PRINCIPLES of PHYSICAL

Week 7

ENERGY CHANGES

Learning Outcomes:

- □ Define:
  - i. Enthalpy (1)
  - ii. Enthalpy change, (ΔH) (1)
  - iii. Activation energy, E_a (1)
  - iv. Exothermic reaction (1)
  - ***v.*** Endothermic reaction (1)
  - ***vi.*** Energy diagram (1)
- □ Describe enthalpy change (ΔH). (2)
- □ Draw labelled energy diagram for an exothermic or endothermic reaction. (showing E_a, Energy of the reactants or ∑ΔH reactants, Energy of the products or ∑ΔH products, Energy of the reaction or ΔH reaction or catalysed pathway) (3)
- □ Determine from a given energy diagram the:
  - a) E_a (1)
  - b) ∑ΔH reactants (1)
  - c) ∑ΔH products (1)
  - d) ΔH reaction (1)
  - e) Catalysed pathway (2)
  - f) Whether the reaction is either endothermic or exothermic (2)
- □ Determine simple everyday reactions as either endothermic or exothermic reactions based on qualitative assessments of the chemical reaction. (e.g. Alchol feeds cool on the skin) (3)
- □ Explain a given everday reaction in terms of heat reactions. (3)
- □ Relate amount (n) to energy change (ΔH) in a thermochemical equation (2)

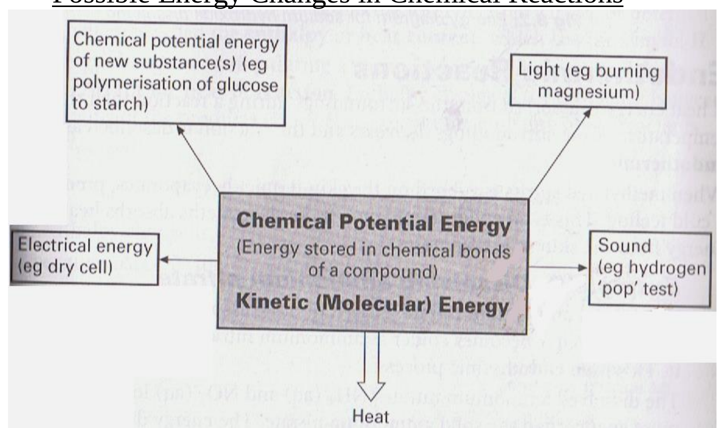
[ΔH reaction = ΣΔH products – ΣΔH reactants]

- □ Evaluate data on energy changes of a reaction or pathways of energy changes of a reaction or thermochemical equations to determine required information or energy diagrams (4)
- □ Define Hess's Law (1)
- □ Write the equation for the desired chemical reacton (2)
- □ Rearrange equations approximately to determine energy change of a chemical reaction including the enthalpy changes (3)
- □ Apply Hess's Law to calculate the energy change for a chemical reaction (involves simple steps) (3)
- □ Evaluate the enthalpy change for a reaction with Hess's Law by adding up steps of reaction and stating whether the overall reaction is exothermic or endothermic (4)

EXOTHERMIC REACTIONS

- Chemical reactions involve energy changes.
- During a reaction, energy changes from one form to another.

Possible Energy Changes in Chemical Reactions



- Most, in not all, chemical reactions involve changes in heat energy.

If heat energy is *produced* during a reaction, the temperature of the surrounding *increases*, and the reaction is described as being **exothermic**.

Such reactions include:

- □ The fusion reaction occurring in the sun (the energy being released warms the earth!!)
- □ The combustion of petrol

Example of Exothermic Reaction

Dissolving sodium hydroxide in water.

When solid sodium hydroxide pellets, NaOH(s), are dissolved in a beaker of water, the water becomes warm. This is an exothermic reaction since heat energy has been produced and transferred to the water, increasing its temperature.

Dissolved sodium hydroxide ($Na^{+}_{(aq)}$ and $OH^{-}_{(aq)}$ ions) contains less energy than the solid sodium hydroxide, $NaOH_{(s)}$.

