1.8 Thermodynamics
Definitions of enthalpy changes

Enthalpy change of formation
The standard enthalpy change of formation of a compound is the energy transferred when 1 mole of the compound is formed from its elements under standard conditions (298K and 100kpa), all reactants and products being in their standard states

\[ \text{Na}(s) + \frac{1}{2}\text{Cl}_2(g) \rightarrow \text{NaCl}(s) \quad [\Delta H = -411.2 \text{ kJ mol}^{-1}] \]

Enthalpy of atomisation
The enthalpy of atomisation of an element is the enthalpy change when 1 mole of gaseous atoms is formed from the element in its standard state

\[ \text{Na}(s) \rightarrow \text{Na}(g) \quad [\Delta_{\text{sub}} H = +148 \text{ kJ mol}^{-1}] \]

\[ \frac{1}{2} \text{O}_2(g) \rightarrow \text{O}(g) \quad [\Delta_{\text{sub}} H = +249 \text{ kJ mol}^{-1}] \]

Bond dissociation enthalpy (bond energy)
The bond dissociation enthalpy is the standard molar enthalpy change when one mole of a covalent bond is broken into two gaseous atoms (or free radicals)

\[ \text{Cl}_2(g) \rightarrow 2\text{Cl}(g) \quad \Delta_{\text{diss}} H = +242 \text{ kJ mol}^{-1} \]

Or

\[ \text{CH}_4(g) \rightarrow \text{CH}_3(g) + \text{H}(g) \quad \Delta_{\text{diss}} H = +435 \text{ kJ mol}^{-1} \]

First ionisation enthalpy
The first ionisation enthalpy is the enthalpy change required to remove 1 mole of electrons from 1 mole of gaseous atoms to form 1 mole of gaseous ions with a +1 charge

\[ \text{Mg}(g) \rightarrow \text{Mg}^+(g) + e^- \quad [\Delta_{\text{I1}} H] \]

Second ionisation enthalpy
The second ionisation enthalpy is the enthalpy change to remove 1 mole of electrons from one mole of gaseous 1+ ions to produce one mole of gaseous 2+ ions

\[ \text{Mg}^+(g) \rightarrow \text{Mg}^{2+}(g) + e^- \quad [\Delta_{\text{I2}} H] \]

First electron affinity
The first electron affinity is the enthalpy change that occurs when 1 mole of gaseous atoms gain 1 mole of electrons to form 1 mole of gaseous ions with a -1 charge

\[ \text{O}(g) + e^- \rightarrow \text{O}^-(g) \quad [\Delta_{\text{ea1}} H] = -141.1 \text{ kJ mol}^{-1}] \]

The first electron affinity is exothermic for atoms that normally form negative ions. This is because the ion is more stable than the atom, and there is an attraction between the nucleus and the electron.

Enthalpy of lattice formation
The Enthalpy of lattice formation is the standard enthalpy change when 1 mole of an ionic crystal lattice is formed from its constituent ions in gaseous form.

\[ \text{Na}^+(g) + \text{Cl}^-(g) \rightarrow \text{NaCl}(s) \quad [\Delta_{\text{latt}} H = -787 \text{ kJ mol}^{-1}] \]

Enthalpy of lattice dissociation
The Enthalpy of lattice dissociation is the standard enthalpy change when 1 mole of an ionic crystal lattice form is separated into its constituent ions in gaseous form.

\[ \text{NaCl}(s) \rightarrow \text{Na}^+(g) + \text{Cl}^-(g) \quad [\Delta_{\text{latt}} H = +787 \text{ kJ mol}^{-1}] \]

Enthalpy of Hydration \( \Delta_{\text{hydr}} H \)
Enthalpy change when one mole of gaseous ions become aqueous ions.

\[ X^+(aq) + \text{aq}^-(aq) \rightarrow X^+(aq) \quad [\Delta_{\text{hydr}} H = -519 \text{ kJ mol}^{-1}] \]

or

\[ X^-_2(aq) \rightarrow X^-_2(aq) \quad [\Delta_{\text{hydr}} H = -506 \text{ kJ mol}^{-1}] \]

This always gives out energy (exothermic, -ve) because bonds are made between the ions and the water molecules.

