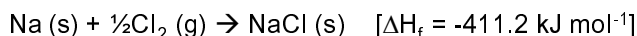# 1: THERMODYNAMICS

Definitions of enthalpy changes

Enthalpy change of formation
The enthalpy change when 1 mole of a compound is formed from its elements in their standard states under standard conditions

$Na (s) + \frac{1}{2}Cl_2 (g) \rightarrow NaCl (s)$ [ΔH_f = -411.2 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

$Na (s) \rightarrow Na(g)$ [ΔH_{at} = +148 kJ mol^{-1}]
$\frac{1}{2} O_2 (g) \rightarrow O (g)$ [ΔH_{at} = +249 kJ mol^{-1}]
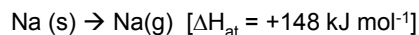
The enthalpy change for a solid metal turning to gaseous atoms can also be called the **Enthalpy of sublimation** and will numerically be the same as the enthalpy of atomisation
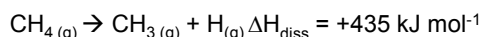
$Na (s) \rightarrow Na(g)$ [$\Delta Hsub$ = +148 kJ mol^{-1}]
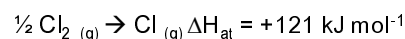

Bond dissociation enthalpy
The bond dissociation enthalpy is the standard molar enthalpy change when one mole of a covalent bond is broken into two gaseous atoms (or free radicals)

$Cl_{2\ (g)} \rightarrow 2Cl_{\ (g)}$ ΔH_{diss} = +242 kJ mol^{-1}
Or
$CH_{4\ (g)} \rightarrow CH_{3\ (g)} + H_{(g)}$ ΔH_{diss} = +435 kJ mol^{-1}

For diatomic molecules the $\Delta Hdiss$ of the molcule is the same as 2x ΔH_{at} of the element

$Cl_{2\ (g)} \rightarrow 2Cl_{\ (g)}$ ΔH_{diss} = +242 kJ mol^{-1}
$\frac{1}{2} Cl_{2\ (g)} \rightarrow Cl_{\ (g)} \Delta H_{at}$ = +121 kJ mol^{-1}

First Ionisation energy
The first ionisation energy is the energy required to remove 1 mole of electrons from 1 mole of gaseous atoms to form 1 mole of gaseous ions with a +1 charge (under standard conditions)
$Mg (g) \rightarrow Mg^+ (g) + e^-$ [ΔH Ie]
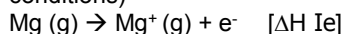
Second Ionisation energy
The second ionisation energy is the energy required to remove 1 mole of electrons from one mole of gaseous 1+ ions to produces one mole of gaseous 2+ ions.
$Mg^+ (g) \rightarrow Mg^{2+} (g) + e^-$ [Δ Hi]
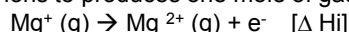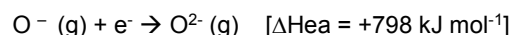

First Electron affinity
The first electron affinity is the enthalpy change that occurs when 1 mole of gaseous atoms gain 1 mole of electrons to form 1 mole of gaseous ions with a –1 charge (under standard conditions)

$O (g) + e^- \rightarrow O^- (g)$ [ΔHea] = -141.1 kJ mol^{-1}]

The first electron affinity is exothermic for atoms that normally form negative ions because the ion is more stable than the atom

second electron affinity
The second electron affinity is the enthalpy change when one mole of gaseous 1- ions gains one electron per ion to produce gaseous 2- ions.

$O^- (g) + e^- \rightarrow O^{2-} (g)$ [ΔHea = +798 kJ mol^{-1}]

The second electron affinity for oxygen is **endothermic** because it take energy to overcome the **repulsive force** between the **negative ion 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 (under standard conditions).

$Na^+(g) + Cl^- (g) \rightarrow NaCl (s)$ [ΔH_{Latt} = -787 kJ mol^{-1}]
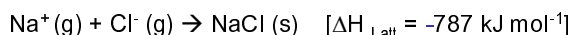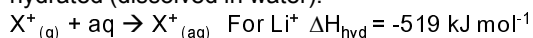
Enthalpy of lattice dissociation
The Enthalpy of lattice dissociation is the standard enthalpy change when 1 mole of an ionic crystal lattice is separated into its constituent ions in gaseous form (under standard conditions).
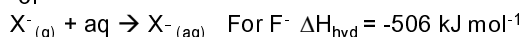
$NaCl (s) \rightarrow Na^+(g) + Cl^- (g)$ [ΔH_{Latt} = +787 kJ mol^{-1}]
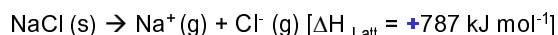
Note the conflicting definitions and the sign that always accompanies the definitions

