

Topic 6 – Energetics Revision Notes

1) Introduction

- An enthalpy change is a change in heat energy measured at constant pressure.
- Enthalpy changes refer to the chemicals **not** the surroundings.
- The symbol for an enthalpy change is ΔH (Δ = change, H = heat energy)
- The units for enthalpy change are kJ mol^{-1}

2) Exothermic Reactions

- In exothermic reactions, the chemicals lose energy so ΔH is negative. The surroundings gain energy so feel hotter.
- In terms of a reaction with oxygen, **oxidation** is an exothermic process. Examples include:

Combustion – burning fuels for heating and in engines e.g.
 $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$

Respiration – oxidation of carbohydrates in living things
 $\text{C}_6\text{H}_{12}\text{O}_6 + 6\text{O}_2 \rightarrow 6\text{CO}_2 + 6\text{H}_2\text{O}$

3) Endothermic Reactions

- In endothermic reactions, the chemicals gain energy so ΔH is positive. The surroundings lose energy so feel cooler.
- Endothermic reactions require an input of heat energy or they will stop.
- Examples of endothermic processes include:

Thermal decomposition of calcium carbonate
 $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$

Photosynthesis
 $6\text{CO}_2 + 6\text{H}_2\text{O} \rightarrow \text{C}_6\text{H}_{12}\text{O}_6 + 6\text{O}_2$

4) Measuring Enthalpy Changes

- Measuring enthalpy changes by experiment is called calorimetry
- Calorimetry works by using the energy released or absorbed in a reaction to change the temperature of a known mass of water
- Calculating an enthalpy change from experimental results involves a two step process
- Firstly $q = -mc\Delta T/1000$, where m = mass of water in g, c = specific heat capacity of water ($4.18 \text{ Jg}^{-1}\text{K}^{-1}$), ΔT = change in temperature)
- Secondly, $\Delta H = q/n$ where n = number of moles of reactant
- Enthalpy changes measured from calorimetry are smaller than the expected values because of heat loss to the apparatus and the environment
- Other reasons for differences from standard values are non-standard conditions and evaporation of water
- In the case of measuring enthalpies of combustion, differences from standard values can occur through incomplete combustion and evaporation of the fuel (if it is a liquid)

Example 1:

The combustion of 0.15g of ethanol, C₂H₅OH, in a spirit burner increased the temperature of 75 cm³ of water by 12.5°C. Calculate the enthalpy of combustion of ethanol in kJ mol⁻¹.

$$\begin{aligned}q &= -mc\Delta T/1000 \\ &= -75 \times 4.18 \times 12.5/1000 \\ &= -3.919 \text{ kJ}\end{aligned}$$

$$\begin{aligned}n &= 0.15/46 \\ &= 3.26 \times 10^{-3} \text{ mol}\end{aligned}$$

$$\begin{aligned}\Delta H &= -3.919/3.26 \times 10^{-3} \\ &= -1202 \text{ kJ mol}^{-1}\end{aligned}$$

Example 2:

The heat capacity of a calorimeter and the water it contains is 400 J K⁻¹. A student burned 0.47g of ethanol, C₂H₅OH, and the heat produced increased the temperature of the calorimeter and contents from 19.0°C to 53.3°C. Calculate the enthalpy of combustion of ethanol.

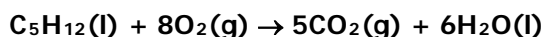
$$\begin{aligned}q &= -c\Delta T/1000 \quad (\text{m not needed if heat capacity rather than specific heat capacity}) \\ &= -400 \times 34.3/1000 \\ &= -13.72 \text{ kJ}\end{aligned}$$

$$\begin{aligned}n &= \text{mass/molar mass} \\ &= 0.47/46.0 \\ &= 0.0102\end{aligned}$$

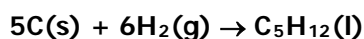
$$\begin{aligned}\Delta H &= q/n \\ &= -13.72/0.0102 \\ &= -1343 \text{ kJ mol}^{-1}\end{aligned}$$

5) Enthalpy Changes

- Standard conditions for measuring enthalpy changes are a pressure of 100 kPa and a temperature of 298K
- Enthalpy change of combustion, ΔH_c , is the enthalpy change when one mole of a substance is completely burned in oxygen under standard conditions e.g.



- Some substances cannot be burnt and have zero enthalpy of combustion e.g. O₂, CO₂, H₂O
- Enthalpy change of formation, ΔH_f , is the enthalpy change when one mole of a substance is formed from its elements under standard conditions e.g.



- ΔH_f for elements is zero

8) Mean Bond Enthalpy

- This is the energy needed to break a covalent bond averaged over many different molecules
- Average bond enthalpies have a positive sign because energy is needed to break a bond
- The strength of a covalent bond depends on the strength of attraction between the shared pair of electrons and the positive nuclei of the atoms

9) Calculating Enthalpy Changes

- Hess's Law = enthalpy change is independent of route
- Enthalpy changes can be calculated in three ways, based on Hess's Law.
- The three ways can be used for to calculate any enthalpy change – formation, combustion, other types of reaction
- **The data provided determines which method to use**
- If the data is enthalpy changes of formation, use:

$$\Delta H = \Sigma \Delta H_f (\text{products}) - \Sigma \Delta H_f (\text{reactants})$$

- If the data is enthalpy changes of combustion, use:

$$\Delta H = \Sigma \Delta H_c (\text{reactants}) - \Sigma \Delta H_c (\text{products})$$

- If the data is bond enthalpies, use:

$$\Delta H = \Sigma(\text{bonds broken}) - \Sigma(\text{bonds formed})$$

- Using average bond enthalpies gives less accurate results than the other two methods because bond enthalpies are average values from a range different compounds

Example – data is enthalpy of formation

Calculate the enthalpy change for the following reaction.

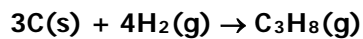


Enthalpies of formation (kJ mol^{-1}) $\text{Li}_2\text{CO}_3(\text{s})$ -1216, $\text{Li}_2\text{O}(\text{s})$ -596, $\text{CO}_2(\text{g})$ -394

$$\begin{aligned}\Delta H &= \Sigma \Delta H_f(\text{products}) - \Sigma \Delta H_f(\text{reactants}) \\ \Delta H &= (-596 + (-394)) - (-1216) \\ &= -990 + 1216 \\ &= 226 \text{ kJ mol}^{-1}\end{aligned}$$

Example – data is enthalpy of combustion

Calculate the enthalpy change for the following reaction.

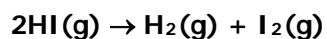


Enthalpies of combustion (kJ mol^{-1}) C(s) -394, $\text{H}_2\text{(g)}$ -286, $\text{C}_3\text{H}_8\text{(g)}$ -2220

$$\begin{aligned}\Delta H &= \Sigma \Delta H_c(\text{reactants}) - \Sigma \Delta H_c(\text{products}) \\ \Delta H &= (3 \times -394) + (4 \times -286) - (-2220) \\ &= -2326 + 2220 \\ &= -106 \text{ kJ mol}^{-1}\end{aligned}$$

Example – data is bond energies

Calculate the enthalpy change for the following reaction.



Bond enthalpies (kJ mol^{-1}) H-I 299, H-H 436, I-I 151

$$\begin{aligned}\Delta H &= \Sigma (\text{bonds broken}) - \Sigma (\text{bonds formed}) \\ \Delta H &= (299 \times 2) - (436 + 151) \\ &= 598 - 587 \\ &= 11 \text{ kJ mol}^{-1}\end{aligned}$$