Note the conflicting definitions and the sign that always accompanies the definitions

Enthalpy of solution
The enthalpy of solution is the standard enthalpy change when one mole of an ionic solid dissolves in a large enough amount of water to ensure that the dissolved ions are well separated and do not interact with one another.

\[ \text{NaCl}(s) + \text{aq} \rightarrow \text{Na}^+(aq) + \text{Cl}^-(aq) \]
Born Haber cycles

The lattice enthalpy cannot be determined directly. We calculate it indirectly by making use of changes for which data are available and link them together in an enthalpy cycle the Born-Haber cycle.

**Born-Haber cycle: sodium Chloride**

\[
\begin{array}{c}
\text{Na}^+ (g) + e^- + \text{Cl}^- (g) \\
\Delta_a H (\text{Cl}) \\
\text{Na}^+ (g) + \frac{1}{2} \text{Cl}_2(g) \\
\Delta_e a H (\text{Cl}) \\
\text{Na} (g) + \frac{1}{2} \text{Cl}_2(g) \\
\Delta_a H (\text{Na}) \\
\text{Na} (s) + \frac{1}{2} \text{Cl}_2(g) \\
\Delta H (\text{NaCl}) \\
\text{NaCl} (s)
\end{array}
\]

By applying Hess’s law the heat of formation equals to the sum of everything else

\[
\Delta_r H = \Delta_{at} H (\text{Na}) + \Delta_{at} H (\text{Na}) + \Delta_a H (\text{Cl}) + \Delta_e a H (\text{Cl}) + \Delta_{Lat} H
\]

Rearrange to give

\[
\Delta_{Lat} H = \Delta H - (\Delta_{at} H (\text{Na}) + \Delta_{at} H (\text{Na}) + \Delta_a H (\text{Cl}) \Delta_e a H (\text{Cl}) )
\]

\[
\Delta_{Lat} H = -411 - (+107 + 496 + 122 + -349) = -787 \text{ kJ mol}^{-1}
\]

**Born-Haber cycle: magnesium Chloride**

\[
\begin{array}{c}
\text{Mg}^{2+} (g) + 2e^- + 2\text{Cl}^- (g) \\
2 \times \Delta_a H (\text{Cl}) \\
\text{Mg}^{2+} (g) + 2e^- + \text{Cl}_2(g) \\
\Delta_{at} H (\text{Mg}) \\
\text{Mg}^+ (g) + e^- + \text{Cl}_2(g) \\
\Delta_{at} H (\text{Mg}) \\
\text{Mg} (g) + \text{Cl}_2(g) \\
\Delta_a H (\text{Mg}) \\
\text{Mg} (s) + \text{Cl}_2(g) \\
\Delta H (\text{MgCl}_2) \\
\text{MgCl}_2 (s)
\end{array}
\]

The data for the \( \Delta_{at} H (\text{Cl}) \) could also be given as the bond energy for \( \text{E(Cl-Cl)} \) bond.

Remember :

\[
\text{E(Cl-Cl)} = 2 \times \Delta_{at} H (\text{Cl})
\]

Note in this example the first and second ionisation energies of magnesium are needed as Mg is a +2 ion.
### Born Haber cycle: calcium oxide

\[
\begin{align*}
\text{Ca}^{2+} (g) + 2e^- + \text{O}^- (g) &\rightarrow \Delta_{\text{f}}H(\text{CaO}) + \Delta_{\text{ea1}}H(\text{O}) + \Delta_{\text{ea2}}H(\text{O}) \\
\text{Ca}^{2+} (g) + 2e^- + \frac{1}{2} \text{O}_2(g) &\rightarrow \Delta_{\text{IE1}}H(\text{Ca}) \\
\text{Ca} (g) + e^- + \frac{1}{2} \text{O}_2(g) &\rightarrow \Delta_{\text{IE2}}H(\text{Ca}) \\
\text{Ca} (g) + \frac{1}{2} \text{O}_2(g) &\rightarrow \Delta H(\text{CaO}) \\
\text{CaO} (s) &
\end{align*}
\]

### Trends in Lattice Enthalpies

The strength of a enthalpy of lattice formation depends on the following factors:

1. **The sizes of the ions**
   - The larger the ions, the less negative the enthalpies of lattice formation (i.e. a weaker lattice). As the ions are larger the charges become further apart and so have a weaker attractive force between them.

