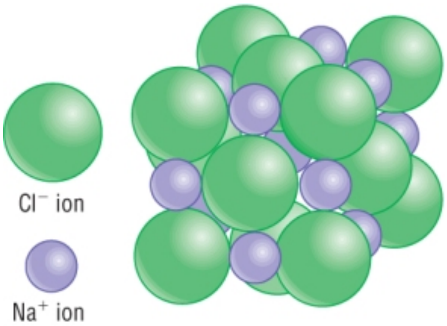
1.8 Thermodynamics

Review:

- In 1.3 we looked at ionic bonding and learned that:

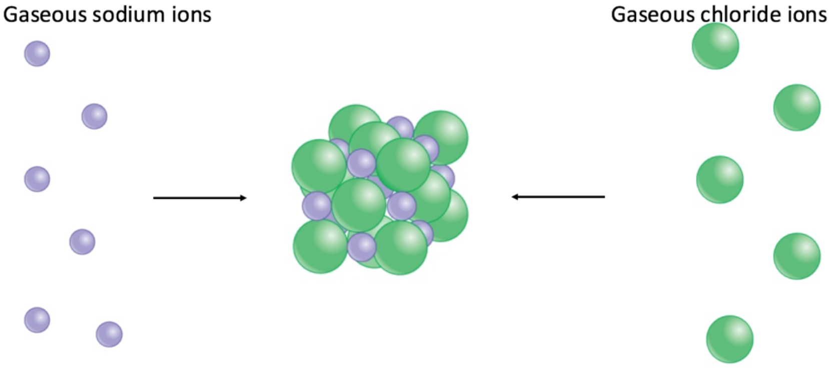
Giant ionic lattice structure



Ionic bonding: ***Strong electrostatic force of attraction between oppositely charged ions that are arranged in a regular repeating pattern***

Enthalpy of lattice formation, $\Delta_{LF}H_{lattice}$

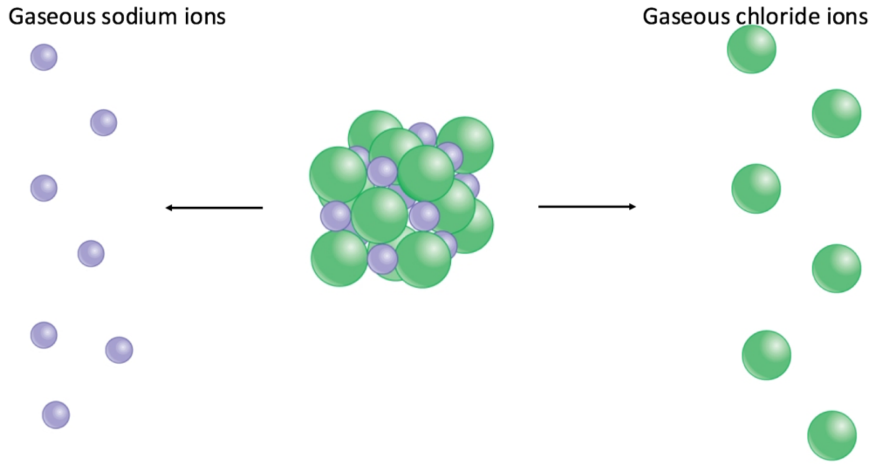
Lattice formation enthalpy ***Enthalpy change when 1 mole of a solid ionic compound is formed from its gaseous ions***



- This is an **exothermic process** as large amounts of energy are released upon the formation of the lattice from gaseous ions.

Enthalpy of lattice dissociation, $\Delta_{LD}H$

Lattice dissociation enthalpy
Enthalpy change when 1 mole of a solid ionic compound is completely dissociated into its gaseous ions



- This is an **endothermic process** as large amounts of energy are required upon the dissociation of the lattice to form gaseous ions.

Lattice enthalpies, $\Delta_{L}H$

- These enthalpies must be **equal and opposite** to each other as the amount of energy will be the **same**, the sign is due to the direction of the movement of energy:

Energy is required to break bonds – **endothermic** – dissociation

Energy is given out when bonds form – **exothermic** – formation

Enthalpy of lattice formation, $\Delta_{LF}H$ = - Enthalpy of lattice dissociation, $\Delta_{LD}H$

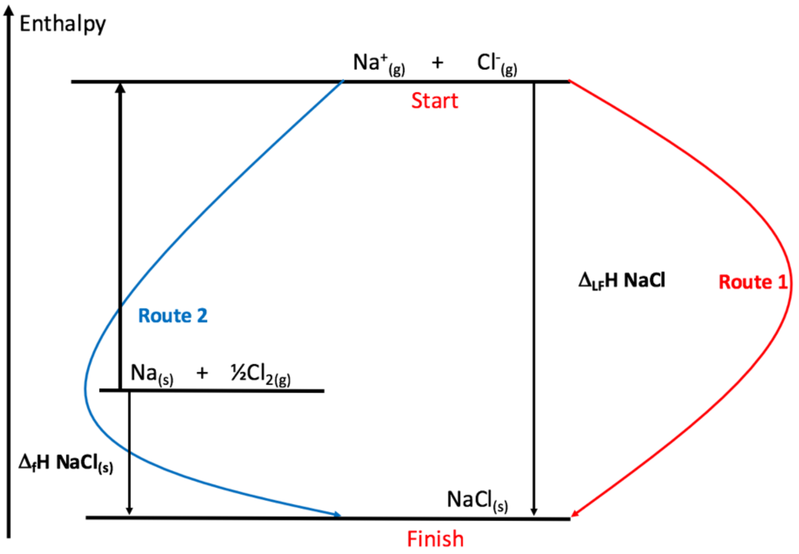
Enthalpy of lattice formation, $\Delta_{LF}H$ NaCl = -781 Kj mol^{-1}

Enthalpy of lattice dissociation, $\Delta_{LD}H$ NaCl = +781 Kj mol^{-1}

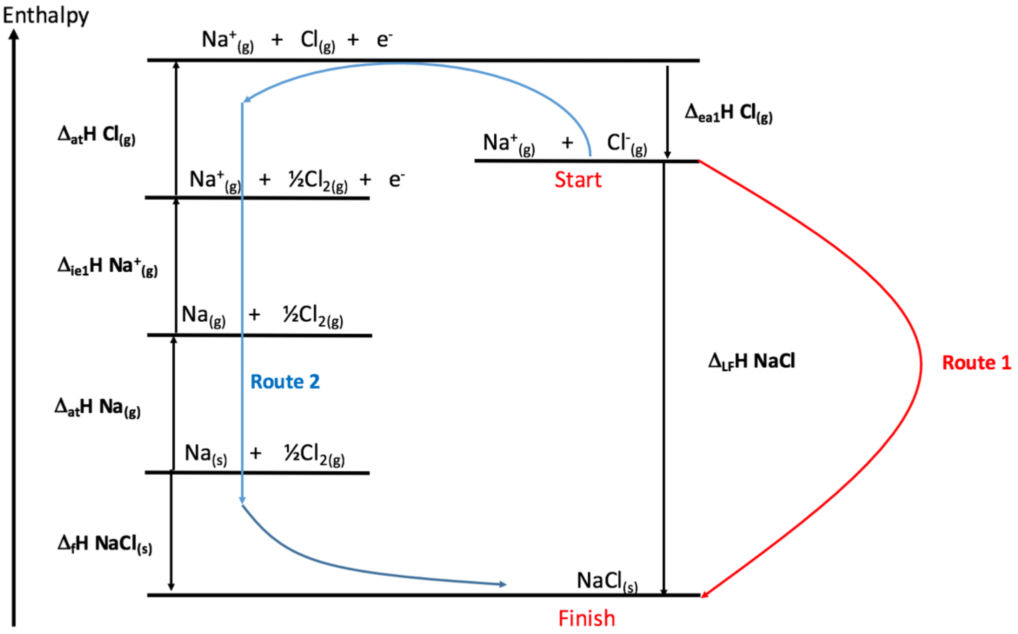
- Lattice enthalpy indicates the strength of ionic bonds.
- Lattice enthalpy is impossible to measure directly due to gaseous ions.
- A special type of **Hess's cycle** is used.
- This is called a **Born - Haber cycle**.
- These cycles require many more types of enthalpy changes:

Born - Haber cycles

- Start with the elements at **'zero energy'**
- **Endothermic processes go up**
- **Exothermic processes go down**
- A simple view of a **Born - Haber** cycle is shown below:



- **Route 2** is actually a multi - step process changing elements to gases then to ions



- **Route 2** has to be calculated in stages using a combination of enthalpy changes.
- Apply Hess's cycle to calculate the **Lattice energy**.