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Enthalpy of Hydration ΔH_{hyd}
Enthalpy change when one mole of gaseous ions become hydrated (dissolved in water).
$X^+_{\ (g)} + aq \rightarrow X^+_{\ (aq)}$ For Li$^+$ ΔH_{hyd} = -519 kJ mol^{-1}
or
$X^-_{\ (g)} + aq \rightarrow X^-_{\ (aq)}$ For F$^-$ ΔH_{hyd} = -506 kJ mol^{-1}

This always gives out energy (exothermic, -ve) because bonds are made between the ions and the water molecules
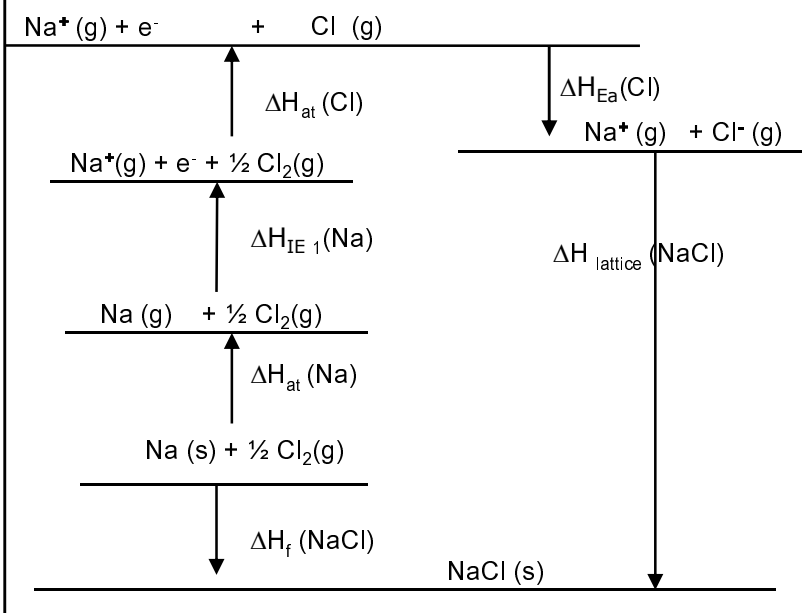
Enthalpy of solution
The enthalpy of solution is the standard enthalpy change when one mole of an ionic solid dissolves in an large enough amount of water to ensure that the dissolved ions are well separated and do not interact with one another

$NaCl (s) + aq \rightarrow Na^+ (aq) + Cl^-(aq)$

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

Na^+ (g) + e^- +   Cl (g)

ΔH_{at} (Cl)

ΔH_{Ea}(Cl)

Na^+ (g) + Cl^- (g)

Na^+(g) + e^- + ½ Cl_2(g)

$\Delta H_{IE\ 1}$(Na)

$\Delta H_{lattice}$ (NaCl)

Na (g)   + ½ Cl_2(g)

ΔH_{at} (Na)

Na (s) + ½ Cl_2(g)

ΔH_f (NaCl)

NaCl (s)

Pay attention to state symbols and direction of arrows.

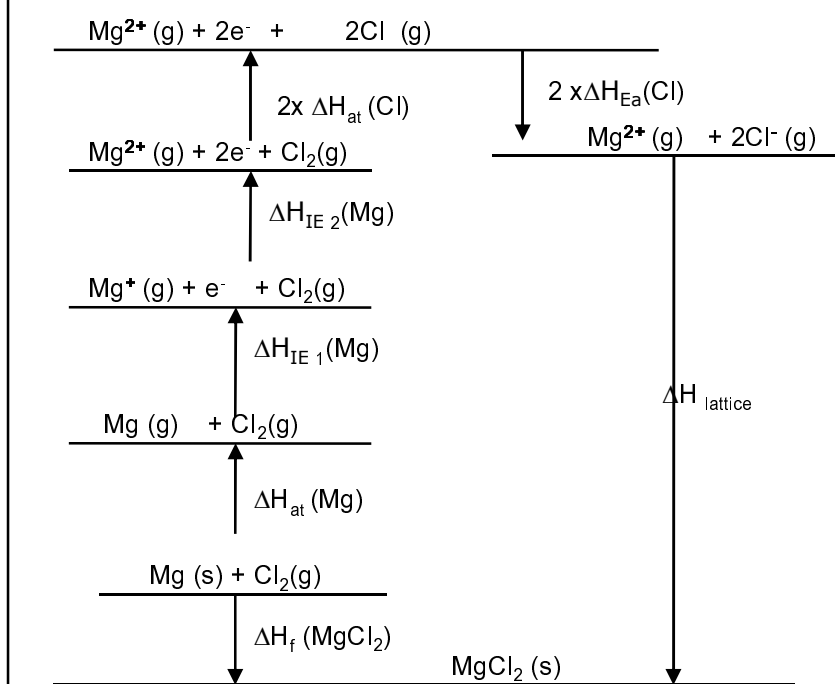Usually all pieces of data are given except the one that needs to be calculated