2. **The charges on the ion**
   - The bigger the charge of the ion, the greater the attraction between the ions so the stronger the lattice enthalpy (more negative values).

### Perfect Ionic Model

Theoretical lattice enthalpies assumes a **perfect ionic model** where the ions are **100% ionic** and **spherical** and the attractions are purely electrostatic.

### Differences between theoretical and Born Haber (experimental) lattice enthalpies

The Born Haber lattice enthalpy is the real experimental value. When a compound shows covalent character, the theoretical and the born Haber lattice enthalpies differ. The more the covalent character the bigger the difference between the values.

When the negative ion becomes distorted and more covalent we say it becomes polarised. The metal cation is called polarising if it polarises the negative ion.

### Notice the second electron affinity for oxygen is **endothermic** because it take energy to overcome the **repulsive force** between the negative ion and the electron

In this case the cycle has been constructed using

\[
\Delta H_{\text{lattice dissociation}}
\]

The calculation therefore is:

\[
\left[ \Delta_{\text{f}}H(\text{CaO}) + \Delta H_{\text{lattice dissociation}} \right] = \\
\left[ \Delta_{\text{f}}H(\text{Ca}) + \Delta_{\text{IE1}}H(\text{Ca}) + \Delta_{\text{IE2}}H(\text{Ca}) \\
+ \Delta_{\text{at}}H(\text{O}) + \Delta_{\text{ea1}}H(\text{O}) + \Delta_{\text{ea2}}H(\text{O}) \right]
\]

The lattice enthalpies become less negative down any group.

- e.g. LiCl, NaCl, KCl, RbCl
- e.g. group 1 halides (eg NaF, KI) have lattice enthalpies of around –700 to -1000 kJmol⁻¹
- group 2 halides (eg MgCl₂) have lattice enthalpies of around –2000 to –3500 kJmol⁻¹
- group 2 oxides (eg MgO) have lattice enthalpies of around –3000 to –4500 kJmol⁻¹

There is a tendency towards covalent character in ionic substances when

- the positive ion is small
- the positive ion has multiple charges
- the negative ion is large
- the negative ion has multiple negative charges.

When a compound has some covalent character- it tends towards giant covalent so the lattice is stronger than if it was 100% ionic. Therefore the Born-Haber value would be larger than the theoretical value.

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Why does Calcium chloride have the formula CaCl₂ and not CaCl or CaCl₃?

It is useful to draw out the born haber cycles for each potential case.

We need to calculate an enthalpy of formation for each case. The one with the most exothermic enthalpy of formation will be the one that forms as it will be the most thermodynamically stable.

Theoretical lattice enthalpies have been calculated for each case:

- \( \Delta H_{\text{latt}} \text{CaCl} = -719 \text{ kJ mol}^{-1} \)
- \( \Delta H_{\text{latt}} \text{CaCl}_2 = -2218 \text{ kJ mol}^{-1} \)
- \( \Delta H_{\text{latt}} \text{CaCl}_3 = -4650 \text{ kJ mol}^{-1} \)

These get larger as the positive charge on the calcium ion becomes bigger.

The enthalpy of formation is largely a balance of the ionisation energy and lattice enthalpy.

\( \Delta H (\text{CaCl}) \) is \(-163.1 \text{ kJ mol}^{-1} \). This is exothermic.

The increased ionisation enthalpy to form \( \text{Ca}^{2+} \) is more than compensated for by the stronger lattice enthalpy of formation. The enthalpy of formation is therefore more exothermic. This is the most stable form.

\( \Delta H (\text{CaCl}_2) \) is \(-739.2 \text{ kJ mol}^{-1} \). This is exothermic.

The big increase in ionisation enthalpy to remove the 3rd electron is not compensated for by the stronger lattice enthalpy of formation. The enthalpy of formation is therefore endothermic. This is the least stable form.
**Free-energy change (ΔG) and entropy change (ΔS)**

A SPONTANEOUS PROCESS (e.g. diffusion) will proceed on its own without any external influence.

