

# AQA Chemistry A-level

## Topic 1.8 - Thermodynamics

### Flashcards



# What does Hess's Law state?



# What does Hess's Law state?

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



# Define standard enthalpy of formation.



Define standard enthalpy of formation.

The enthalpy change when one mole of a compound is formed from its constituent elements in standard conditions, with all products and reactants in their standard states.



What is the standard  
enthalpy of an element?



What is the standard enthalpy of an element?

Zero, by definition.



# Define standard enthalpy of combustion



## Define standard enthalpy of combustion

The enthalpy change when one mole of a substance is completely burnt in (excess) oxygen



# Define standard enthalpy of atomisation



# Define standard enthalpy of atomisation

Enthalpy change when one mole of gaseous atoms is formed from a compound in its standard state in standard conditions.



# Define first ionisation energy



Define first ionisation energy

Enthalpy change when one mole of electrons is removed from one mole of gaseous atoms to form one mole of gaseous  $1+$  ions.



# Define second ionisation energy.



Define second ionisation energy.

Enthalpy change when one mole of electrons is removed from one mole of gaseous  $1+$  ions to form one mole of gaseous  $2+$  ions



# Define first electron affinity



## Define first electron affinity

Enthalpy change when one mole of gaseous atoms gains one mole of electrons to form one mole of gaseous 1- ions.



Define second electron  
affinity.



Define second electron affinity.

Enthalpy change when one mole of gaseous 1- ions gains one mole of electrons to form one mole of gaseous 2- ions



# Define lattice enthalpy of formation



# Define lattice enthalpy of formation

Enthalpy change when one mole of solid ionic lattice is formed from its constituent gaseous ions.



# Define lattice enthalpy of dissociation.



Define lattice enthalpy of dissociation.

Enthalpy change when one mole of solid ionic lattice is dissociated (broken into) into its gaseous ions



# Define enthalpy of hydration.



Define enthalpy of hydration.

Enthalpy change when one mole of gaseous ions become hydrated/dissolved in water to infinite dilution [water molecules totally surround the ion]



# Define enthalpy of solution



## Define enthalpy of solution

Enthalpy change when one mole of solute dissolves completely in a solvent to infinite dilution.



# Define mean bond dissociation enthalpy



# Define mean bond dissociation enthalpy

Enthalpy change when one mole of (a certain type of) covalent bonds is broken, with all species in the gaseous state



# Write example equations for:

Standard enthalpy of formation  
Standard enthalpy of combustion  
Standard enthalpy of atomisation  
First ionisation energy  
Second ionisation energy  
First electron affinity



# Write example equations for:

Standard enthalpy of formation  $\text{Mg (s)} + \frac{1}{2} \text{O}_2 \text{ (g)} \rightarrow \text{MgO (s)}$

Standard enthalpy of combustion  $\text{CH}_4 \text{ (g)} + 2\text{O}_2 \text{ (g)} \rightarrow \text{CO}_2 \text{ (g)} + 2\text{H}_2\text{O (g)}$

Standard enthalpy of atomisation  $\frac{1}{2}\text{I}_2 \text{ (g)} \rightarrow \text{I (g)}$

First ionisation energy  $\text{Li (g)} \rightarrow \text{Li}^+ \text{ (g)} + \text{e}^-$

Second ionisation energy  $\text{Mg}^+ \text{ (g)} \rightarrow \text{Mg}^{2+} \text{ (g)} + \text{e}^-$

First electron affinity  $\text{Cl (g)} + \text{e}^- \rightarrow \text{Cl}^- \text{ (g)}$



# Write example equations for:

Second electron affinity

Lattice enthalpy of formation

Lattice enthalpy of dissociation

Enthalpy of hydration

Enthalpy of solution

Mean bond dissociation enthalpy



# Write example equations for:

Second electron affinity  $\text{O}^- (\text{g}) + \text{e}^- \rightarrow \text{O}^{2-} (\text{g})$

Lattice enthalpy of formation  $\text{Na}^+ (\text{g}) + \text{Cl}^- (\text{g}) \rightarrow \text{NaCl} (\text{s})$

Lattice enthalpy of dissociation  $\text{NaCl} (\text{s}) \rightarrow \text{Na}^+ (\text{g}) + \text{Cl}^- (\text{g})$

Enthalpy of hydration  $\text{Na}^+ (\text{g}) \rightarrow \text{Na}^+ (\text{aq})$

