of limestone:

- Limestone contains calcium carbonate:

$$CaCO_{3(s)} \rightarrow CaCO_{(s)} + CO_{2(g)} \quad \Delta H = +178 \text{ Kj mol}^{-1}$$

- The positive sign means that the reaction is **endothermic**.
- This means that $H_{products} > H_{reactants}$
- The units tell you that **178 Kj** of energy is taken in per **mole** of calcium carbonate.
- Changes in enthalpy are given in **molar quantities**:

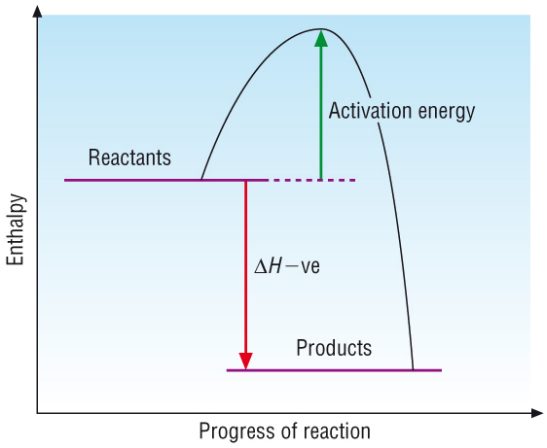
Enthalpy profile diagrams

Activation energy:

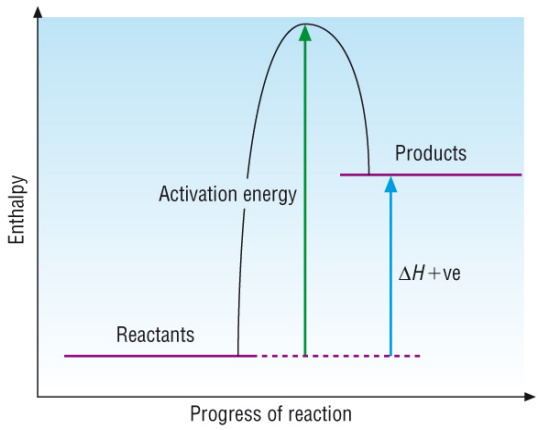
Activation Energy, E_a
Is the minimum energy required to start a reaction by the breaking of bonds

- Most reactions do not occur **'spontaneously'** but need a little bit of energy to get them going, a spark is needed to set methane alight.
- This **'bit of energy'** is called the **activation energy**. It is the energy required to **break the bonds in the reactants**.

Exothermic reactions:



Endothermic reactions:



Exothermic reactions:

Where $H_{products} < H_{reactants}$

ΔH = negative

- Even though the products are lower in energy than the reactants, a small amount of energy is needed to break the reactant bonds, the **activation energy**.
- We say the reaction has to overcome the energy barrier.
- After that, the exothermic nature of the reaction is enough to break more reactant bonds.
- A way of thinking of this is like a cyclist riding a bike up a small hill to gain a large amount of 'free wheeling' or energy.

Endothermic reactions:

Where $H_{reactants} > H_{products}$

ΔH = positive

- The products are higher in energy than the reactants, a small amount of energy is still needed to break the reactant bonds, the **activation energy**.
- The reaction has to overcome the energy barrier.
- This time there is no excess energy to break more reactant bonds, a sustained amount of energy needs to be continually supplied to keep the reaction going.
- This time the cyclist rides a bike up a large hill to gain a small amount of 'free wheeling' or energy.

Standard enthalpy changes

Standards:

- Enthalpy changes for reactions will vary slightly depending upon the conditions under which the reaction is carried out.
- Chemists use **standard conditions** to ensure that all reactions and corresponding enthalpy changes are carried out under the same conditions.
- They are as close to normal lab conditions as possible.

Standard conditions, θ :

P = 100 kPa

1 Mole or 1 Molar solutions

Normal physical states at standard temperature and pressure

Temperature at 298K (25°C)

Written as:

ΔH^{θ}_{298}

Δ - Change in

H - Enthalpy

θ symbol represents **standard conditions**

298 Represents 298K

Standard states:

- Standard enthalpy changes must have substances in their **standard states** under these standard conditions:

Substance	Chemical symbol and state	Explanation
Magnesium	$Mg_{(s)}$	Magnesium is a solid under standard conditions, **(s)**
Hydrogen	$H_{2(g)}$	Hydrogen is a gas under standard conditions, **(g)**
Water	$H_2O_{(l)}$	Water is a liquid under standard conditions, **(l)**

Bond enthalpies

Bond enthalpy:

> **Bond Enthalpy**
> ***Is the energy required to break one mole of a given covalent bond in the molecule in the gaseous state***

Breaking bonds = Energy is put in to break bonds **Endothermic process**

Forming bonds = Energy released when bonds are formed **Exothermic process**

Examples:

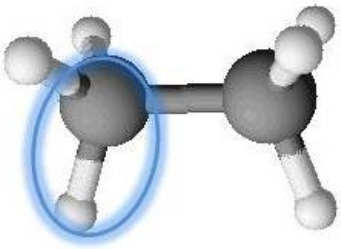
$H - H_{(g)} \rightarrow 2H_{(g)}$ $\Delta H = +436\ KjMol^{-1}$

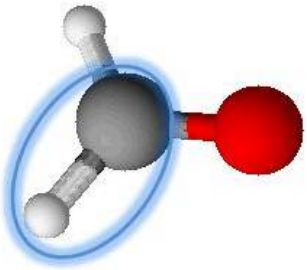
$H - Cl_{(g)} \rightarrow H_{(g)} + Cl_{(g)}$ $\Delta H = +432\ KjMol^{-1}$