Route 1 = Route 2

$$\Delta_{LF}H \quad = \quad \Delta_fH \quad - \quad (\Sigma\Delta H\text{'s})$$

Top tip: Remember, $\Delta_{LF}H = -\Delta_{LD}H$ **So probably better to remember:**

$$\Delta_{LF}H = -\Delta_{LD}H \quad = \quad \Delta_fH \quad - \quad (\Sigma\Delta H\text{'s})$$

The other enthalpy changes:

1) Standard enthalpy change of formation, $\Delta_f H$:

$$\boxed{\text{Elements}} \xrightarrow{\Delta_f H} \boxed{\text{1 mole compound}}$$

Standard enthalpy change of formation, $\Delta_f H$:
1 mole of compound is formed from its constituent elements in their standard state.

$$K_{(s)} + {}^{1}/_{2}Cl_{2(g)} \rightarrow KCl_{(s)}$$

2) The standard enthalpy change of atomisation, $\Delta_{at}H$:

$$\boxed{\text{Elements}} \xrightarrow{\Delta_{at} H} \boxed{\text{1 mole gaseous atoms}}$$

Standard enthalpy change of atomisation, $\Delta_{at}H$:
1 mole of gaseous atoms are formed from its element in its standard state.

$$Mg_{(s)} \rightarrow Mg_{(g)}$$

3) First ionisation energy, $\Delta_{ie1}H$:

$$\boxed{\text{1 mole gaseous atoms}} \xrightarrow{\Delta_{ie1} H} \boxed{\text{1 mole gaseous 1+ ions}}$$

First ionisation energy, $\Delta_{ie1}H$:
1 mole of gaseous 1+ ions is formed from 1 mole of gaseous atoms

$$Li_{(g)} \rightarrow Li^{+}_{(g)} + e^{-}$$

4) Second ionisation energy, $\Delta_{ie2}H$:

$$\boxed{\text{1 mole gaseous 1+ ions}} \xrightarrow{\Delta_{ie2} H} \boxed{\text{1 mole gaseous 2+ ions}}$$

Second ionisation energy, $\Delta_{ie2}H$:
1 mole of gaseous 2+ ions is formed from 1 mole gaseous 1+ ions

$$Ca^{1+}_{(g)} \rightarrow Ca^{2+}_{(g)} + e^{-}$$

5) First Electron affinity, $\Delta_{ea1}H$:

$$\boxed{\text{1 mole gaseous atoms}} \xrightarrow{\Delta_{ea1} H} \boxed{\text{1 mole gaseous 1- ions}}$$

First Electron affinity, $\Delta_{ea1}H$:
1 mole of gaseous 1- ions formed from 1 mole of gaseous atoms

$$Cl_{(g)} + e^{-} \rightarrow Cl^{-}_{(g)}$$

6) Second Electron affinity, $\Delta_{ea2}H$:

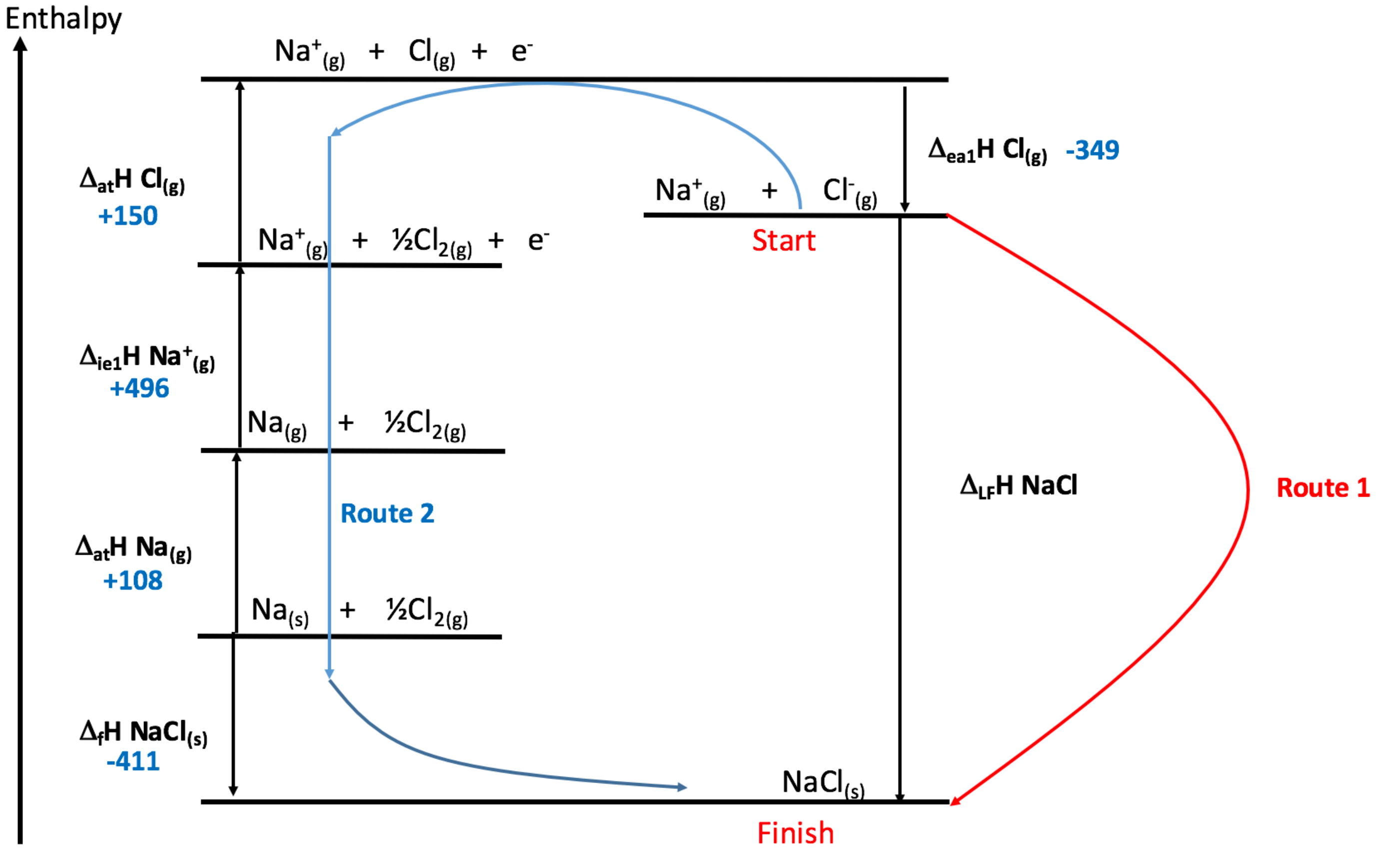
$$\boxed{\text{1 mole gaseous 1- ions}} \xrightarrow{\Delta_{ea2} H} \boxed{\text{1 mole gaseous 2- ions}}$$

Second Electron affinity, $\Delta_{ea2}H$:
1 mole of gaseous 2- ions formed from 1 mole gaseous 1- ions

$$O^{-}_{(g)} + e^{-} \rightarrow O^{2-}_{(g)}$$

Born - Haber cycle calculation

- These tend to use lattice enthalpy changes of formation, **$\Delta_{LF}H$** but you must be able to use **$\Delta_{LF}H$** and **$\Delta_{LD}H$**



Route 1 = Route 2

$\Delta_{LF}H \quad = \quad \Delta_f H \quad - \quad (\Sigma\Delta H\text{'s})$

$\Delta_{LF}H \quad = \quad -411 \quad - \quad (108 + 496 + 150 + -349)$

$\Delta_{LF}H \quad = \quad -816 \text{ kJ mol}^{-1}$

Top tips:

1) Mg**Cl_2** – All enthalpies involving chlorine only need to be multiplied by 2. Do this in the table of data FIRST

2) **Mg**Cl_2 – Mg is in group 2, therefore 2 ionisation energies required here.