A problem with ΔH

A reaction that is exothermic will result in products that are more thermodynamically stable than the reactants. This is a driving force behind many reactions and causes them to be spontaneous (occur without any external influence).

Some spontaneous reactions, however, are endothermic. How can this be explained?

We need to consider something called entropy.

**Entropy, S˚**

Substances with more ways of arranging their atoms and energy (more disordered) have a higher entropy.

<table>
<thead>
<tr>
<th>Elements</th>
<th>Compounds</th>
</tr>
</thead>
<tbody>
<tr>
<td>Simpler compounds</td>
<td>Complex compounds</td>
</tr>
<tr>
<td>Pure substances</td>
<td>Mixtures</td>
</tr>
</tbody>
</table>

Solids have lower entropies than liquids, which are lower than gases.

- When a solid increases in temperature its entropy increases as the particles vibrate more.
- There is a bigger jump in entropy with boiling than that with melting.
- Gases have large entropies as they are much more disordered

**Predicting Change in entropy ‘ΔS˚’ Qualitatively**

An increase in disorder and entropy will lead to a positive entropy change $\Delta S˚ = +ve$.

Balanced chemical equations can often be used to predict if $\Delta S˚$ is positive or negative.

$\text{NH}_3\text{Cl} (s) \rightarrow \text{HCl} (g) + \text{NH}_3 (g)$

$\Delta S˚ = +ve$

- change from solid reactant to gaseous products
- increase in number of molecules
- both will increase disorder

$\text{Na (s) + }\frac{1}{2}\text{ Cl}_2 (g) \rightarrow \text{NaCl (s)}$

$\Delta S˚ = -ve$

- change from gaseous and solid reactant to solid
- decrease in number of molecules
- both will decrease disorder

**Calculating ΔS˚ quantitatively**

Data books lists standard entropies (S˚) per mole for a variety of substances. It is not possible for a substance to have a standard entropy of less than zero.

Elements in their standard states do not have zero entropy. Only perfect crystals at absolute zero ($T = 0 K$) will have zero entropy.

The unit of entropy is $J K^{-1} mol^{-1}$

$\Delta S˚ = \Sigma S˚\text{products} - \Sigma S˚\text{reactants}$
Example

Calculate $\Delta S$ for the following reaction at 25°C:

$$2\text{Fe}_2\text{O}_3 (s) + 3\text{C} (s) \rightarrow 4\text{Fe} (s) + 3\text{CO}_2 (g)$$

$\Delta S = S_{\text{products}} - S_{\text{reactants}} = (3 x 213.6 + 4 x 27.3) - (2 x 87.4 + 3 x 5.7) = +558.1 \text{ J K}^{-1} \text{ mol}^{-1}$ (3 S.F.)

Note: the entropy change is very positive as a large amount of gas is being created, increasing disorder

Gibbs Free Energy Change, $\Delta G$

The balance between entropy and enthalpy determines the feasibility of a reaction. This is given by the relationship:

$$\Delta G = \Delta H - T\Delta S$$

For any spontaneous change, $\Delta G$ will be negative.

Units: KJ mol$^{-1}$

Need to convert $S$ to KJ mol$^{-1}$(÷1000)

Converting from°C to K (+273)

$$\Delta G = \Delta H - T\Delta S$$

Example: Data for the following reaction, which represents the reduction of aluminium oxide by carbon, are shown in the table.

$$\text{Al}_2\text{O}_3(s) + 3\text{C}(s) \rightarrow 2\text{Al}(s) + 3\text{CO}(g)$$

Calculate the values of $\Delta H$, $\Delta S$, and $\Delta G$ for the above reaction at 298 K

1. Calculate $\Delta S$

$$\Delta S = \sum S_{\text{products}} - \sum S_{\text{reactants}}$$

$$\Delta S = (2 x 28 + 3 x 198) - (51 + 3 x 5.7) = +581 \text{ J K}^{-1} \text{ mol}^{-1}$$

2. Calculate $\Delta H$

$$\Delta H = \sum \Delta f_{\text{H}} [\text{products}] - \sum \Delta f_{\text{H}} [\text{reactants}]$$

$$\Delta H = (3 x -111) - (-1669) = +1336 \text{ kJ mol}^{-1}$$

3. Calculate $\Delta G$

$$\Delta G = \Delta H - T\Delta S$$

$$\Delta G = +1336 - 298 x 0.581 = +1163 \text{ kJ mol}^{-1}$$

$\Delta G$ is positive. The reaction is not feasible.