Bond enthalpies are always positive values as bonds are always broken – Energy is always put in to break the bonds

Mean bond enthalpy:

- Bonds like C - H exist in all hydrocarbons but their bond enthalpies will vary slightly depending on their environment:





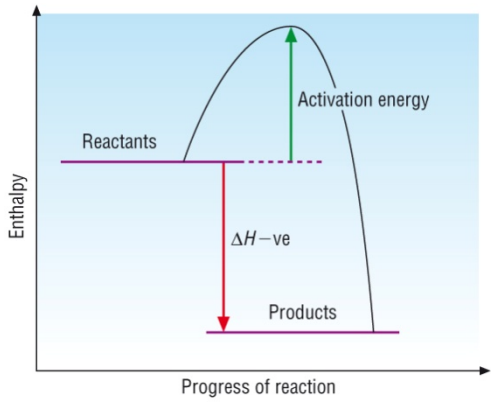
- For this reason we use **Mean bond enthalpies:**

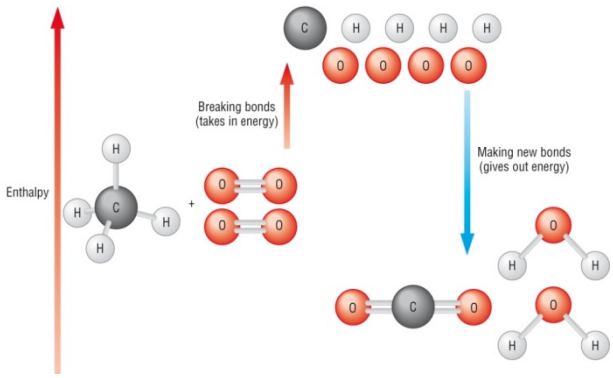
Bond	*Average bond enthalpy / KJ Mol^{-1}*
C - H	+413
O = O	+497
O - H	+463
C = C	+612
H - H	+436

Mean bond enthalpy:

> **Mean bond Enthalpy**
> ***Is the average value for the bond enthalpy over the range of compounds it is found in***

Breaking and making bonds:





- In this and any reaction, reactant bonds will be broken and new product bonds will be formed.
- When **bonds are broken** energy is required making it an **endothermic** process, **the activation energy**.
- When new **bonds are formed** energy is release making it an **exothermic** process.

Exothermic or endothermic reactions:

Exothermic:

- If the **energy released** when new bonds form making the products **is greater** than the **energy needed** to break the bonds of the reactants, the reaction is **exothermic**.

E_{in} Breaking weak bonds < E_{out} Forming strong bonds Exothermic process

Endothermic:

- If the **energy released** when new bonds form making the products **is less** than the **energy needed** to break the bonds of the reactants, the reaction is **endothermic**.

E_{in} Breaking strong bonds > E_{out} Forming weak bonds Endothermic process

Using bond enthalpies to determine enthalpy changes:

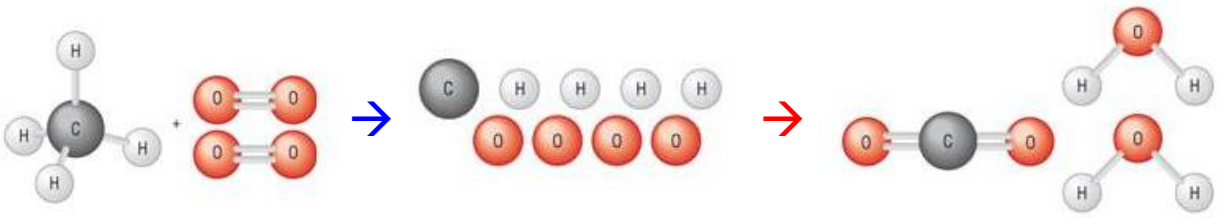
Energy required to break bonds = **Σ (Bond enthalpies of bonds broken)**

Breaking weak bonds = **Σ (Bond enthalpies of bonds formed)**

Formula:

ΔH = Σ (Bond enthalpies of bonds broken) - Σ (Bond enthalpies of bonds formed)

Worked example:



Σ (Bond enthalpies of bonds broken) **Σ (Bond enthalpies of bonds formed)**

Reactants				Products			
Bond	Number	Bond energy	Total	Bond	Number	Bond energy	Total
C - H	4	413	1652	C = O	2	805	1610
O = O	2	497	994	O - H	4	463	1852
Σ (Bond enthalpies of bonds broken) =			2646	Σ (Bond enthalpies of bonds formed) =			3462

ΔH = Σ (Bond enthalpies of bonds broken) - Σ (Bond enthalpies of bonds formed)

ΔH = 2646 - 3462

ΔH = - 816 Kj Mol^{-1}

- These values are often **less accurate** as the **mean** value and not the **actual** bond enthalpies are used.