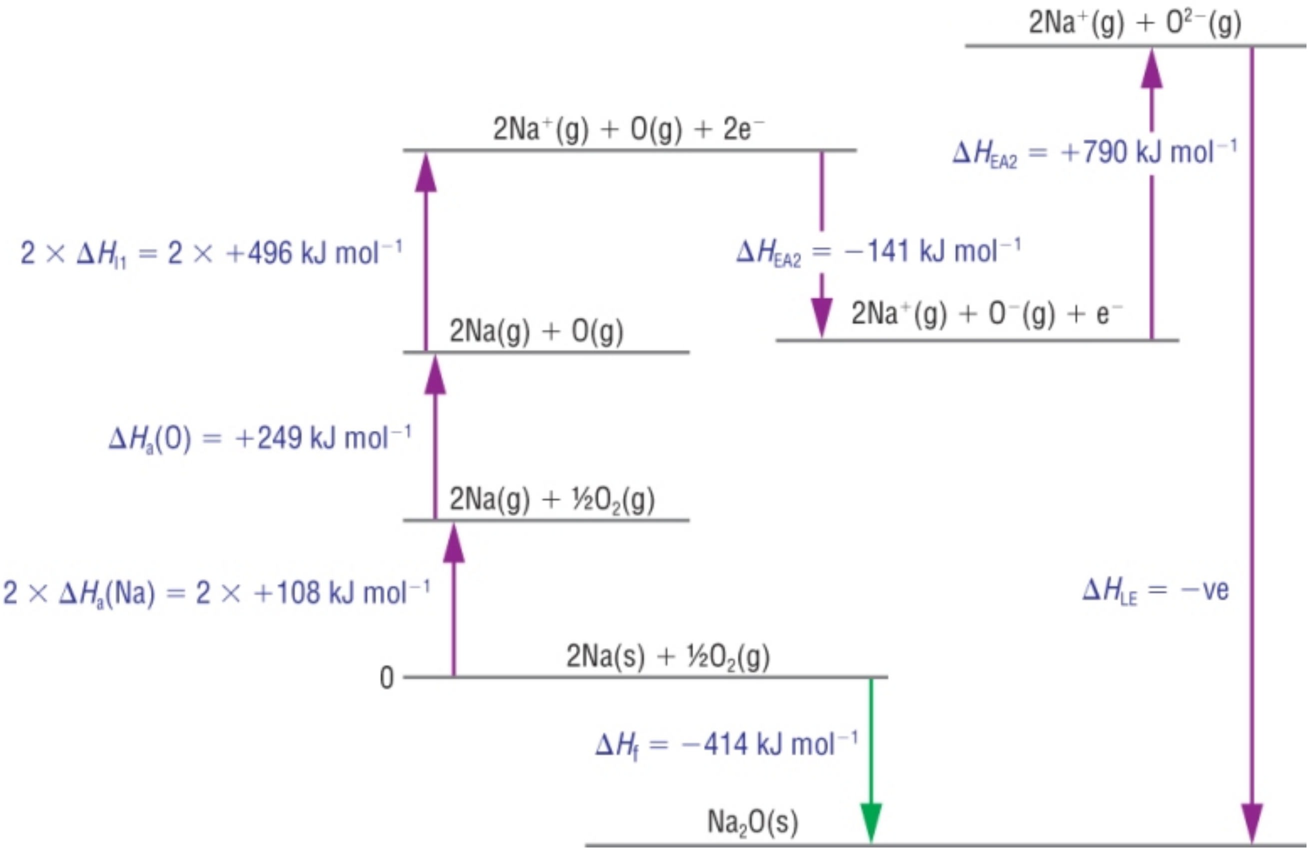
$$Mg \rightarrow Mg^+ \rightarrow Mg^{2+}$$

3) Mg**O** – O is in group 6, therefore 2 electron affinities required here.

$$O \rightarrow O^- \rightarrow O^{2-}$$

Other worked examples:

1) Sodium oxide:



Calculation:

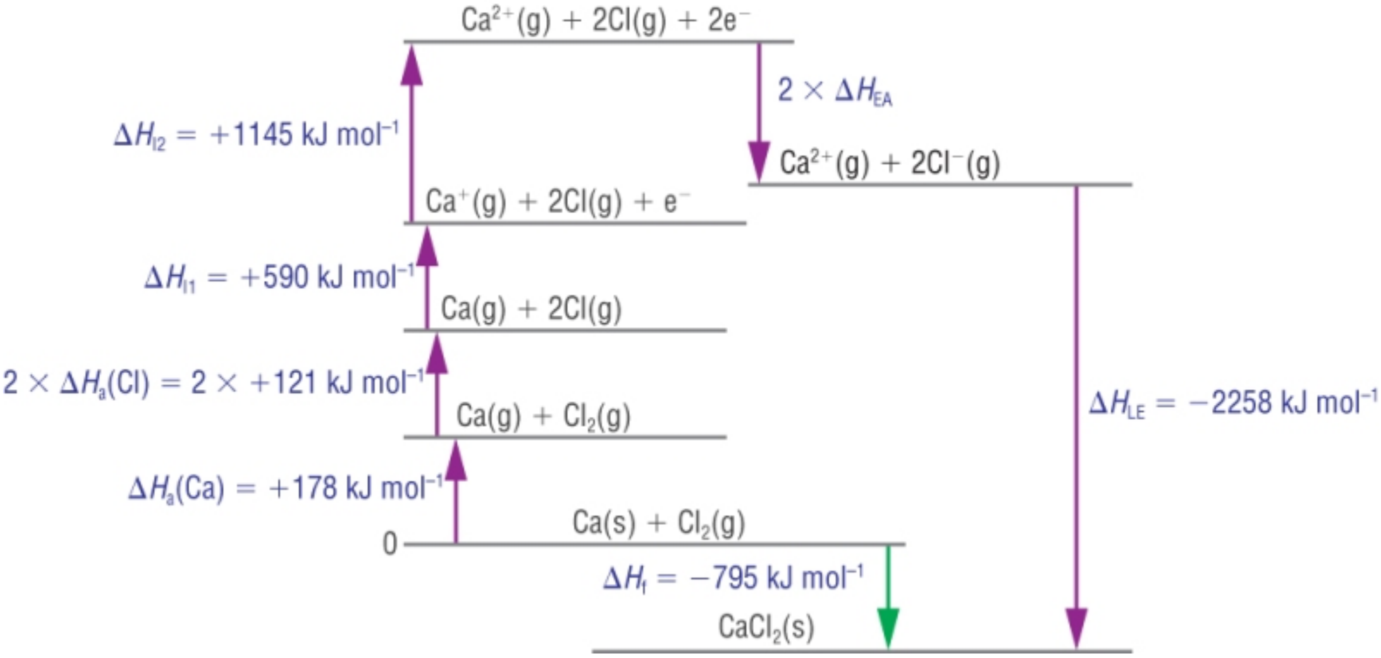
$\Delta_{LF}H \quad = \quad \Delta_f H \quad - \quad (\Sigma\Delta H\text{'s})$

$\Delta_{LF}H \quad = \quad -414 \quad - \quad (216 + 249 + 992 + -141 + 790)$

$\Delta_{LF}H \quad = \quad -1520\ kJ\ mol^{-1}$

2) Calcium chloride:

- These cycles can be used to calculate other unknowns:



Calculation:

$\Delta_{LF}H = \Delta_f H - (\Sigma\Delta H\text{'s})$

$-2258 = -795 - (178 + 242 + 590 + 1145 + 2\Delta_{ea1}H)$

$-2258 = -795 - (2155 + 2\Delta_{ea1}H)$

$-2258 = -795 - 2155 - 2\Delta_{ea1}H$

$-2258 = -2950 - 2\Delta_{ea1}H$

$2\Delta_{ea1}H - 2258 = -2950$

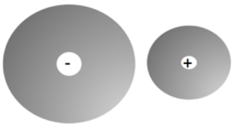
$2\Delta_{ea1}H = -2950 + 2258$

$2\Delta_{ea1}H = -692$

$2\Delta_{ea1}H = -346$ kJ mol^{-1}

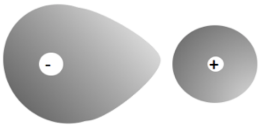
Theoretical vs experimental enthalpies:

- These calculations work on the assumption that the compound is **purely ionic**:
  - Ions are spherical
  - Charge is evenly distributed



Perfect ionic model:

- Spherical ions
- Charge evenly distributed



Partial covalent bonding:

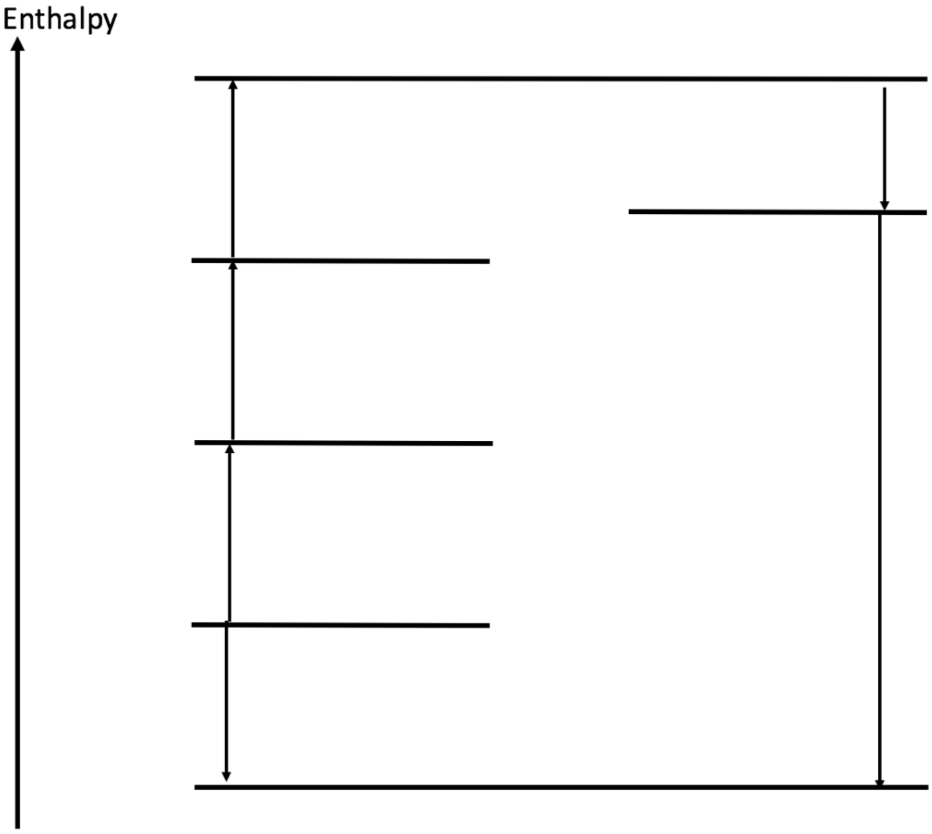
- Positive ion attracts electrons in the negative ions
- Ions therefore not spherical
- Charge therefore not evenly distributed

Questions:

For each of the questions below, complete the table, fill in the Born – Haber cycle and calculate the value with a ?:

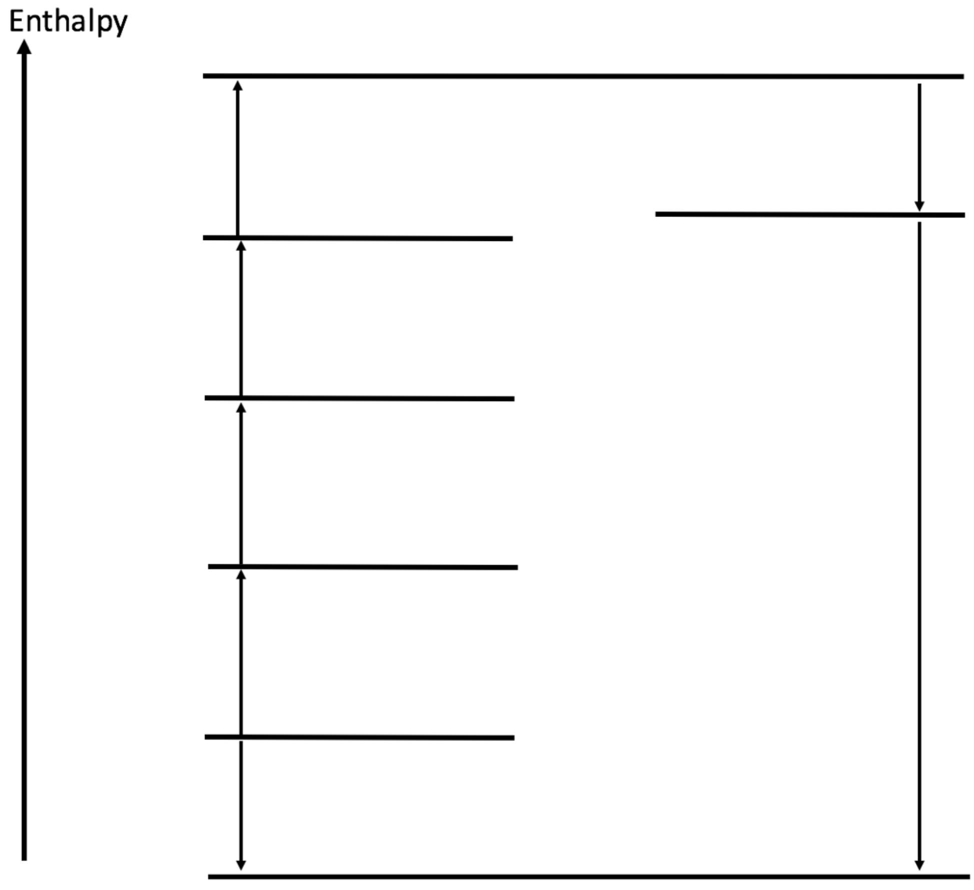
1. **sodium bromide**

Enthalpy change	**Equation**	**kJmol^{-1}**
Formation of NaBr		- 360
Atomisation of sodium		+109
Atomisation of bromine		+112
1st ionisation of sodium		+494
1st electron affinity of Br		-325
Lattice enthalpy of formation NaBr		?



2. magnesium chloride

Enthalpy change	**Equation**	**kJmol^{-1}**
Formation of $MgCl_2$		-641
Atomisation of Mg		+148
Atomisation of chlorine		+121
1st ionisation of Mg		+738
2nd ionisation of Mg		+1451
1st electron affinity of chlorine		-364
Lattice enthalpy of formation $MgCl_2$		?



3. Potassium oxide

Enthalpy change	Equation	kJmol^{-1}
Formation of K_2O		- 414
Atomisation of potassium		+109
Atomisation of oxygen		+249
1st ionisation of potassium		+494
1st electron affinity of O		-141
2nd electron affinity of O		+791
Lattice enthalpy of formation of K_2O		?

Calculate the electron affinity of fluorine

Enthalpy change	**Equation**	**kJmol^{-1}**
Formation of CaF_2		- 1220
Atomisation of calcium		+177
Atomisation of fluorine		+79
1st ionisation of calcium		+590
2nd ionisation of calcium		+1100
Electron affinity of fluorine		?
Lattice enthalpy of formation of CaF_2		-2630

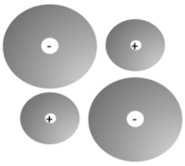
Enthalpy change of solution, $\Delta_{sol}H$

What happens when a solid dissolves?

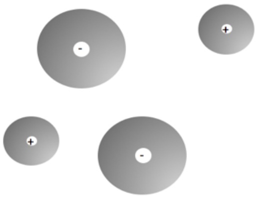
- If the ionic bonds are so strong, how is it possible to break this apart?
- The clue comes with the temperature changes that occur when they dissolve.
- Some are slightly exothermic while others are slightly endothermic.
- The energy produced when water surrounds the ions must be about the same as the electrostatic forces of attraction between the ions.
- In terms of a Hess's cycle, this is what happens when a solid dissolves:

1) Ionic lattice breaks down into gaseous ions, $\Delta_{LD}H$:

- Ionic compound forms gaseous ions, $\Delta_{LD}H$ **- Bond breaking – Endothermic – E_{in}**

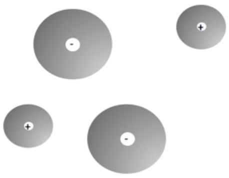


Breaking bonds
→

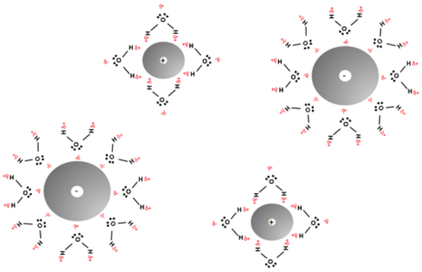


2) Hydration of gaseous ions, $\Delta_{hyd}H$:

- Gaseous ions form aqueous ions, $\Delta_{hyd}H$ **- Bond forming – Exothermic – E_{out}**