This energy difference can be shown using an **energy diagram:**

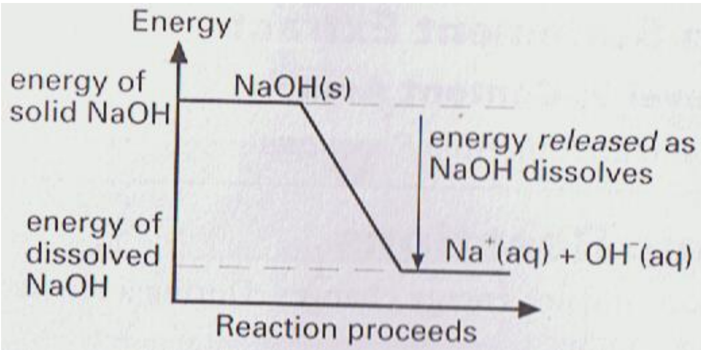


Figure 1: Energy diagram for sodium hydroxide dissolving

ENDOTHERMIC REACTIONS

- If heat enegy is *absorbed* from the surroundings during a reaction then the temperature of the surroundings *decreases*.
- The reaction is described therefore as **endothermic**.
- When methylated spirits is rubbed on the skin it quickly evaporates, producing a cold feeling.
- This is an endothermic process because meths absorbs heat energy from the skin as it evaporates.

Example of Endothermic Reactions

Dissolving ammonium nitrate

When solid ammonium nitrate, $NH_4NO_{3(s)}$ is dissolved in a beaker of water, the water becomes colder as ammonium nitrate absorbs energy from it. This is an endothermic process.

The dissolved ammonium nitrate ($NH_4^+{}_{(aq)}$ and $NO_3^-{}_{(aq)}$ ions) contains more energy than the solid ammonium nitrate.

The energy diagram for this process is:

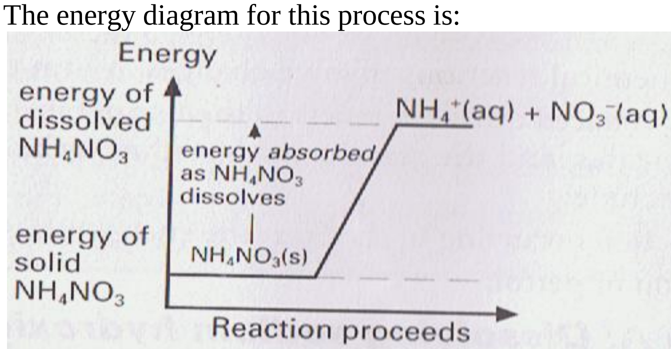


Figure 2: Energy diagram fro ammonium nitrate dissolving

If a reaction is endothermic in one direction (eg left to right in an equation), it will be exothermic in the opposite direction (eg right to left).

Example: Heating hydrated Copper Sulphate

Heat + $CuSO_4.5H_2O$ ⟶ $CuSO_{4(s)}$ + $5H_2O_{(l)}$ (Endothermic)

blue crystals *white powder*

For the "opposite" direction:

$CuSO_{4(s)}$ + $5H_2O_{(l)}$ ⟶ Heat + $CuSO_4.5H_2O$ (Exothermic)

Burns from steam (water or gas) are worse than scolds (water as a liquid) – the energy absorbed when water (liquid) turned to steam (gas) is given out again when the steam condenses on the skin to a liquid.

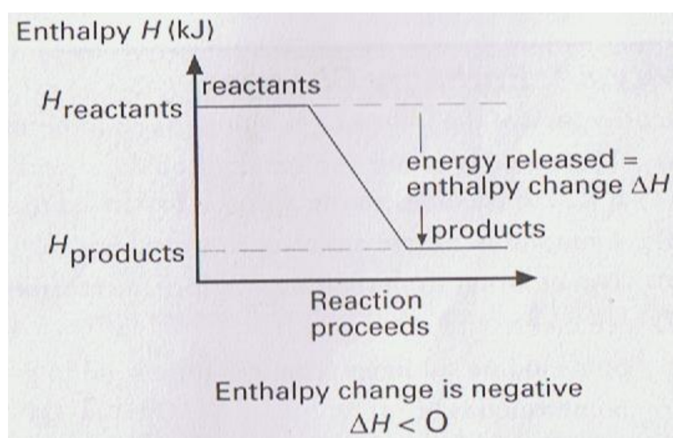
ENTHALPHY

- Substances have an amount of energy, in the form of either kinetic or potential energy, which is called the **enthalpy** or **heat content**.
- Enthalpy has the symbol H.
- The energy change occurring during a reaction is called the **enthalpy change** (symbol **ΔH**) or the **heat of reaction.**
- Enthalpy change is the difference between the enthalpy of the products (H_p or $H_{products}$) and the enthalpy of the reactants (H_r or $H_{reactants}$):