Questions:

1) **Calculate the enthalpy change for the reaction:**

$$H_{2(g)} + Br_{2(g)} \rightarrow 2HBr_{(g)}$$

Bond	H – H	Br – Br	H – Br
Mean bond enthalpy / Kj Mol^{-1}	+438	+193	+366

2) **Calculate the enthalpy change for the reaction:**

$$C_4H_{10(g)} + 6.5O_{2(g)} \rightarrow 4CO_{2(g)} + 5H_2O_{(g)}$$

Bond	C – H	C – C	O = O	C=O	O – H
Mean bond enthalpy / Kj Mol^{-1}	+413	+347	+497	+805	+463

3) **Calculate the enthalpy change for the reaction:**

$$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$$

Bond	N_2	H – H	N – H
Mean bond enthalpy / Kj Mol^{-1}	+945	+436	+391

Types of standard enthalpy changes:

- There are 3 standard enthalpy changes that you need to know:

1) Standard enthalpy change of formation, $\Delta_f H^\theta$

> ***The enthalpy change that occurs when 1 mole of a compound is formed from its constituent elements in their standard states under standard conditions.***

$$H_{2(g)} + \tfrac{1}{2}O_{2(g)} \rightarrow H_2O_{(l)}$$

- $\Delta_f H^\theta$ of an **element** must be **zero**
- The more negative a $\Delta_f H^\theta$ **the more stable** that compound is.

2) Standard enthalpy of combustion – $\Delta_c H^\theta$

> ***The enthalpy change that occurs when 1 mole of a compound reacts completely with oxygen under standard conditions where all reactants and products are in their standard states.***

$$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}$$

3) Standard enthalpy change of reaction, $\Delta_r H^\theta$

> ***The enthalpy change when a reaction occurs in the molar quantities shown in the equation under standard conditions where all reactants and products are in their standard states.***

$$Fe_2O_{3(s)} + 2Al_{(s)} \rightarrow 2Fe_{(s)} + Al_2O_{3(s)}$$

- For this enthalpy change we need a reaction to refer to.

Questions on standard enthalpy of formation, $\Delta_f H^\theta$:

1) **Write the equations for the standard enthalpy change of formation, $\Delta_f H^\theta$ for the following:**

a) $\rightarrow$ $C_2H_{4(g)}$

b) $\rightarrow$ $C_4H_9Br_{(l)}$

c) $\rightarrow$ $K_2CO_{3(s)}$

2) Write the equations for the standard enthalpy change of formation, $\Delta_f H^\theta$ for the following:

a) Propane

b) Sulphuric acid

c) Potassium nitrate

3) Ammonia is made by the Haber process and the equation is shown below

$$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)} \qquad \Delta H = -93 \text{ KJ mol}^{-1}$$

a) What is the enthalpy change of formation of ammonia?

b) 170g of ammonia is made. What is the enthalpy change to make this amount of ammonia?

c) What is the enthalpy change for the decomposition of 10 moles of ammonia?

Questions on standard enthalpy of combustion, $\Delta_c H^\theta$:

4) Write the equations for the standard enthalpy change of combustion, $\Delta_c H^\theta$ for the following:

a) $C_5H_{12(l)}$ →

b) $C_4H_9OH_{(l)}$ →

c) $C_6H_{12}O_{6(s)}$ →

5) Write the equations for the standard enthalpy change of combustion, $\Delta_c H^\theta$ for the following:

a) Propane

b) Ethene

c) Ethanol

6) Methane is used in the homes as a gas supply for heating and has an enthalpy change of combustion, $\Delta_c H^\theta$ = -882 Kj mol^{-1}

a) Write a balance chemical equation for the enthalpy change of combustion of methane.

b) What is the enthalpy change of combustion when 160g of methane undergoes complete combustion?

c) What mass of methane would be required to produce 4.41 Kj of energy when it undergoes complete combustion?

Questions on standard enthalpy of reaction, $\Delta_r H^\theta$:

7) The reaction below is used for welding in isolated places (railway lines)

$$Fe_2O_{3(s)} + 2Al_{(s)} \rightarrow 2Fe_{(s)} + Al_2O_{3(s)} \qquad \Delta_r H^\theta = -849 \text{ Kj mol}^{-1}$$

a) What is the enthalpy change per mole of Iron, Fe?

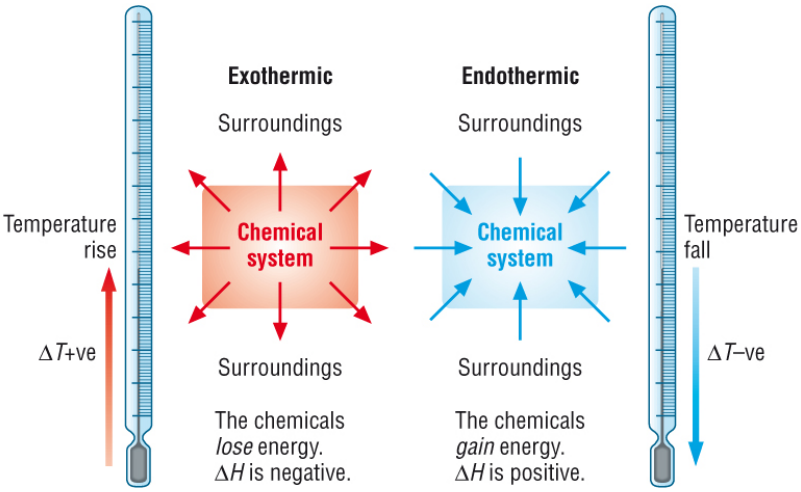
b) What is the enthalpy change for the production of 1 tonne of Iron, Fe?

Determination of enthalpy changes - calorimetry

- Enthalpy content of reactants and products cannot be measured directly.
- You can measure the enthalpy change between the reactants and products - enthalpy change for a reaction.