Forming bonds
→



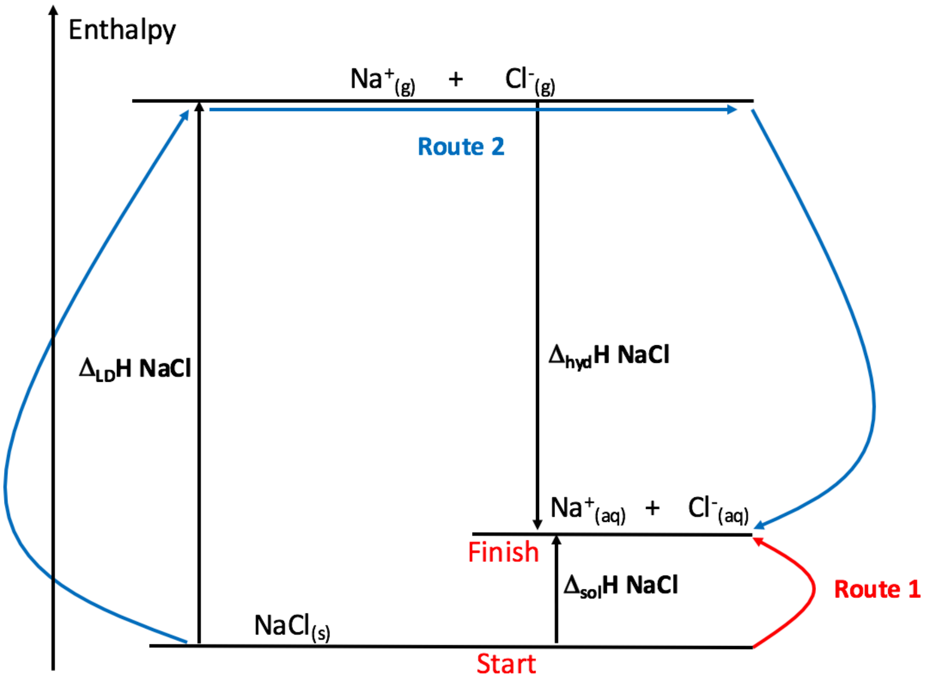
3) The resulting enthalpy change is known as the **Enthalpy change of solution, $\Delta_{sol}H$**:

- If there is energy left over, it is given to the surroundings - exothermic
- If there is not enough energy, it is taken in from the surroundings - endothermic
- These can all be calculated in another type of Hess's cycle:

Exothermic	**Endothermic**
$\Delta_{LD}H < \Delta_{hyd}H$	$\Delta_{LD}H > \Delta_{hyd}H$
Bond breaking < Bond forming	Bond breaking > Bond forming
$E_{in} < E_{out}$	$E_{in} > E_{out}$

Calculating enthalpy changes of solution, $\Delta_{sol}H$:

- These can all be calculated in another type of Hess's cycle:



- Some new enthalpy changes are required first:

1) Enthalpy change of solution, $\Delta_{sol}H$

1 mole compound $\xrightarrow{\Delta_{sol}H}$ Aqueous ions

Enthalpy change of solution, $\Delta_{sol}H$:
1 mole of an ionic compound completely dissolves in enough solvent that the ions no longer interact with each other

$$NaCl_{(s)} + aq \rightarrow Na^+_{(aq)} + Cl^-_{(aq)}$$

2) Enthalpy change of hydration, $\Delta_{hyd}H$

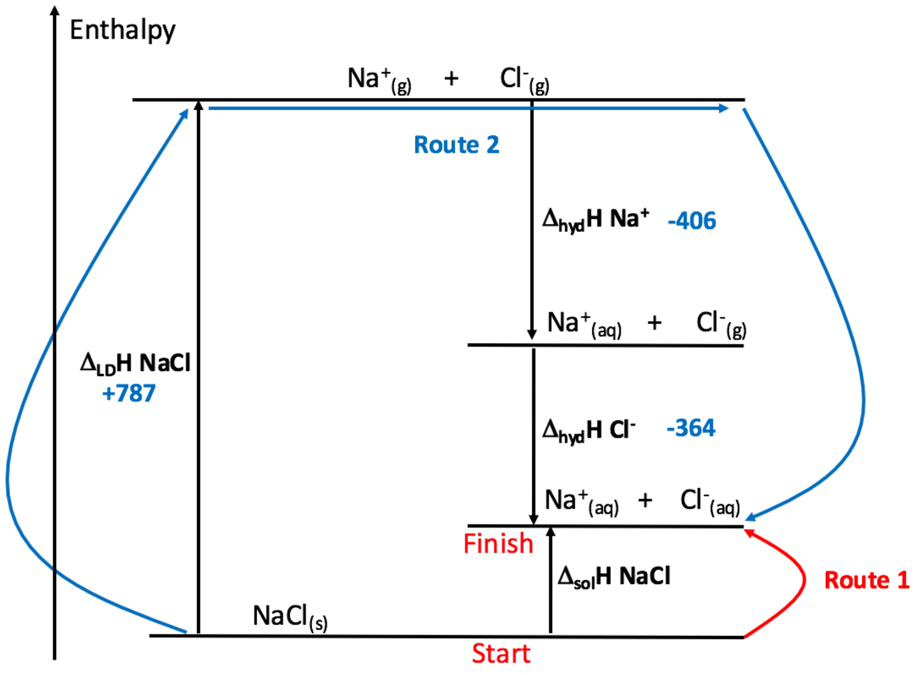
Gaseous ions $\xrightarrow{\Delta_{hyd}H}$ 1 mole aqueous ions

Enthalpy change of hydration, $\Delta_{hyd}H$:
1 mole of aqueous ions are formed from gaseous ions

$$Na^+_{(g)} + aq \rightarrow Na^+_{(aq)}$$

$$Cl^-_{(g)} + aq \rightarrow Cl^-_{(aq)}$$

- The hydration enthalpies on the cycle are therefore actually 2 changes:



Route 1 = Route 2

$\Delta_{sol}H \quad = \quad \Delta_{LD}H \quad + \quad (\Sigma\Delta_{hyd}H\text{'s})$

Top tip: Remember, $\Delta_{LF}H = - \Delta_{LD}H$ **So also remember:**

$\Delta_{sol}H \quad = \quad - \Delta_{LF}H \quad + \quad (\Sigma\Delta_{hyd}H\text{'s})$

$$\Delta_{sol}H = \Delta_{LD}H + (\Sigma\Delta_{\mathbf{hyd}}H\text{'s})$$

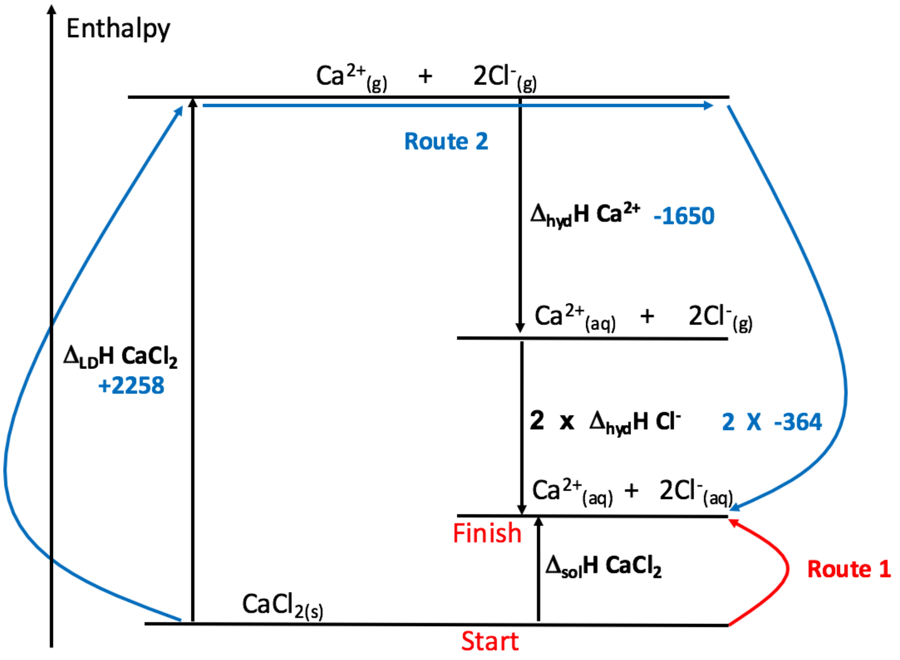
$$\Delta_{sol}H = 787 + (-406 + -364)$$

$$\Delta_{sol}H = 787 + (-770)$$

$$\Delta_{sol}H = 787 - 770$$

$$\Delta_{sol}H = 17 \text{ kJ mol}^{-1}$$

Another worked example:



$\Delta_{sol}H = \Delta_{LD}H + (\Sigma\Delta_{\mathbf{hyd}}H\text{'s})$

$\Delta_{sol}H = 2258 + (-1650 + (2 \times -364))$

$\Delta_{sol}H = 2258 + (-1650 + -728)$

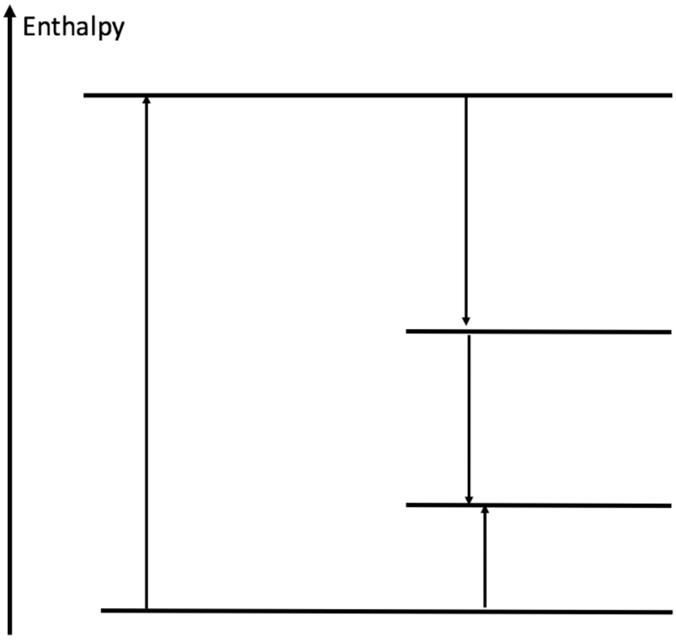
$\Delta_{sol}H = 2258 - 2378$

$\Delta_{sol}H = -120 \text{ kJ mol}^{-1}$

Questions:

1. **Calculate the enthalpy of solution of sodium chloride**

$\Delta_{LD}H$ [NaCl] +771 kJmol^{-1}; $\Delta_{hyd}H$ [Na^{+}] -406 kJmol^{-1} ; $\Delta_{hyd}H$ [Cl^{-}] -363 kJmol^{-1}



2. **Calculate the enthalpy of solution of calcium chloride**

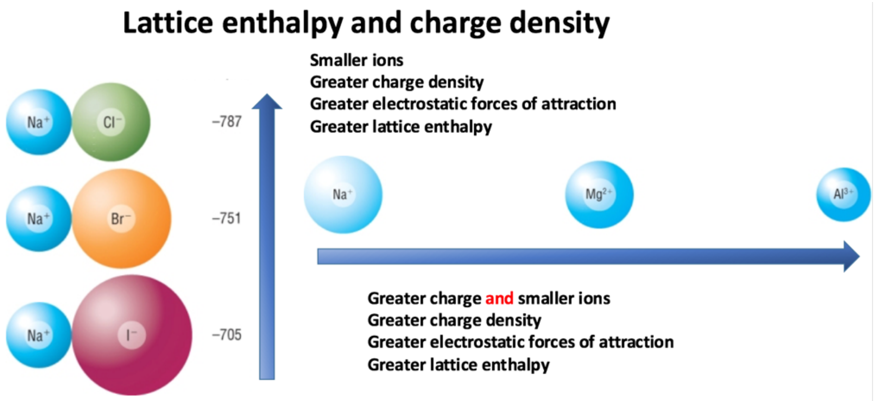
$\Delta_{LD}H$ [CaCl$_2$] +2258 kJmol^{-1}; $\Delta_{hyd}H$ [Ca^{2+}] -1577 kJmol^{-1} ; $\Delta_{hyd}H$ [Cl^{-}] -363 kJmol^{-1}

Factors affecting hydration and lattice enthalpies

1) Lattice enthalpies

- Lattice enthalpy is due to strong electrostatic forces of attraction between oppositely charged ions.
- The **charge density** affects how strong these forces of attractions are
- 2 things affect charge density:

  - **Size of ion**
  - **Size of charge**


Lattice enthalpy and charge density

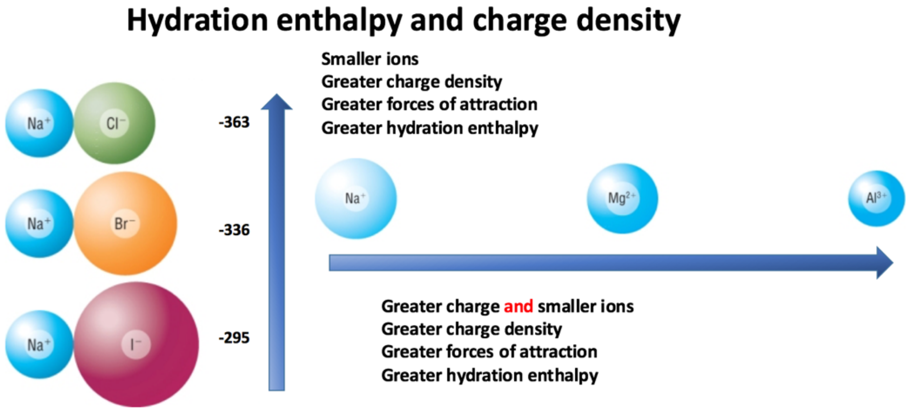
Smaller ions
Greater charge density
Greater electrostatic forces of attraction
Greater lattice enthalpy

Greater charge and smaller ions
Greater charge density
Greater electrostatic forces of attraction
Greater lattice enthalpy

2) Hydration enthalpies

- Hydration enthalpy is due to forces of attraction between ions and water.
- The **charge density** affects how strong these forces of attractions are
- 2 things affect charge density:

  - **Size of ion**
  - **Size of charge**


Hydration enthalpy and charge density

Smaller ions
Greater charge density
Greater forces of attraction
Greater hydration enthalpy

Greater charge and smaller ions
Greater charge density
Greater forces of attraction
Greater hydration enthalpy

<u>Entropy, S</u> – Feasibility of a reaction

What is entropy:

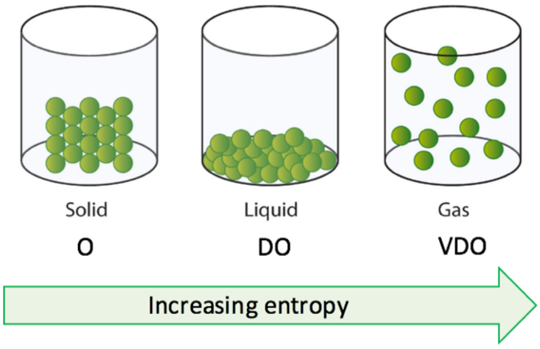
- Entropy is a measure of **disorder** in a system.
- It is a measure of the number of ways that energy can be arranged.
- There is always a degree of disorder as particles always have energy, values = **positive**.
- **Generally, all reactions (processes) move to disorder**

Entropy:

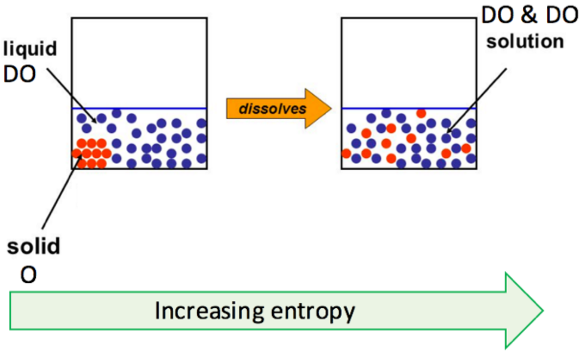
- Naturally occurring events that lead to disorder:

Melting ice cream	**Smell of cooking spreading**
Your bedroom	**Expanding Universe**

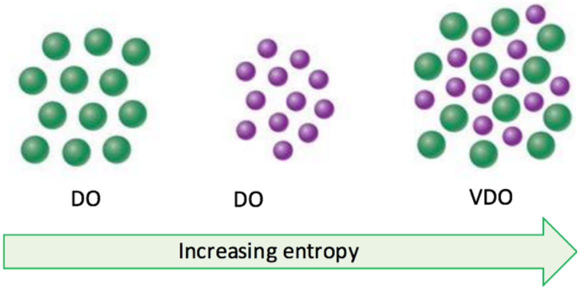
States of matter:



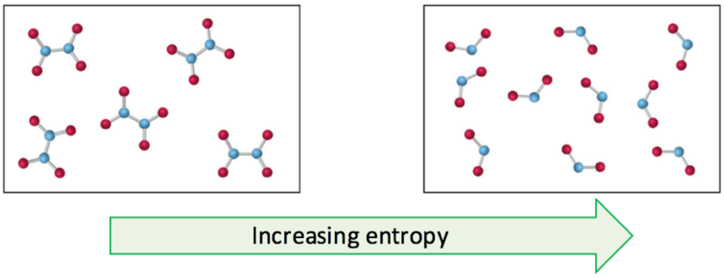
Dissolving:



Mixtures:



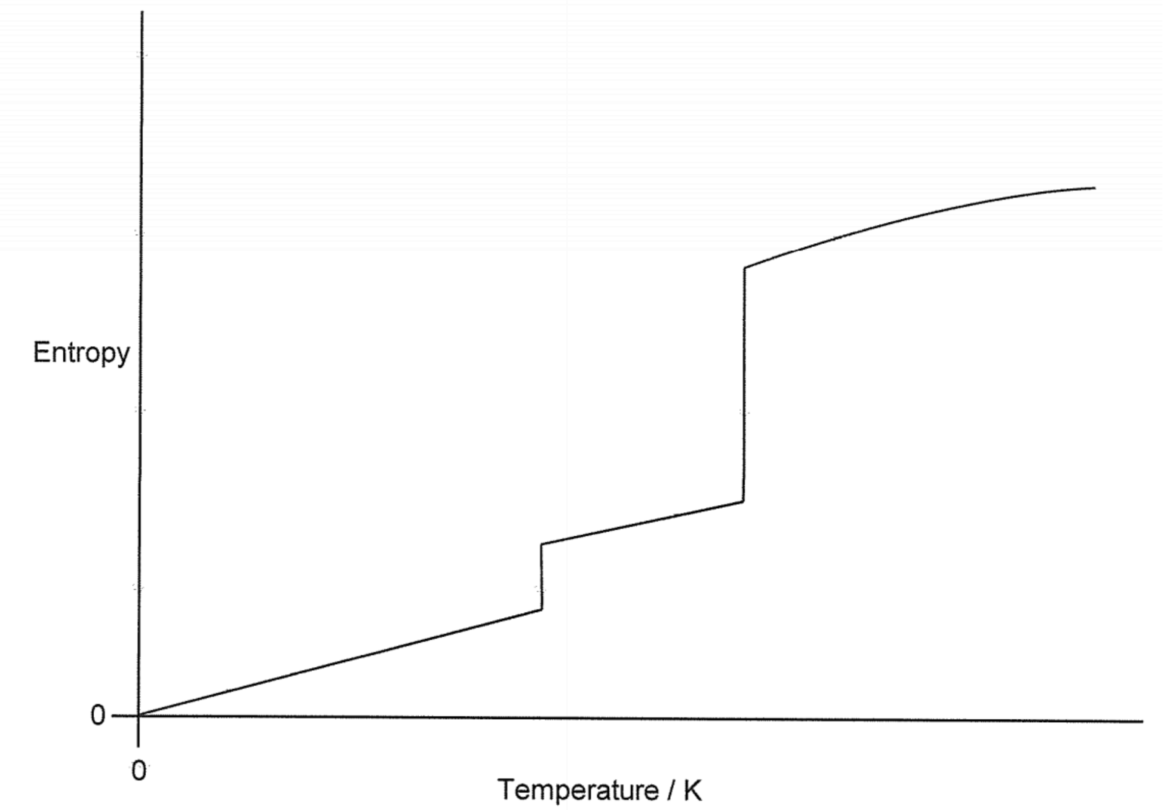
Number of particles:



- **Generally, all reactions (processes) move to disorder / an increase in entropy**

Effect of temperature on entropy:

- Mark on the chart:
  - Solid
  - Liquid
  - Gas
  - Melting
  - Boiling



Label and explain:
Solid:

Liquid:

Gas:

Melting:

Boiling:

What can entropy tell us?

- All **feasible / spontaneous** processes start off **ordered** and become **disordered**.
- The entropy starts of with a low value (ordered) and leads to a higher value (disordered)
- This means that reactants and products will have an **entropy content, S**
- We are interested in the change in entropy of a chemical system.
- This gives us an idea as to whether a reaction is **feasible / spontaneous** or not?

Using Entropy:

Reaction A	$NH_4NO_{3(s)}$	+	aq	→	$NH_4^+{}_{(aq)}$	+	$NO_3^-{}_{(aq)}$
Entropy	O		DO	→	DO		DO
S Overall	O			→		DO	
				Feasible			

Reaction B	$MgCO_{3(s)}$	→	$MgO_{(s)}$	+	$CO_{2(g)}$
Entropy	O	→	O		VDO
S Overall	O	→		DO	
			Feasible		

Calculating entropy:

- Standard entropy content of a substance can be looked up in a data book.
- Units **J mol K^{-1}**
- The difference in entropy content between the products and reactants will give a value.

$$\boxed{\Delta S \quad = \quad \Sigma S_p \quad - \quad \Sigma S_r}$$

- **Positive** means that the system is moving to more **disorder - Feasible**
- **Negative** means that the system is moving towards more **order – Not feasible**

Worked example:

Reaction C	$2Mg_{(s)}$	+	$O_{2(g)}$	→	$2MgO_{(s)}$
Entropy content	+ 32.7		+ 205		+ 26.9

$$\Delta S = \Sigma S_p - \Sigma S_r$$

$$\Delta S = (2 \times +26.9) - [\,(+205) + (2 \times +32.7)\,]$$

$$\Delta S = -216.6 \text{ J K}^{-1} \text{ mol}^{-1}$$

Not feasible

Questions:

1) Which of the following reactions are likely to be feasible?

$C_6H_{12}O_{6(aq)} \rightarrow 6CO_{2(g)} + 6H_2O_{(l)}$

$2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$

$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$

$2NH_{3(g)} \rightarrow N_{2(g)} + 3H_{2(g)}$

$Ag^+_{(aq)} + Cl^-_{(aq)} \rightarrow AgCl_{(s}$

$Cl_{2(aq)} + 2KI_{(aq)} \rightarrow I_{2(aq)} + 2KCl_{(aq)}$

2) Calculate ΔS for the following reactions:

a) $RbCl_{(s)} + aq \rightarrow Rb^+_{(aq)} + Cl^-_{(aq)}$

S for $RbCl_{(s)}$ = + 95.9 J mol^{-1} K^{-1}

S for $Rb^+_{(aq)}$ = + 121.5 J mol^{-1} K^{-1}

S for $Cl^-_{(aq)}$ = + 56.5 J mol^{-1} K^{-1}

b) $C_2H_4(g) + H_2O(g) \rightarrow C_2H_5OH(g)$

	S in J K^{-1} mol^{-1}
Ethene (gas)	+ 226
Steam (gas)	+ 195
Ethanol (gas)	+ 276

A feasibility problem

- Look again at reactions B and C:

Reaction B: **Endothermic reaction**

	$MgCO_{3(s)}$	→	$MgO_{(s)}$	+	$CO_{2(g)}$
Entropy	O	→	O		VDO
S Overall	O	→		DO	

ΔS = Positive

Feasible

Not actually feasible (unless at high temperatures)

Reaction C: **Exothermic reaction**

	$2Mg_{(s)}$	+	$O_{2(g)}$	→	$2MgO_{(s)}$
Entropy content	+ 32.7		+ 205		+ 26.9

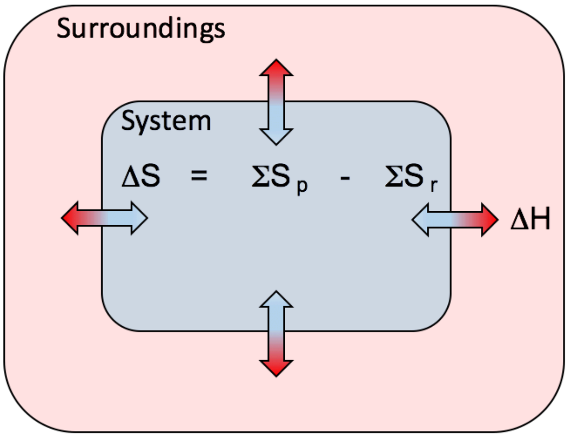
ΔS = - 216.6 J K^{-1} mol^{-1}

Not feasible

Actually feasible

The effect of enthalpy change, ΔH

- There must be more to what makes a reaction feasible.
- ΔS deals with the system.
- ΔH also plays a part and deals with the entropy of the surroundings:



- An exothermic reaction gives energy to the surroundings.
- The entropy of the surroundings therefore increases.