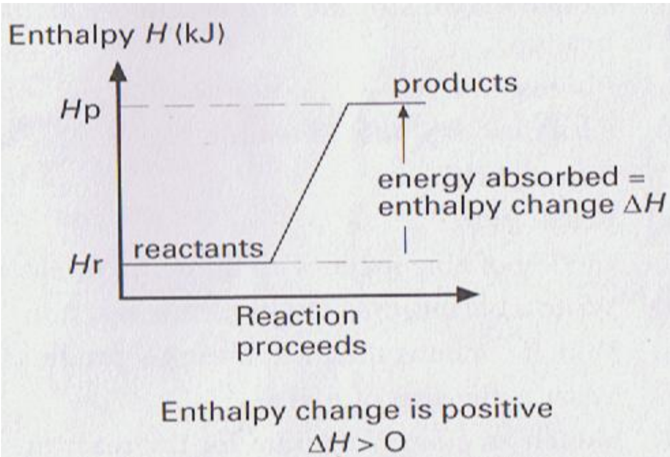
Enthalpy Change (ΔH) = $H_{products}$ - $H_{reactants}$

The enthalpy change is:

- **Negative (ΔH< 0) for Exothermic reactions**



- **Positive (ΔH> 0) for Endothermic reactions**



Enthalpy change in a reaction depends upon the making and breaking of bonds

- Bond-forming is exothermic
- Bond-breaking is endothermic

In a reaction that is overall exothermic, more energy is given out as the bonds of the products form than is taken in as the bonds of the reactants break.

In a reaction that is overall endothermic, more energy is taken in as the bonds of the reactants break than is given out as the bonds of the products form.

Example: Carbon dioxide subliming

$$CO_{2(s)} \longrightarrow CO_{2(g)}$$

- Dry ice subliming is overall an endothermic process.
- To break the van der Waals bonds between CO_2 molecules in the solid state requires energy ('bonds breaking, energy in').
- No bonds form between CO_2 molecules in the gaseous state ('bonds forming, energy out').

Since **energy in** > **energy out**, reaction is overall **endothermic.**

Note:

1. Enthalpy is measured in joules (J) or, more commonly in kilojoules (kJ).
2. Although enthalpy *change*,ΔH, can be measured for a reaction, the actual enthalpy, H for the reactant(s) and/or product(s) *cannot* be measured.

HESS'S LAW

Learning Outcomes:

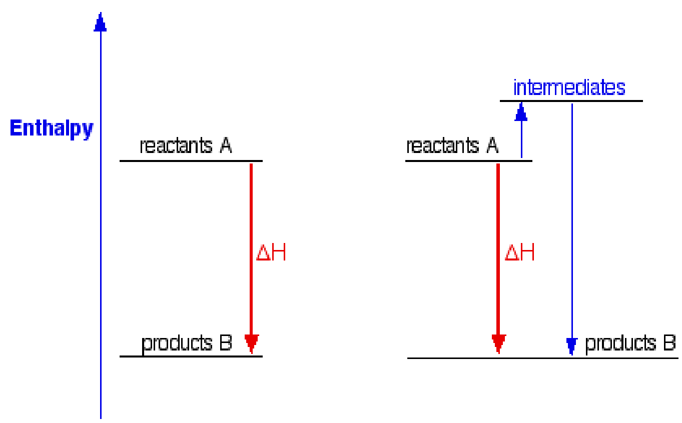
- □ Define Hess's Law (1)
- □ Write the equation for the desired chemical reaction (2)
- □ Rearrange equations approximately to determine energy change of a chemical reaction including the enthalpy changes (3)
- □ Apply Hess's Law to calculate the energy change for a chemical reaction (involves simple steps) (3)
- □ Evaluate the enthalpy change for a reaction with Hess's Law by adding up steps of reaction and stating whether the overall reaction is exothermic or endothermic (4)

Stating Hess's Law

Hess's Law is the most important law in this part of chemistry. Most calculations follow from it. It says . . .

Explaining Hess's Law

- Hess's Law is saying that if you convert reactants **A** into products **B**, the overall enthalpy change will be exactly the same whether you do it in one step or two steps or however many steps.
- If you look at the change on an enthalpy diagram, that is actually fairly obvious.



- This shows the enthalpy changes for an exothermic reaction using two different ways of getting from reactants **A** to products **B**.
- In one case, you do a direct conversion; in the other, you use a two-step process involving some intermediates.
- In either case, the overall enthalpy change must be the same, because it is governed by the relative positions of the reactants and products on the enthalpy diagram.

- If you go via the intermediates, you do have to put in some extra heat energy to start with, but you get it back again in the second stage of the reaction sequence.
- However many stages the reaction is done in, ultimately the overall enthalpy change will be the same, because the positions of the reactants and products on an enthalpy diagram will always be the same.