Enthalpy of solution  $\text{NaCl} (\text{s}) \rightarrow \text{Na}^+ (\text{aq}) + \text{Cl}^- (\text{aq})$

Mean bond dissociation enthalpy  $\text{Br}_2 (\text{g}) \rightarrow 2\text{Br} (\text{g})$



# What is a Born-Haber cycle?



# What is a Born-Haber cycle?

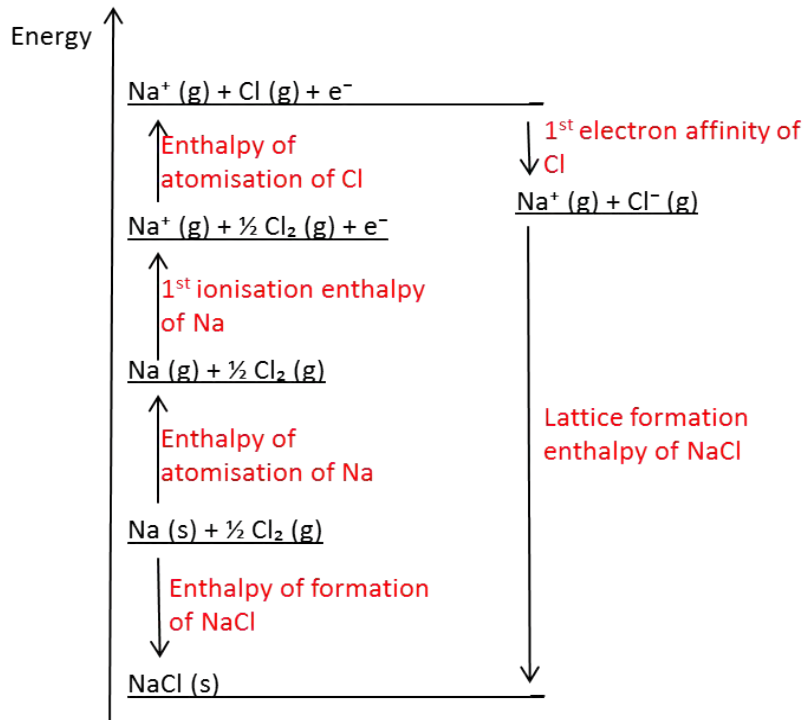
Thermochemical cycle showing all the enthalpy changes involved in the formation of an ionic compound. Start with elements in their standard states (enthalpy of 0)



Draw a labelled Born-Haber cycle for the formation of sodium chloride lattice, including ionisation, electron affinity, formation, lattice formation etc.



Draw a labelled Born-Haber cycle for the formation of sodium chloride lattice, including ionisation, electron affinity, formation, lattice formation etc.



# What factors affect the lattice enthalpy of an ionic compound?



What factors affect the lattice enthalpy of an ionic compound?

Size of the ions, charge on the ions



How can you increase the  
lattice enthalpy of a  
compound? Why does this  
increase it?



# How can you increase the lattice enthalpy of a compound? Why does this increase it?

Smaller ions, since the charge centres will be closer together.

Increased charge, since there will be a greater electrostatic force of attraction between the oppositely charged ions. N.B.

Increasing the charge on the anion has a much smaller effect than increasing the charge on the cation, since increasing anion charge also has the effect of increasing ionic size.



How can Born-Haber cycles  
be used to see if  
compounds could  
theoretically exist?



How can Born-Haber cycles be used to see if compounds could theoretically exist?

Use known data to predict certain values of theoretical compounds, and then see if these compounds would be thermodynamically stable. Was used to predict the existence of the first noble gas containing compound.



What actually happens when a solid is dissolved in terms of interactions of the ions with water molecules?



What actually happens when a solid is dissolved in terms of interactions of the ions with water molecules?

Break lattice → gaseous ions; dissolve each gaseous ion in water. The aqueous ions are surrounded by water molecules (which have a permanent dipole due to polar O-H bond)



# What is the perfect ionic model?



# What is the perfect ionic model?

Assumes that ions are perfectly spherical and that there is an even charge distribution (100% polar bonds). Act as point charges.



# Why is the perfect ionic model often not accurate?



# Why is the perfect ionic model often not accurate?

Ions are not perfectly spherical. Polarisation often occurs when small positive ions or large negative ions are involved, so the ionic bond gains covalent character. Some lattices are not regular and the crystal structure can differ.