Calculating the temperature a reaction will become feasible

Calculate the temperature range that this reaction will be feasible.

$$\text{N}_2(g) + \text{O}_2(g) \rightarrow 2 \text{NO}(g)$$

$\Delta H = 180 \text{ kJ mol}^{-1}$ $\Delta S = 25 \text{ J K}^{-1} \text{ mol}^{-1}$

The reaction will be feasible when $\Delta G \leq 0$

Make $\Delta G = 0$ in the following equation $\Delta G = \Delta H - T\Delta S$

$0 = \Delta H - T\Delta S$

So $T = \Delta H / \Delta S$

$T = 180 / (25/1000)$

$= 7200$K

If $\Delta G$ is negative, there is still a possibility, however, that the reaction will not occur or will occur so slowly that effectively it doesn’t happen. If the reaction has a high activation energy the reaction will not occur.

$\Delta G$ during phase changes

As physical phase changes like melting and boiling are equilibria, the $\Delta G$ for such changes is zero.

What temperature would methane melt at?

$$\text{CH}_4(\text{g}) \rightarrow \text{CH}_4(\text{l})$$

$\Delta H = 0.94 \text{ kJ mol}^{-1}$ $\Delta S = 10.3 \text{ J mol}^{-1} \text{K}^{-1}$

Make $\Delta G = 0$ in the following equation $\Delta G = \Delta H - T\Delta S$

$0 = \Delta H - T\Delta S$

So $T = \Delta H / \Delta S$

$T = 0.94 / (10.3 x 1000)$

$T = 91$K

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Effect of temperature on feasibility

\[ \Delta G = \Delta H - T\Delta S \]

Changing temperature will change the value of \(-T\Delta S\) in the above equation.

If the reaction involves an increase in entropy (\(\Delta S\) is positive) then increasing temperature will make it more likely that \(\Delta G\) is negative and more likely that the reaction occurs. E.g., \(\text{NaCl} + \text{aq} \rightarrow \text{Na}^+_{(aq)} + \text{Cl}^-_{(aq)}\)

If the reaction involves a decrease in entropy (\(\Delta S\) is negative) then increasing temperature will make it less likely that \(\Delta G\) is negative and less likely for the reaction to occur. E.g., \(\text{HCl}(g) + \text{NH}_3(g) \rightarrow \text{NH}_4\text{Cl}(s)\)

If the reaction has a \(\Delta S\) close to zero then temperature will not have a large effect on the feasibility of the reaction, as \(-T\Delta S\) will be small and \(\Delta G\) won’t change much. E.g., \(\text{N}_2(g) + \text{O}_2(g) \rightarrow 2\text{NO}(g)\)

This graph shows how the free-energy change for formation of ammonia varies with temperature above 240 K. \(\frac{1}{2} \text{N}_2(g) + \frac{3}{2} \text{H}_2(g) \rightarrow \text{NH}_3(g)\)

Applying the equation of a straight line

\[ y = mx + c \]

\[ c = \Delta H \]

The gradient of this graph is equal to \(-\Delta S\)

The positive gradient means \(\Delta S\) is negative which corresponds to the equation above showing increasing order.

When \(\Delta G < 0\) then the reaction is spontaneous. In this case at temperatures below around 460K, the slope of the line would change below 240K because ammonia would be a liquid and the entropy change would be different.

Enthalpies of solution

Using Hess’s law to determine enthalpy changes of solution

When an ionic substance dissolves in water, it involves breaking up the bonds in the lattice and forming new bonds between the metal ions and water molecules.