Remember:

Heat loss in a chemical system	= **Heat gain to the surroundings**	**Temperature increases**
Heat gain in a chemical system	= **Heat loss to the surroundings**	**Temperature decreases**

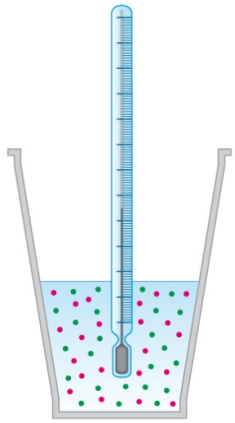


- The enthalpy change for any reaction, $\Delta_r H$ is measured in **KJ mol^{-1}**, which is **energy** per **mole.** This means we need to know 2 things:

  i) The energy change - for which we use temperature change and mass of stuff the heat energy is exchanged between

  ii) The amounts in moles of the limiting reagent

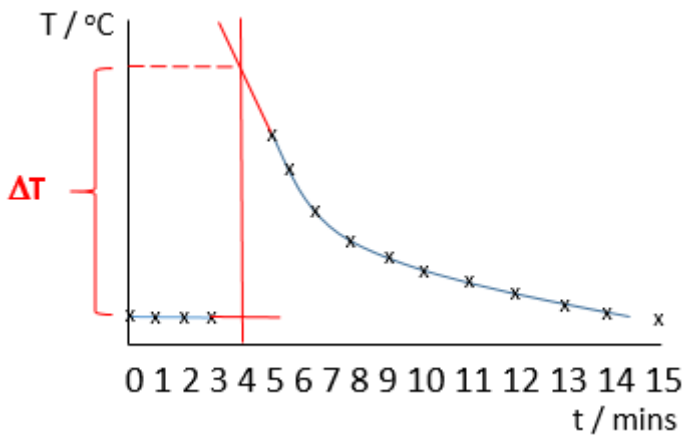
Calorimetry – Enthalpy changes involving solutions:



The particles are dissolved in the solution. The heat is exchanged from the particles into the solution. The heat exchange is measured using the thermometer.

- A calorimeter is an insulated reaction vessel which minimises heat exchange to the air.
- We use the water as the surroundings so we measure the amount of **energy** exchanged with the water.
- The simplest calorimeter is a polystyrene cup but they can get quite complex.

Calculating ΔT to avoid heat loss errors:



- The initial solution is added and the temperature taken for 3 minutes.
- On the 4th minute the other solution or solid is added but temperature not recorded.
- 5 – 15 minutes the temperature is recorded.
- The points are extrapolated to the 4 minute mark and the two temperatures read off the axis.
- This is ΔT.

Calculation

Step 1: The energy change, q:

$$q = m \; c \; \Delta T$$

q	**Quantity of energy exchanged, joules (divide by 1000 converts J to kJ)**
m	**Mass / volume of the water grammes (g = cm^3 as the density of water is 1 gcm^{-3})**
c	**Specific heat capacity j g^{-1} K^{-1} (water = 4.18)**
ΔT	**Change in temperature K,**

- As **enthalpy** is in **Kj** mol^{-1} and **q** is in j, convert the energy calculated by dividing by 1000).

Step 2: Calculate the number of moles used:

$$\text{Moles} = \frac{\text{Mass}}{\text{Mr}} \quad \text{or} \quad C \times V\,(dm^3)$$

Step 3: Calculate the enthalpy, ΔH:

$$\text{Enthalpy, } \Delta H = \frac{\text{Energy, q}}{\text{Moles}}$$

Step 4: Check you have the sign correct:

(-)ve for exothermic reactions

(+)ve for endothermic reactions

Example:

Mg is added to 100cm^3 of 2.00 Mol dm^{-3} $CuSO_4$, the temperature rose from 20.0^{o}C to 65°C.

Step 1: The energy change, q:

$$q = m \quad c \quad \Delta T$$

$$q = 100 \times 4.18 \times 45$$

$$q = 18810 \text{ joules}$$

$$q = 18.81 \text{ Kj}$$

Step 2: Calculate the number of moles used:

$$\text{Moles} = C \times V\ (dm^3)$$

$$\text{Moles} = 2 \times 0.1$$

$$\text{Moles} = 0.200$$

Step 3: Calculate the enthalpy, ΔH:

$$\text{Enthalpy, } \Delta H = \frac{\text{Energy, q}}{\text{Moles}}$$

$$\Delta H = \frac{18.81}{0.2}$$

$$\Delta H = 94.05 \text{ Kj mol}^{-1}$$

Step 4: Check you have the sign correct:

(-)ve for exothermic reactions

(+)ve for endothermic reactions

$$\Delta H = -94.05 \text{ Kj mol}^{-1}$$

Questions:

1) Ca is added to 50 cm^3 of 1.00 Mol dm^{-3} H_2SO_4, the temperature rose from 20.0°C to 40°C. Calculate the enthalpy change for this reaction:

Step 1: The energy change, q:

$$q = m \quad c \quad \Delta T$$

Step 2: Calculate the number of moles used:

$$\text{Moles} = C \times V\ (dm^3)$$

Step 3: Calculate the enthalpy, ΔH:

$$\text{Enthalpy, } \Delta H = \frac{\text{Energy, q}}{\text{Moles}}$$

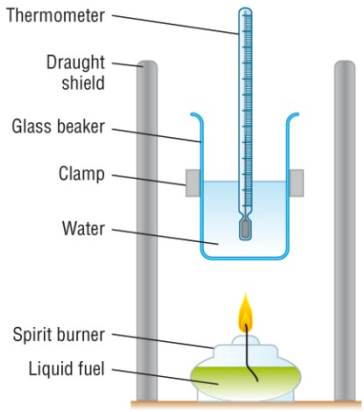
Step 4: Check you have the sign correct:

$$\Delta H =$$

2) 50 cm^3 of 2.00 Mol dm^{-3} HCl was added to 50 cm^3 of 2.00 Mol dm^{-3} NaOH, the temperature rose from 22.0°C to 47°C. Calculate the enthalpy change for this reaction:

Determination of enthalpy changes of combustion - calorimetry

- The enthalpy change of combustion, $\Delta_c H$ is measured in **KJ mol^{-1}**, which is **energy** per **mole.** The calculation follows the same process:



- To calculate $\Delta_c H$, measure a known mass (volume) of water.
- The temperature rise of the water.
- Weigh the burner before heating.
- Heat to raise the temperature.
- Take a final temperature.
- Reweigh the burner to get a mass of fuel used.

Example: 1.5g of propan-1-ol heated 250cm^3 of water by 45°C.

Step 1: The energy change, q:

$$q = m \quad c \quad \Delta T$$

$$q = 250 \times 4.18 \times 45$$

$$q = 47025 \text{ joules}$$

$$q = 47.025 \text{ Kj}$$

Step 2: Calculate the number of moles used:

$$\text{Moles} = \frac{\text{mass}}{\text{Mr}}$$

$$\text{Moles} = \frac{1.5}{60}$$

$$\text{Moles} = 0.025$$

Step 3: Calculate the enthalpy, ΔH:

$$\text{Enthalpy, } \Delta_c H = \frac{\text{Energy, q}}{\text{Moles}}$$

$$\Delta_c H = \frac{47.025}{0.025}$$

$$\Delta_c H = 1881 \text{ Kj mol}^{-1}$$

Step 4: Check you have the sign correct: Always (-)ve as combustion is always exothermic

$$\Delta_c H = -1881 \text{ Kj mol}^{-1}$$

Questions:

1) 2.3 g of ethanol was burned in air. The heat energy increased the temperature of 100 cm^3 of water by 50°C. Calculate the enthalpy change for this combustion reaction:

Step 1: The energy change, q:

$$q = m \quad c \quad \Delta T$$

Step 2: Calculate the number of moles used:

$$\text{Moles} = \frac{\text{mass}}{\text{Mr}}$$

Step 3: Calculate the enthalpy, ΔH:

$$\text{Enthalpy, } \Delta_c H = \frac{\text{Energy, q}}{\text{Moles}}$$

Step 4: Check you have the sign correct:

$$\Delta_c H =$$

2) 1.1 g of propane was burned in air. The heat energy increased the temperature of 200 cm^3 of water by 25°C. Calculate the enthalpy change for this combustion reaction:

Comparison of experimental value with standard enthalpy change:

Standard enthalpy change of combustion of C_3H_7OH, $\Delta_cH^{\theta}_{298}$ - 2021 Kj Mol^{-1}

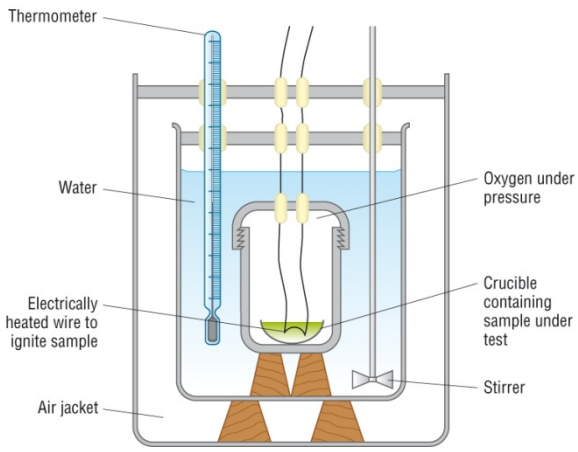
Experimental enthalpy change of combustion of C_3H_7OH, Δ_cH - 1881 Kj Mol^{-1}

Sources of errors:

- Heat absorbed by the container
- Heat loss to the surroundings

- Incomplete combustion of the fuel - less heat energy given out
- Evaporation of the fuel

Improvements:



- Use a **Bomb calorimeter** (left).
- This apparatus reduces heat loss s the water is insulated from the surroundings.
- It is burnt in oxygen to ensure complete combustion.

Hess's law:

Measuring enthalpy changes indirectly:

- Sometimes it is not possible to measure an enthalpy change directly due to:

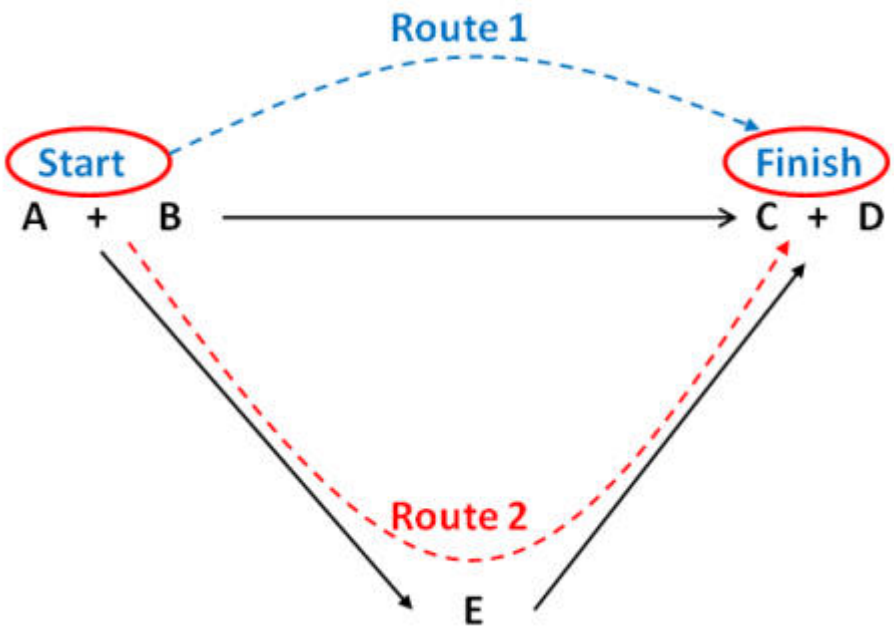
  High activation energy

  Slow rate of reaction

- Hess's law allows us to work out enthalpy changes that are not possible to measure:

Hess's Law:

The total enthalpy change for a reaction is independent of the route taken.



- If this was not true, it could be possible to gain energy via route A instead of going via route B. This would break the **1st law of thermodynamics**.
- The total energy for each route is the same: These are called **Enthalpy cycles**.

Using Hess's law – Enthalpy cycles

- Enthalpy cycles are used to determine enthalpy changes for reactions that are not easily measured directly.
- This can be done in 2 ways:

  1. **Enthalpy changes of formation**

  2. **Enthalpy changes of combustion**

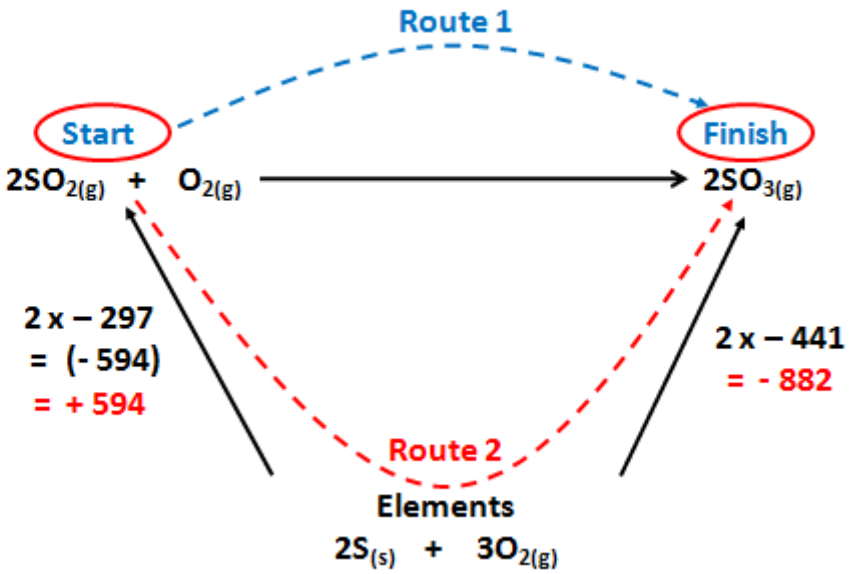
1) Calculating enthalpy changes from Enthalpy changes of formation, $\Delta_f H^\theta$

- Enthalpy changes of formation can be used in an enthalpy cycle to work out enthalpy changes.
- Consider the reaction:

$$2SO_{2(g)} + O_{2(g)} \rightarrow 2SO_{3(g)}$$

- The enthalpy cycle would be:

Enthalpy cycle:



- we use the opposite sign for **- 594 = + 594** as our route, **Route 2** goes the **opposite direction** to the **enthalpy cycle**

Route 2 = + 594 + - 882

Route 2 = - 288 Kj Mol^{-1}

Remember

Route 1 = Route 2

Route 1 = - 288 Kj Mol^{-1}

Using the formula:

$$\Delta H = \Sigma \Delta_f H^\theta \text{ (products)} - \Sigma \Delta_f H^\theta \text{ (reactants)}$$

$$\Delta H = (2 \times -441) - (2 \times -297)$$

$$\Delta H = -288 \text{ Kj Mol}^{-1}$$

Key point: Look at the table of data → Formation → Arrows up

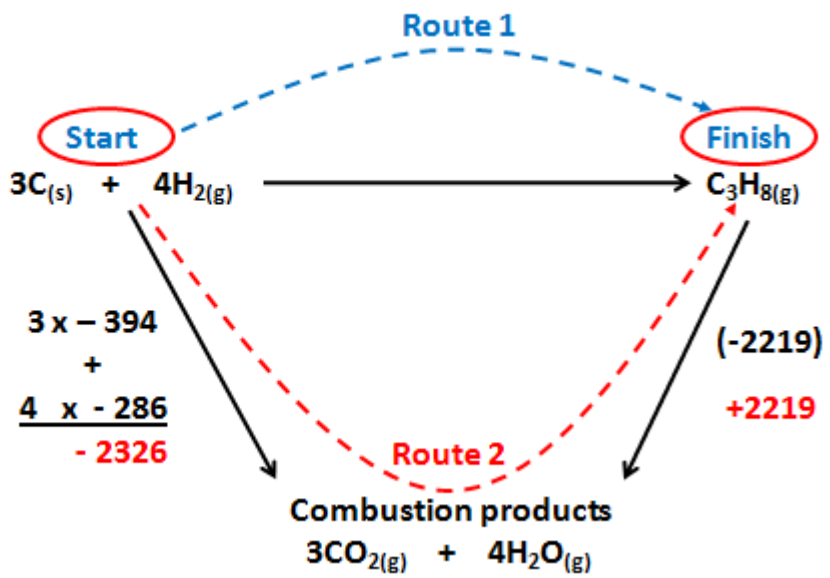
2) Calculating enthalpy changes from Enthalpy changes of combustion, $\Delta_c H^\theta$