By applying hess's law the heat of formation equals to the sum of everything else

$\Delta H_f = \Delta H_{at}Na + \Delta H_{IE}Na + \Delta H_{at}Cl + \Delta H_{Ea}Cl + \Delta H_{latt}$

Rearrange to give $\Delta H_{latt} = \Delta H_{formation} - (\Delta H_{atm}Na + \Delta H_{IE}Na + \Delta H_{atm}Cl + \Delta H_{Ea}Cl)$

$\Delta H_{lattice} = -411 - (+107 + 496 + 122 + -349) = -787 \ kJmol^{-1}$
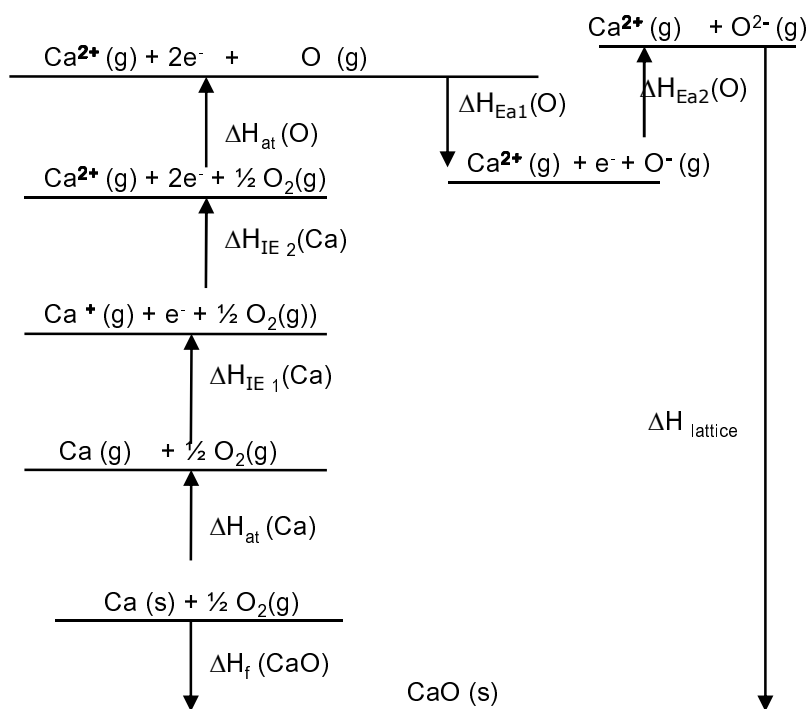
Born Haber cycle: magnesium Chloride

Mg^{2+} (g) + $2e^-$ +   2Cl (g)

2x ΔH_{at} (Cl)

2 xΔH_{Ea}(Cl)

Mg^{2+} (g)   + $2Cl^-$ (g)

Mg^{2+} (g) + $2e^-$ + Cl_2(g)

$\Delta H_{IE\ 2}$(Mg)

Mg^+ (g) + e^- + Cl_2(g)

$\Delta H_{IE\ 1}$(Mg)

$\Delta H_{lattice}$

Mg (g)   + Cl_2(g)

ΔH_{at} (Mg)

Mg (s) + Cl_2(g)

ΔH_f ($MgCl_2$)

$MgCl_2$ (s)

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

Remember :

E(Cl-Cl) = 2 x ΔH_{at} (Cl)

Born Haber cycle: calcium oxide

Ca^{2+} (g) + 2e$^-$ + O (g)

Ca^{2+} (g) + O^{2-} (g)

ΔH_{at} (O)

ΔH_{Ea1}(O)

ΔH_{Ea2}(O)

Ca^{2+} (g) + 2e$^-$ + ½ O_2(g)

Ca^{2+} (g) + e$^-$ + O$^-$ (g)

$\Delta H_{IE\ 2}$(Ca)

Ca $^+$ (g) + e$^-$ + ½ O_2(g))

$\Delta H_{IE\ 1}$(Ca)

$\Delta H_{lattice}$

Ca (g) + ½ O_2(g)

ΔH_{at} (Ca)

Ca (s) + ½ O_2(g)

ΔH_f (CaO)

CaO (s)

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**

Trends in Lattice Enthalpies

The strength of a lattice enthalpy depends on the following factors

1.   **The sizes of the ions:**

   The larger the ions, the less negative the lattice enthalpy (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).