Which kind of bonds will be the most ionic? Why?



Which kind of bonds will be the most ionic? Why?

Between large positive ions and small negative ions e.g. CsF



Define the terms  
spontaneous and feasible



## Define the terms spontaneous and feasible

If a reaction is spontaneous and feasible, it will take place of its own accord; does not take account of rate of reaction.



Is a reaction with a positive or negative enthalpy change more likely to be spontaneous?



Is a reaction with a positive or negative enthalpy change more likely to be spontaneous?

Negative - exothermic



# Define entropy



Define entropy

Randomness/disorder of a system.

Higher value for entropy = more  
disordered



# What units is entropy measured in?



What units is entropy measured in?

$\text{JK}^{-1}\text{mol}^{-1}$



# What is the second law of thermodynamics?



# What is the second law of thermodynamics?

Entropy (of an isolated system) always increases, as it is overwhelmingly more likely for molecules to be disordered than ordered



Is a reaction with positive or negative entropy change more likely to be spontaneous?



Is a reaction with positive or negative entropy change more likely to be spontaneous?

Positive - reactions always try and increase the amount of disorder



# Compare the general entropy values for solids, liquids and gases



Compare the general entropy values for solids,  
liquids and gases

Solids < liquids < gases



How would you calculate  
the entropy change for a  
reaction?



How would you calculate the entropy change for a reaction?

Entropy change = sum of products' entropy - sum of reactants' entropy

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



# Define Gibbs free energy using an equation



Define Gibbs free energy using an equation

$\Delta G = \Delta H - T\Delta S$  (G = Gibbs free energy, H = enthalpy change, S = entropy change, T = temperature)



What does the value for  
Gibbs free energy for a  
reaction show?



What does the value for Gibbs free energy for a reaction show?

If  $G < 0$ , reaction is feasible. If  $G = 0$ , reaction is JUST feasible. If  $G > 0$ , reaction is not feasible.



What is the significance of  
the temperature at which  $G$   
 $= 0$ ?



What is the significance of the temperature at which  $G = 0$ ?

This is the temperature (in Kelvin) at which the reaction becomes feasible.



How would you calculate the temperature at which a reaction becomes feasible?



How would you calculate the temperature at which a reaction becomes feasible?

Rearrange to  $T = (\Delta H)/(\Delta S)$  since  $G = 0$



What are the limitations of using  $G$  as an indicator of whether a reaction will occur?



## What are the limitations of using $G$ as an indicator of whether a reaction will occur?

Gibbs free energy only indicates if a reaction is feasible. It does not take into account the rate of reaction (the kinetics of the reaction). In reality, many reactions that are feasible at a certain temperature have a rate of reaction that is so slow that effectively no reaction is occurring.



If the reaction is exothermic and entropy increases, what is the value of  $G$  and what does this mean?



If the reaction is exothermic and entropy increases, what is the value of  $G$  and what does this mean?

$G$  always negative, so reaction is always feasible  
- product favoured



If the reaction is endothermic and entropy decreases, what is the value of  $G$  and what does this mean?



If the reaction is endothermic and entropy decreases, what is the value of  $G$  and what does this mean?

$G$  always positive, so reaction is never feasible - reactant favoured



If the reaction is exothermic and entropy decreases, what is the value of  $G$  and what does this mean?



If the reaction is exothermic and entropy decreases, what is the value of  $G$  and what does this mean?

Temperature dependent



If the reaction is endothermic and entropy increases, what is the value of  $G$  and what does this mean?



If the reaction is endothermic and entropy increases, what is the value of  $G$  and what does this mean?

Temperature dependent



# Why is entropy zero at 0K?



# Why is entropy zero at 0K?

No disorder - molecules/atoms are not moving or vibrating and cannot be arranged in any other way. Maximum possible state of order



What are the two key things  
to look out for to decide if  
entropy  
increases/decreases/stays  
relatively constant?



What are the two key things to look out for to decide if entropy increases/decreases/stays relatively constant?

Number of moles - more moles made → increase in entropy

Going from solid → liquid/gas or liquid → gas



How is it possible for the temperature of a substance undergoing an endothermic reaction to stay constant?



How is it possible for the temperature of a substance undergoing an endothermic reaction to stay constant?

The heat that is given out escapes to the surroundings