NOTE:*When building an eqution from enthalpies of combustion or ehtnalpies of formation data, the alterations to the given equations and enthalpy changes are possible:*

1. Multiplying the equation and enthalpy change by a number, for example

$C(s) + O_2(g) \longrightarrow CO_2(g) \quad \Delta H = -394 kJmol^{-1}$
$3C(s) + 3O_2(g) \longrightarrow 3CO_2(g) \quad \Delta H = (3 \times -394) kJmol^{-1}$

2. Reversing the direction of the equation changes the sign of the enthalpy change, for example:

$C_3H_{8(g)} + 5O_{2(g)} \longrightarrow 4H_2O_{(l)} + 3CO_{2(g)} \Delta H = -2217 kJmol^{-1}$
$4H_2O_{(l)} + 3CO_{2(g)} \longrightarrow C_3H_{8(g)} + 5O_{2(g)} \Delta H = +2217 kJmol^{-1}$

<u>Example: Calculating Enthalpy Values</u>

1. Calculate the enthalpy for the formation of propane (C_3H_8),

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

Given the following information:

$C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)} \quad \Delta H = -394 kJmol^{-1}$
$H_{2(g)} + ½O_{2(g)} \longrightarrow H_2O_{(l)} \quad \Delta H = -285 kJmol^{-1}$
$C_3H_{8(g)} + 5O_{2(g)} \longrightarrow 4H_2O_{(l)} + 3CO_{2(g)} \Delta H = -2217 kJmol^{-1}$

Solution:

i) Rearrange each balanced equation with reactants and products on the correct side according to the reaction given as the problem.

$C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)}$

$H_{2(g)} + ½O_{2(g)} \longrightarrow H_2O_{(l)}$

$4H_2O_{(l)} + 3CO_{2(g)} \longrightarrow C_3H_{8(g)} + 5O_{2(g)}$

ii) Assign correct ΔH values for each equation.

$C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)} \quad \Delta H = -394 kJmol^{-1}$
$\mathbf{3C_{(s)} + 3O_{2(g)} \longrightarrow 3CO_{2(g)} \quad \Delta H = -788 kJmol^{-1}}$

$H_{2(g)} + ½O_{2(g)} \longrightarrow H_2O_{(l)} \quad \Delta H = -285 kJmol^{-1}$
$\mathbf{4H_{2(g)} + 2O_{2(g)} \longrightarrow 4H_2O_{(l)} \quad \Delta H = -1140 kJmol^{-1}}$

$4H_2O_{(l)} + 3CO_{2(g)} \longrightarrow C_3H_{8(g)} + 5O_{2(g)} \Delta H = +2217 kJmol^{-1}$

iii) Add the re-arranged equation together.

$3C_{(s)}$ + ~~$3O_{2(g)}$~~ $\longrightarrow$ ~~$3CO_{2(g)}$~~ $\quad \Delta H = -788 kJmol^{-1}$
$4H_{2(g)}$ + ~~$2O_{2(g)}$~~ $\longrightarrow$ ~~$4H_2O_{(l)}$~~ $\quad \Delta H = -1140 kJmol^{-1}$
~~$4H_2O_{(l)}$~~ + ~~$3CO_{2(g)}$~~ $\longrightarrow C_3H_{8(g)}$ + ~~$5O_{2(g)}$~~ $\quad \Delta H = +2217 kJmol^{-1}$

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

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iv) Add the ΔH value

$3C_{(s)}$ + ~~$3O_{2(g)}$~~ $\longrightarrow$ ~~$3CO_{2(g)}$~~ $\quad \Delta H = -788 kJmol^{-1}$
$4H_{2(g)}$ + ~~$2O_{2(g)}$~~ $\longrightarrow$ ~~$4H_2O_{(l)}$~~ $\quad \Delta H = -1140 kJmol^{-}$
^{1}~~$4H_2O_{(l)}$~~ + ~~$3CO_{2(g)}$~~ $\longrightarrow C_3H_{8(g)}$ + ~~$5O_{2(g)}$~~ $\Delta H = +2217 kJmol^{-1}$

$3C_{(s)} + 4H_{2(g)} \longrightarrow C_3H_{8(g)} \quad \Delta H = +289 kJmol^{-1}$

<u>Exercise 10: Calculating Enthalpy Changes</u>

1. When 6g of magnesium burns in excess oxygen, magnesium oxide forms and 150kJ of energy is released.
   Find the heat released if 2 mol of magnesium burns [M(Mg) = 24 $gmol^{-1}$].

2. Define these words:
   a) Enthalpy
   b) Enthalpy change
   c) Activation energy
   d) Exothermic reaction
   e) Endothermic reaction
   f) Energy diagram

3. For the reaction:

   $\mathbf{N_{2\,(g)} + 3H_{2\,(g)} \longrightarrow 2NH_{3(g)}, \quad \Delta H = -92 KJmol^{-1}}$

   a) Find the heat energy released when 2 mol of nitrogen react.
   b) Find the energy released when 0.5 mol of ammonia form.
   c) How much nitrogen must react to produce 1840 kJ of energy?