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

group 2 halides (eg $MgCl_2$) have lattice enthalpies of around −2000 to −3500

group 2 oxides eg MgO have lattice enthalpies of around −3000 to −4500 kJmol^{-1}

Differences between theoretical and Born Haber lattice enthalpies
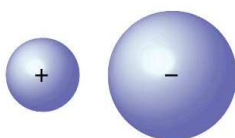
Theoretical lattice enthalpies assume the ions are 100% ionic and spherical. The Born Haber lattice enthalpy is the real experimental value.

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.

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

100% ionic
The theoretical and the born Haber lattice enthalpies will be the same
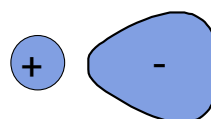
Ionic with covalent character
The theoretical and the Born Haber lattice enthalpies will differ

Why does Calcium chloride have the formula CaCl$_2$ and not CaCl or CaCl$_3$?

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

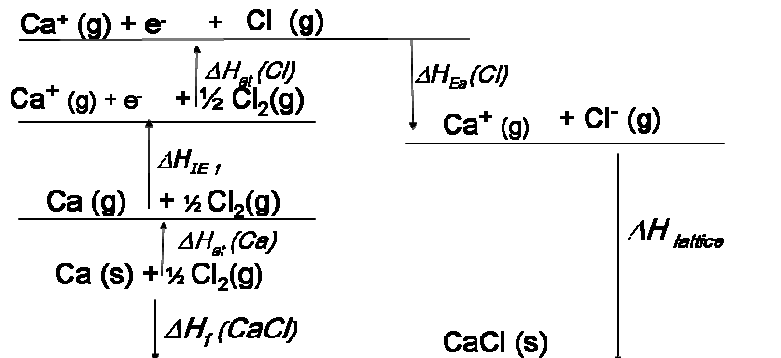Theoretical lattice enthalpies have been calculated for each case

Theoretical lattice enthalpies
ΔH_{latt} CaCl = -719 kJ mol^{-1}
ΔH_{latt} CaCl$_2$ = -2218 kJ mol^{-1}
ΔH_{latt} CaCl$_3$ = -4650 kJ mol^{-1}

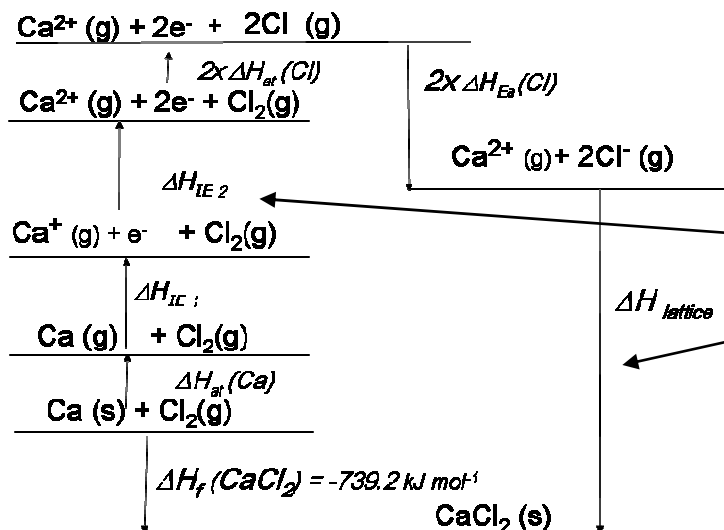These get larger as the positive charge on the calcium ion becomes bigger.