For \(\text{MgCl}_2\) the ionic equation for the dissolving is \(\text{MgCl}_2(s) + \text{aq} \rightarrow \text{Mg}^{2+}_{(aq)} + 2\text{Cl}^-_{(aq)}\)

- \(\Delta H_{\text{lattice dissociation (MgCl}_2)}\)
- \(\Delta_{\text{hyd}} H_{\text{Mg}^{2+}} + 2 \times \Delta_{\text{hyd}} H_{\text{Cl}^-}\)
- \(H_{\text{solution}} \rightarrow \text{Mg}^{2+}_{(aq)} + 2\text{Cl}^-_{(aq)}\)

In general

\[ \Delta H_{\text{solution}} = \Delta H_{\text{L dissociation}} + \Sigma \Delta_{\text{hyd}} H \]

When an ionic substance dissolves the lattice must be broken up. The enthalpy of lattice dissociation is equal to the energy needed to break up the lattice (to gaseous ions). This step is endothermic.

The size of the lattice enthalpy depends on the size and charge of the ion. The smaller the ion and the higher its charge, the stronger the lattice.

OR

\[ \Delta H_{\text{solution}} = - \Delta H_{\text{L formation}} + \Sigma \Delta_{\text{hyd}} H \]

Sometimes in questions \(\Delta H_{\text{L dissociation}}\) is given instead of \(\Delta H_{\text{L formation}}\) in order to catch you out. Remember the difference between the two.
Example. Calculate the enthalpy of solution of NaCl given that the lattice enthalpy of formation of NaCl is $-771 \text{ kJmol}^{-1}$ and the enthalpies of hydration of sodium and chloride ions are $-406$ and $-364 \text{ kJmol}^{-1}$ respectively.

$$
\Delta_{\text{sol}} H = - \Delta H_{\text{lattice, formation}} + \Sigma \Delta \text{hyd} H \\
= (-771) + (-406-364) \\
= + 1 \text{ kJmol}^{-1}
$$

Hydration enthalpies are **exothermic** as energy is given out when water molecules bond to the metal ions.

The negative ions are attracted to the $\delta^+$ hydrogens on the polar water molecules and the positive ions are attracted to the $\delta^-$ oxygen on the polar water molecules.

The higher the **charge density** the greater the hydration enthalpy (e.g. smaller ions or ions with larger charges) as the ions attract the water molecules more strongly.

e.g. Fluoride ions have more negative hydration enthalpies than chloride ions.

Magnesium ions have a more negative hydration enthalpy than barium ions.

**What does $\Delta H_{\text{solution}}$ tell us?**

Generally $\Delta H_{\text{solution}}$ is not very exo or endothermic so the hydration enthalpy is about the same as lattice enthalpy.

In general the substance is more likely to be **soluble** if the $\Delta H_{\text{solution}}$ is **endothermic**.

If a substance is insoluble it is often because the lattice enthalpy is much larger than the hydration enthalpy and it is not energetically favourable to break up the lattice, making $\Delta H_{\text{solution}}$ **endothermic**.

We must consider **entropy**, however, to give us the full picture about solubility.

When a solid dissolves into ions the **entropy increases** as there is more **disorder** as solid changes to solution and the **number of particles increases**.

This positive $\Delta S$ can make $\Delta G$ negative even if $\Delta H$ solution is endothermic, especially at higher temperatures.

For salts where $\Delta H_{\text{solution}}$ is exothermic the salt will always dissolve at all temperatures.

$$
\Delta G = \Delta H - T\Delta S \\
\Delta G \text{ is always negative} \\
\Delta H \text{ is negative} \\
\Delta S \text{ is positive due to the increased disorder as more particles so - T}\Delta S \text{ always negative}
$$

For salts where $\Delta H_{\text{solution}}$ is endothermic the salt may dissolve depending on whether the $-T\Delta S$ value is more negative than $\Delta H$ is positive

$$
\Delta G = \Delta H - T\Delta S \\
\Delta G \text{ is positive} \\
\Delta H \text{ is positive} \\
\Delta S \text{ is positive due to the increased disorder as more particles so - T}\Delta S \text{ always negative}
$$

Increasing the temperature will make it more likely that $\Delta G$ will become negative, making the reaction feasible and the salt will dissolve.