- Consider the enthalpy change below:

$$3C_{(s)} + 4H_{2(g)} \rightarrow C_3H_{8(g)}$$

- Combustion reactions can be used to find this enthalpy change as carbon hydrogen and propane all burn in oxygen.
- Constructing an enthalpy cycle:

Enthalpy cycle:



- we use the **opposite sign for - 2219** as our route, **Route 2** goes in the **opposite direction** to the **enthalpy cycle**

Route 2 = - 2326 + + 2219

Route 2 = - 107 Kj Mol^{-1}

Remember

Route 1 = Route 2

Route 1 = - 107 Kj Mol^{-1}

Using the formula:

$\Delta H = \Sigma \Delta_c H^\theta$ **(reactants)** - $\Sigma \Delta_c H^\theta$ **(products)**

ΔH = **(3 x - 394) + (4 x - 286)** - **- 2219**

ΔH = **- 107 Kj Mol^{-1}**

Key point: Look at the table of data → Combustion → Arrows down

Other enthalpy cycles:

- As long as there is a link between reactants and products, Hess's law can be applied and enthalpy cycles can be constructed.
- Follow the principles as outlined above and you won't go wrong.

Summary of enthalpy cycles:

Step 1:- Write a balanced chemical equation for the reaction.

Step 2:- Construct the Enthalpy cycle.

Formation → Arrows up
Combustion → Arrows down

Step 3:- Decide on your routes and draw them on the cycle

Step 4:- Write in the ΔH^{θ} for each compound / element next to the arrows.

Step 5:- Look up the values of each ΔH^{θ} and write them in. Add them up for each route.

Step 6:- Write out Hess's law – Route 1 = Route 2

Step 7:- Put in your numbers.

Step 8:- Calculate ΔH^{θ}

Questions:

1) Calculate the enthalpy change for the following reaction given the information in the table below:

	CH_4	CH_3Cl	HCl
$\Delta_f H^\theta$	-74.8	-134.5	-92.3

$$CH_{4(g)} + Cl_{2(g)} \rightarrow CH_3Cl_{(g)} + HCl_{(g)}$$

2) Calculate the enthalpy change of formation of ammonia given the information below:

$$H_{2(g)} + \tfrac{1}{2}O_{2(g)} \rightarrow H_2O_{(l)} \qquad \Delta H^\theta = -288 \text{ Kj mol}^{-1}$$

$$4NH_{3(g)} + 3O_{2(g)} \rightarrow 2N_{2(g)} + 6H_2O_{(l)} \qquad \Delta H^\theta = -1530 \text{ Kj mol}^{-1}$$

3) For the following questions you will need the information in the table below:

	$H_2O_{(l)}$	$CO_{2(g)}$	$C_2H_{6(g)}$	$C_2H_{4(g)}$
$\Delta_f H^\theta$	-285.5	-393	-83.6	+52.0

a) Write equations for the complete combustion of

i) Ethane

ii) Ethene

iii) hydrogen

b) Calculate the enthalpy of combustion in each case using the data above in the table.

4) Calculate the enthalpy change for the reaction below given the information in the table below:

	$CH_3CH_2CH_2CH_3$	$CH_3CH_2CH{=}CH_2$	H_2
$\Delta_c H^\theta$	-2877	-2717	-286

$$CH_3CH_2CH{=}CH_2 + H_2 \rightarrow CH_3CH_2CH_2CH_3$$

5) Calculate the enthalpy change of formation of butane, C_4H_{10} given the information in the table below:

	$C_{(s)\ graphite}$	$H_{2\ (g)}$	$C_4H_{10\ (g)}$
$\Delta_c H^\theta$	-393.6	-285.9	-2877.1

$$4C_{(s)} + 5H_{2(g)} \rightarrow C_4H_{10(g)}$$

6) Given the following data:

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$ $\Delta H = -890\ kJmol^{-1}$

$CO(g) + 1/2O_2(g) \rightarrow CO_2(g)$ $\Delta H = -284\ kJmol^{-1}$

$C(s) + O_2(g) \rightarrow CO_2(g)$ $\Delta H = -393\ kJmol^{-1}$

$H_2(g) + 1/2O_2(g) \rightarrow H_2O(l)$ $\Delta H = -286\ kJmol^{-1}$

Calculate:

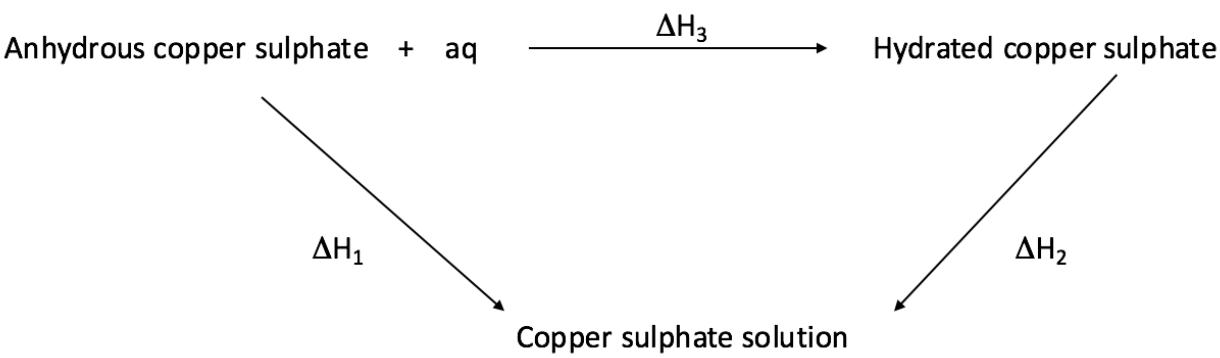
a) The enthalpy of formation of methane

b) The enthalpy change when methane is burned in limited oxygen to form carbon monoxide and water.

c) The enthalpy of formation of carbon monoxide

Required Practical 2 – Measurement of an enthalpy change

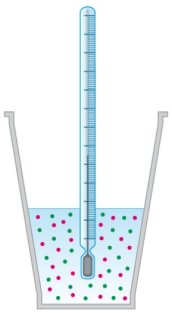
- Hess's Law can be used to find an enthalpy change that cannot be measured directly:

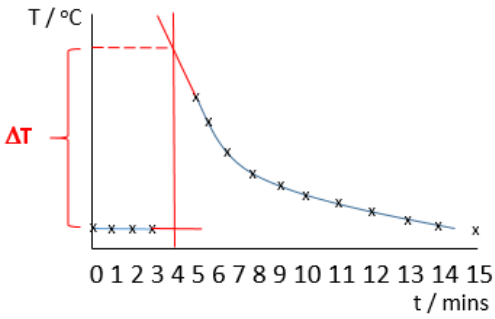


$$\Delta H_3 = \Delta H_1 - \Delta H_2$$

Measuring ΔH_1 and ΔH_2 directly:

1) Weigh 3.90 – 4.10 of anhydrous copper sulphate in a dry stoppered weighing bottle.
2) Place 25cm^3 of water into a polystyrene cup and record its initial temperature, t=0





3) Record the temperature for 3 minutes.
4) On the 4th minute add the anhydrous copper sulphate but do not record the temperature.
5) Stir continuously.
6) 5 – 15 minutes the temperature is recorded.
7) Re weigh the stoppered flask to determine the actually mass of anhydrous copper sulphate added.
8) Plot a graph (as shown above)
9) The points are extrapolated to the 4 minute mark.
10) ΔT is determined from the graph by reading off the Y axis.

Repeat the experiment for hydrated copper sulphate except:

1) Use between 6.20 – 6.30g of hydrated copper sulphate
2) Use 24cm^3 of water. The waters of crystallisation will make the volume up to 25cm^3

Recording results:

	ΔH_1 Mass (g) 2dp	ΔH_2 Mass (g) 2dp
Mass of sample and weighing bottle		
Mass of weighing bottle		
Mass of sample		

Time (mins)	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
ΔH_1 Temp (^oC)															
ΔH_2 Temp (^oC)															

Calculation

Step 1: The energy change, q:

$$q = m \quad c \quad \Delta T$$

- As **enthalpy** is in **Kj** mol^{-1} and **q** is in j, convert the energy calculated by dividing by 1000).

Step 2: Calculate the number of moles used:

$$\text{Moles} = \frac{\text{Mass}}{\text{Mr}} \quad \text{or} \quad C \times V\ (dm^3)$$

Step 3: Calculate the enthalpy, ΔH:

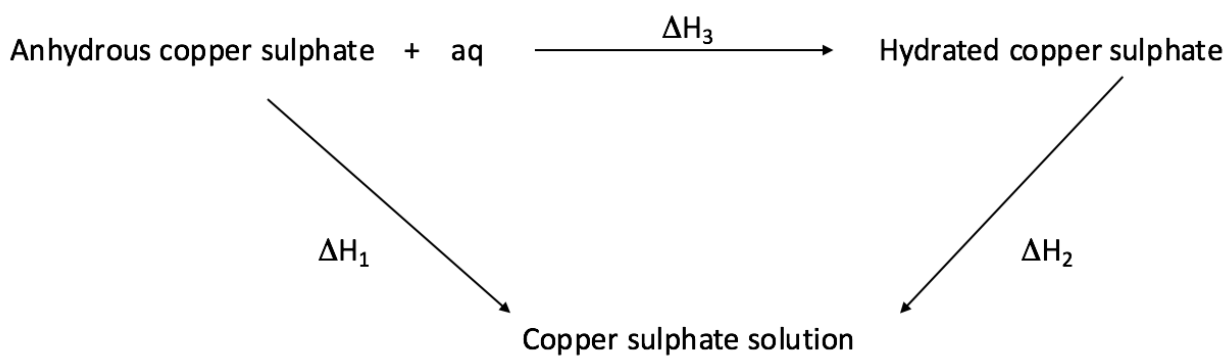
$$\text{Enthalpy, } \Delta H = \frac{\text{Energy, q}}{\text{Moles}}$$

Step 4: Check you have the sign correct:

(-)ve for exothermic reactions

(+)ve for endothermic reactions

Use your answers to ΔH_1 and ΔH_2 to calculate ΔH_3 using Hess's law



$$\Delta H_3 = \Delta H_1 - \Delta H_2$$