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

ΔH_f (CaCl) is -163.1 kJ mol^{-1}. This is exothermic.
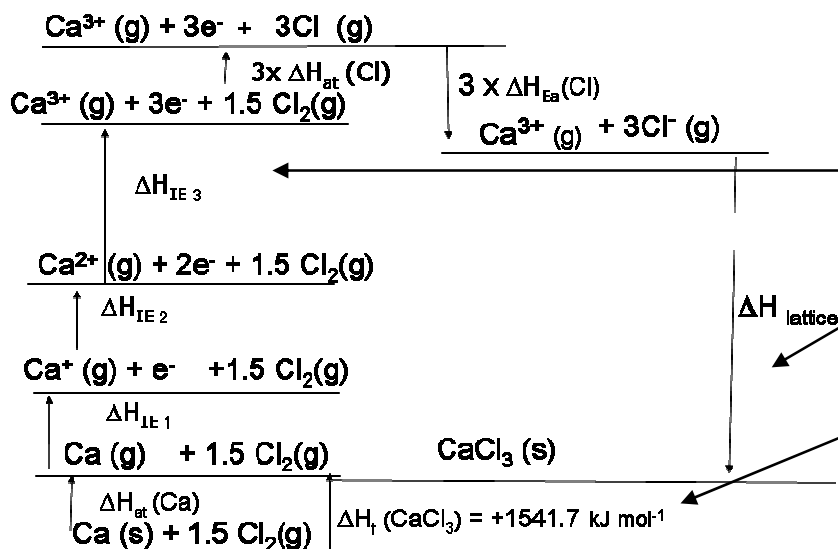
Born Haber Cycle for CaCl

Ca^+ (g) + e$^-$ + Cl (g)

ΔH_{at} (Cl)

ΔH_{Ea}(Cl)

Ca^+ (g) + e$^-$ + ½ Cl$_2$(g)

Ca^+ (g) + Cl$^-$ (g)

$\Delta H_{IE\,1}$

Ca (g) + ½ Cl$_2$(g)

ΔH_{at} (Ca)

$\Delta H_{lattice}$

Ca (s) + ½ Cl$_2$(g)

ΔH_f (CaCl)

CaCl (s)

Born Haber Cycle for CaCl$_2$

Ca^{2+} (g) + 2e$^-$ + 2Cl (g)

2x ΔH_{at} (Cl)

2x ΔH_{Ea}(Cl)

Ca^{2+} (g) + 2e$^-$ + Cl$_2$(g)

Ca^{2+} (g) + 2Cl$^-$ (g)

$\Delta H_{IE\,2}$

Ca^+ (g) + e$^-$ + Cl$_2$(g)

$\Delta H_{IE\,1}$

Ca (g) + Cl$_2$(g)

ΔH_{at} (Ca)

$\Delta H_{lattice}$

Ca (s) + Cl$_2$(g)

ΔH_f (CaCl$_2$) = -739.2 kJ mol^{-1}

CaCl$_2$ (s)

The increased ionisation enthalpy to form Ca^{2+} is more than compensated for by the stronger lattice enthalpy

The enthalpy of formation is therefore more exothermic. This is the most stable form
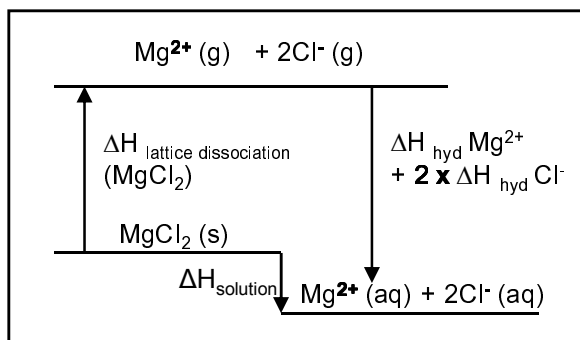
Born Haber Cycle for CaCl$_3$

Ca^{3+} (g) + 3e$^-$ + 3Cl (g)

3x ΔH_{at} (Cl)

3 x ΔH_{Ea}(Cl)

Ca^{3+} (g) + 3e$^-$ + 1.5 Cl$_2$(g)

Ca^{3+} (g) + 3Cl$^-$ (g)

$\Delta H_{IE\,3}$

Ca^{2+} (g) + 2e$^-$ + 1.5 Cl$_2$(g)

$\Delta H_{IE\,2}$

Ca^+ (g) + e$^-$ + 1.5 Cl$_2$(g)

$\Delta H_{IE\,1}$

Ca (g) + 1.5 Cl$_2$(g)

ΔH_{at} (Ca)

$\Delta H_{lattice}$

Ca (s) + 1.5 Cl$_2$(g)

ΔH_f (CaCl$_3$) = +1541.7 kJ mol^{-1}

CaCl$_3$ (s)

The big increase in ionisation enthalpy to remove the 3rd electron is not compensated for by the stronger lattice enthalpy

The enthalpy of formation is therefore endothermic. This is the least stable form

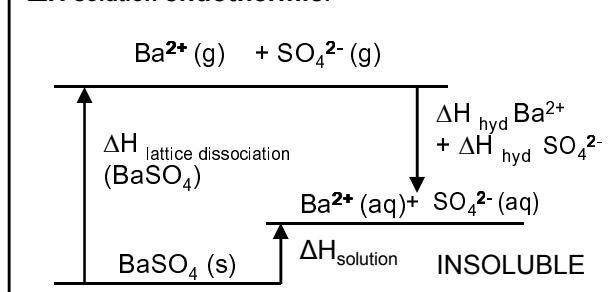## Using Hess's law to determine enthalpy changes of solution

Mg^{2+} (g)   + 2Cl$^-$ (g)

$\Delta H_{\text{lattice dissociation}}$ (MgCl$_2$)

ΔH_{hyd} Mg^{2+} + 2 x ΔH_{hyd} Cl$^-$

MgCl$_2$ (s)

$\Delta H_{\text{solution}}$ Mg^{2+} (aq) + 2Cl$^-$ (aq)

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.**

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

The higher the **charge density** the greater the hydration enthalpy (e.g. smaller ions or ions with larger charges)

In general

$\Delta H \text{ sol} = \Delta H_{\text{L dissociation}} + \Sigma \Delta H_{\text{hyd}}$

OR

$\Delta H \text{ sol} = - \Delta H_{\text{L formation}} + \Sigma \Delta H_{\text{hyd}}$

Sometimes in questions $\Delta H_{\text{L formation}}$ is given instead of $\Delta H_{\text{L dissociation}}$ in order to catch you out. Remember the difference between the two.

ΔH solution endothermic.

Ba^{2+} (g)   + SO$_4$$^{2-}$ (g)

$\Delta H_{\text{lattice dissociation}}$ (BaSO$_4$)

ΔH_{hyd} Ba^{2+} + ΔH_{hyd} SO$_4$$^{2-}$

Ba^{2+} (aq)+ SO$_4$$^{2-}$ (aq)

BaSO$_4$ (s) $\Delta H_{\text{solution}}$ INSOLUBLE

What does $\Delta H_{\text{Solution}}$ tell us?
Generally **ΔH solution** is not very exo or endothermic so the hydration enthalpy is about the same as lattice enthalpy.
In general the substance is thought to **be soluble** if the **ΔH solution is exothermic**
If a substance is insoluble it is often because the lattice enthalpy is larger than the hydration enthalpy and it is not energetically favourable to break up the lattice, making **ΔH solution endothermic**.
We must consider entropy, however, to give us the full picture about solubility

Example 1. Calculate the enthalpy of solution of NaCl given that the lattice enthalpy of formation of NaCl is -771 kJmol^{-1} and the enthalpies of hydration of sodium and chloride ions are -406 and -364 kJmol^{-1} respectively

$\Delta H \text{ sol} = - \Delta H_{\text{L formation}} + \Sigma \Delta H_{\text{hyd}}$

= - (-771) + (-406-364)

= + 1 kJmol^{-1}

Mean Bond energies

The Mean bond energy is the enthalpy needed **to break** the bond into **gaseous atoms**, **averaged over different** molecules

We use values of **mean** bond energies because in reality every single bond in a compound has a slightly different value. E.g. In CH4 there are 4 C-H bonds. Breaking each one will require a different amount of energy. However, we use an average value for the C-H bond for all hydrocarbons.
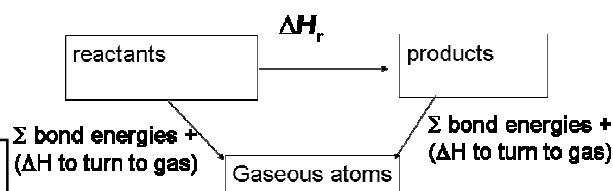
These values are positive because energy is required to break a bond.

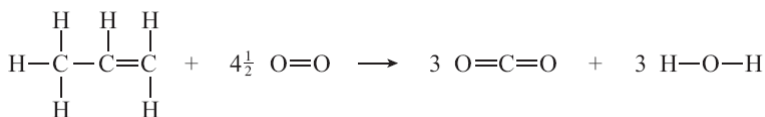The definition only applies when the **substances start and end in the gaseous state.**

In general (if all substances are gases**)**

ΔHr = Σ **bond energies broken** - Σ **bond energies made**

ΔH values calculated using this method will be less accuate than using formation or combustion data because the mean bond energies are not exact

ΔH_r

reactants → products

Σ bond energies + (ΔH to turn to gas)

Σ bond energies + (ΔH to turn to gas)

Gaseous atoms

Example 2. Using the following mean bond enthalpy data to calculate the heat of combustion of propene

Bond	Mean enthalpy (kJ mol-1)
C=C	612
C-C	348
O=O	496
O=C	743
O-H	463
C-H	412

$H-\overset{\overset{\displaystyle H}{|}}{\underset{\underset{\displaystyle H}{|}}{C}}-\overset{\overset{\displaystyle H}{|}}{C}=\overset{\overset{\displaystyle H}{|}}{\underset{\underset{\displaystyle H}{|}}{C}}$ + $4\frac{1}{2}$ O=O $\longrightarrow$ 3 O=C=O + 3 H—O—H

ΔH = Σ bond energies broken - Σ bond energies made

= [E(C=C) + E(C-C) + 6 x E(C-H) + 4.5 x E(O=O)] – [6 xE(C=O) + 6 E(O-H)]

= [612 + 348 + (6 x 412) + (4.5 x 496)] – [(6 x 743) + (6 X 463)]

= - 1752 kJmol^{-1}

Example 3. Using the following mean bond enthalpy data to calculate the heat of formation of NH_3

½ N_2 + 1.5 H_2 → NH_3 (note the balancing is to agree with the definition of heat of formation (i.e. one mole of product)

E(N≡N) = 944 kJ mol^{-1} E(H-H) = 436 kJ mol^{-1} E(N-H) = 388 kJ mol^{-1}

ΔH = Σ bond energies broken - Σ bond energies made

= [0.5 x E(N≡N) + 1.5 x E(H-H)]   – [3 xE(N-H)]

= [(0.5 x 944) + (1.5 x 436)] – [3 x 388)]

= - 38 kJmol^{-1}

More complicated examples that may occur at A2

Working out ΔH_f of a compound using bond energies and other data
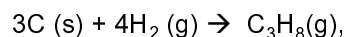
This is a more complicated example of the type in example 3

The ΔH's can be combinations of different data

elements → Compound in standard state

ΔH to turn to elements into gaseous atoms

ΔH to turn to compound into gaseous atoms

Gaseous atoms

Can be bond energies

E(Cl-Cl) Cl_2 → 2Cl

Or atomisation energies (if the substance is not diatomic

C(s) → C(g)

Σ bond energies of compound + (ΔH to turn to gas if compound is not gaseous)

Example 4
Calculate ΔH_f for propane, C_3H_8(g), given the following data.
C(s) → C(g) ΔH = 715 kJ mol^{-1}

Bond	C–C	C–H	H–H
kJ mol^{-1}	348	412	436

3C (s) + 4H_2 (g) → C_3H_8(g),

ΔHf = ΔH to turn to elements into gaseous atoms – ΔH to turn to compound into gaseous atoms

ΔH_f = (3xΔH_{at} [C] + 4 x E[H-H]) – (2 x E[C-C]+ 8 x E[C-H])

= (3x715 + 4 x 436) – (2 x 348+ 8 x 412)

=-103 kJ mol^{-1}

Free-energy change (ΔG) and entropy change (ΔS)

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

Energy / **Progress of Reaction**

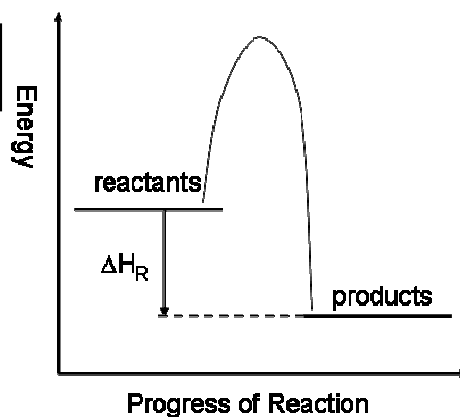reactants / ΔH_R / products

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?**

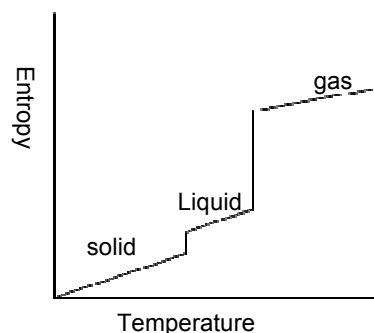We need to consider something called **entropy**

Entropy, $S°$

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

Entropy is a description of the number of ways atoms can share quanta of energy. If number of ways of arranging the energy (**W**) is <u>high</u>, then system is <u>disordered</u> and entropy (**S**) is <u>high</u>.

Elements		Compounds
Simpler compounds	*…tend to have lower entropies than…*	Complex compounds
Pure substances		Mixtures

Solids have lower entropies than liquids which are lower than gases. There is a bigger jump in entropy with boiling than that with melting. **Gases have large entropies**
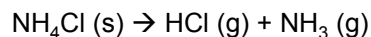

Entropy / **Temperature** / gas / Liquid / solid

Predicting Change in entropy 'ΔS' Qualitatively

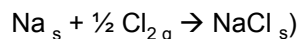An **increase** in disorder and **entropy** will lead to a positive entropy change $\Delta S° = +ve$

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

In general, a significant increase in the entropy will occur if:
- there is a *change of state* from **solid or liquid to gas**
- there is a significant *increase in number of molecules* between products and reactants.

NH_4Cl (s) → HCl (g) + NH_3 (g)
$\Delta S° = +ve$
•change from solid reactant to gaseous products
•increase in number of molecules
both will increase disorder

$Na_s + ½ Cl_{2g}$ → $NaCl_s$)
$\Delta S° = -ve$
•change from gaseous and solid reactant to solid
•decrease in number of molecules
both will decrease disorder

Calculating $\Delta S°$ 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.
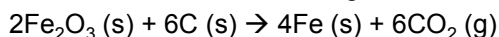
Elements in their standard states do *not* have zero entropy. **Only perfect crystals at absolute zero ($T = 0$ K) will have zero entropy:**

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

$$\Delta S° = \Sigma\ S°_{products} - \Sigma S°_{reactants}$$

Example

Calculate $\Delta S°$ for the following reaction at 25°C:
$2Fe_2O_3 (s) + 6C (s) \rightarrow 4Fe (s) + 6CO_2 (g)$

$\Delta S°$ = $S°_{products}$ - $S°_{reactants}$

= (6 x 213.6 + 4 x 27.3) – (2 x 87.4 + 6 x 5.7)

= + 1181.8 J K^{-1} mol^{-1} = **+ 1180 J K^{-1} mol^{-1}** (3 S.F.)

$S [Fe_2O_3]$ = 87.4 J K^{-1} mol^{-1}
$S [C]$ = 5.7 J K^{-1} mol^{-1}
$S [Fe]$ = 27.3 J K^{-1} mol^{-1}
$S [CO_2]$ = 213.6 J K^{-1} mol^{-1}

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

Gibbs Free Energy Change, ΔG

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, ΔG will be **negative**.

Gibbs free energy is a term that combines the effect of enthalpy and entropy into one number

A reaction that has increasing entropy (+ve ΔS) and is exothermic (-ve ΔH) will make ΔG be **negative** and will always be feasible

Convert from °C to **K** *(+ 273)*

Unit of S= **J K^{-1} mol^{-1}**

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

Need to convert to **KJ K^{-1} mol^{-1}** *(÷1000)*

Units: **KJ mol^{-1}**

Units: **KJ mol^{-1}**

Example : Data for the following reaction, which represents the reduction of aluminium oxide by carbon, are shown in the table.
$Al_2O_3(s) + 3C(s) \rightarrow 2Al(s) + 3CO(g)$
Calculate the values of ^-H , ^-S and ^-G for the above reaction at 298 K
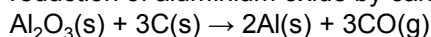
Substance	ΔHf / kJmol⁻¹	ΔS / JK⁻¹mol⁻¹
$Al_2O_3(s)$	-1669	51
C(s)	0	6
Al(s)	0	28
CO(g)	-111	198

1. Calculate ΔS

$\Delta S°$ = $S°_{products}$ - $S°_{reactants}$

= (2 x 28 + 3x198) – (51 + 3 x 6)

= +581J K^{-1} mol^{-1} (3 S.F.)

2. Calculate $\Delta H°$

$\Delta H°$ = $\Delta H°_f$ [products] - $\Delta H°_f$ [reactants]

= (3 x -111) – -1669

= +1336 kJ mol^{-1}

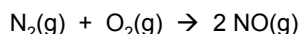**3. Calculate ΔG**

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

= +1336 – 298x 0.581

= +1163kJ mol^{-1}

ΔG 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

$N_2(g) + O_2(g) \rightarrow 2 NO(g)$

ΔH = 180 kJ mol^{-1} ΔS = 25 J K^{-1} mol^{-1}

The reaction will be feasible when ΔG ≤0

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

0 = ΔH - TΔS

So T= ΔH / ΔS

T = 180/ (25/1000)
= 7200K

The T must be >7200K

ΔG during phase changes

As physical phase changes like melting and boiling are equilibria, the ΔG for such changes is zero.

What temperature would methane melt at?

$CH_{4(s)} \rightarrow CH_{4 (l)}$ ΔH = 0.94 kJmol⁻¹ ΔS = 10.3 Jmol⁻¹K⁻¹

Make ΔG = 0 in the following equation $\Delta G = \Delta H - T\Delta S$
0 = ΔH - TΔS
So T= ΔH / ΔS

T= 0.94 / (10.3÷1000)

T= 91K