- An endothermic reaction takes energy from the surroundings.
- The entropy of the surroundings therefore decreases.

Enthalpy	**Deals with the entropy of the surroundings**
Entropy	**Deals with the entropy of the system**

Entropy of the system and temperature:

- The entropy contribution depends on temperature, **T** (K) at which the reaction takes place.

$T\Delta S$

- As temperature increases, the entropy becomes more significant (as there is more energy).

Gibbs free energy and feasibility:

- 3 factors contribute to whether a reaction will be feasible:

1 Temperature in k	**T**	K
2 Entropy change of the system	**ΔS**	$J\ K^{-1}\ mol^{-1}$
3 Enthalpy change with the surroundings	**ΔH**	$kJ\ mol^{-1}$

- The relationship between these 3 factors is expressed by Gibbs **Free energy, ΔG:**

$$\boxed{\Delta G \quad = \quad \Delta H \quad - \quad T\Delta S}$$

Surroundings System

- For there to be a **decrease in (free) energy**, leading to **more stability**:

$$\Delta G \quad < \quad 0$$

- ΔG must be equal or less than 0

The effect of temperature

ΔH	-	T ΔS	=	ΔG		Feasibility
(-)ve		**(+)ve**		**Always (-)ve**		**Always feasible**
(+)ve		**(-)ve**		**Always (+)ve**		**Never feasible**
(-)ve		**(-)ve**	As T decreases, TΔS becomes less (-)ve.	**(-)ve at low T**	so ΔG becomes more (-)ve.	**Only feasible at low T (C)**
(+)ve		**(+)ve**	As T increases, TΔS becomes more (+)ve.	**(-)ve at High T**	so ΔG becomes more (-)ve.	**Only feasible at high T (B)**

- Most exothermic reactions are feasible / spontaneous as enthalpy contributes more to **ΔG**.

How do endothermic reactions take place:

- For an endothermic reaction to take place, **ΔS must be (+)ve – see table above**
- The temperature must be high enough to make, **TΔS > ΔH**
- This will make **ΔG (-)ve** and therefore the reaction **feasible**

ΔH	-	T ΔS	=	ΔG		Feasibility
(+)ve		**(+)ve**	As T increases, TΔS becomes more (+)ve.	**(-)ve at High T**	so ΔG becomes more (-)ve.	**Only feasible at high T**

Calculations:

At what temperature will $ZnCO_3$ decompose?

Reaction	$ZnCO_{3(s)}$	$\rightarrow$	$ZnO_{(s)}$	+	$CO_{2(g)}$	$\Delta H = +71$ KJ mol^{-1}
Entropy, ΔS	+82		+44		+214	

1) Calculate ΔS:

$\Delta S = \Sigma S_p - \Sigma S_r$

$\Delta S = (+44 + +214) - +82$

$\Delta S = 176\ J\ K^{-1}\ mol^{-1}$

$\Delta S = 0.176\ kJ\ K^{-1}\ mol^{-1}$ **COMMON ERROR - MUST COVERT TO kJ /1000**

The units of ΔG and ΔH are kJ mol^{-1}, so ΔS must be converted to kJ $K^{-1} mol^{-1}$

The temperature is in Kelvin (K): *temperature in K = temperature in °C + 273*

2) Assume $\Delta G = 0$ to calculate T:

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

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

$T\Delta S = \Delta H$

$$T = \frac{\Delta H}{\Delta S}$$

$$T = \frac{71}{0.176}$$

T = 403 K **(-273 to convert to °C)**

T = 130 °C

Expect to calculate any expression using either or both of the expressions

Free energy graphs

- These are a plot of ΔG vs T (K).
- As the plot is always a straight line we can use the 'equation of a straight line'.

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

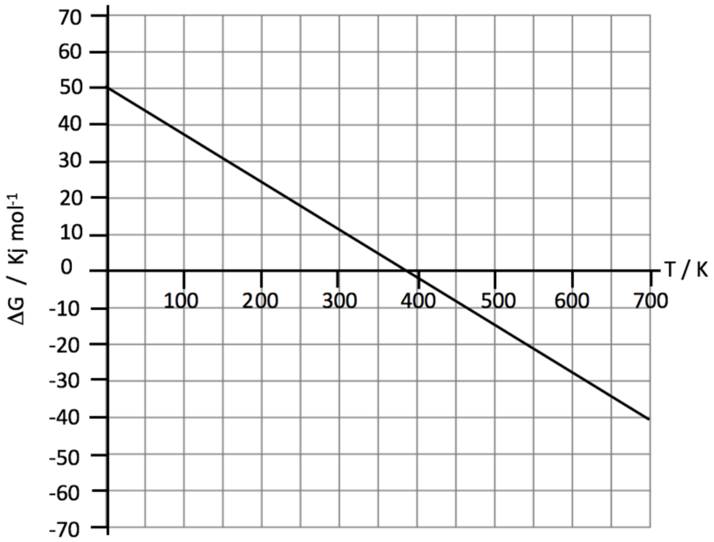
$$\Delta G \quad = \quad -T\Delta S \quad + \quad \Delta H$$

$$\boldsymbol{\Delta G \quad = \quad -\Delta S T \quad + \quad \Delta H}$$

$$\boldsymbol{y \quad = \quad mx \quad + \quad c}$$

- This allows us to calculate ΔH or ΔS.

A free energy graph:



y axis: y = ΔG

Gradient: m = $-\Delta S$

x axis: x = T

Intercept: c = ΔH

Intercept: c = ΔH = 50 Kj mol^{-1}

Gradient: m = $-\Delta S$ = $\frac{\Delta Y}{\Delta x}$

$$-\Delta S = \frac{(-40 \quad - \quad 50)}{(700 \quad - \quad 0)}$$

$$-\Delta S = \frac{-90}{700}$$

$-\Delta S$ = - 0.12857 kJ mol^{-1} *Convert to J: x1000

ΔS = 128.57 J mol^{-1}

*Note you can see from the graph that this experiment only becomes feasible above a certain temperature. ΔG becomes (-)ve above 390 K (ish).

Questions:

1) The decomposition of $ZnCO_3$ only happens at high temperatures.
   a. Write a balanced chemical equation for the decomposition of $ZnCO_3$

   b. Use the data in the table to calculate the enthalpy change for this decomposition reaction:

Substance	**$\Delta_f H$ / kJ mol^{-1}**
$ZnCO_3$	-813
ZnO	-348
CO_2	-394

2) The reaction of sodium hydrogencarbonate and hydrochloric acid has the following standard data:

Substance	**$\Delta_f H$ / kJ mol^{-1}**	**S / J K^{-1}mol^{-1}**
$NaHCO_3$ (s)	-951	102
HCl (aq)	-167	57.0
NaCl(aq)	-407	72.4
H_2O (l)	-286	70.0
CO_2 (g)	-394	213

   a. Write a balanced chemical equation for this reaction

   b. Use the data in the table to calculate the enthalpy change:

   c. Use the data in the table to calculate the entropy change:

   d. Use the data in the table to calculate Gibbs free energy and state whether the reaction in feasible under standard conditions:

3) The rusting of iron has the following data:

Substance	**$\Delta_f H$ / kJ mol^{-1}**	**S / J K^{-1}mol^{-1}**
Fe (s)	0	27.3
O_2 (g)	0	205
Fe_2O_3 (s)	-825	87.4

a. Write a balanced chemical equation for the rusting of iron

b. Use the data in the table to calculate the enthalpy change:

c. Use the data in the table to calculate the entropy change:

d. Use the data in the table to calculate Gibbs free energy and state whether the reaction in feasible under standard conditions:

e. Explain why a reaction with a negative (unfavourable) entropy change can be feasible at room temperature.

4) The decomposition of $MgCO_3$ only happens at high temperatures.
   a. Write a balanced chemical equation for the decomposition of $MgCO_3$

   b. Use the data in the table to calculate the enthalpy change for this decomposition reaction:

Substance	**$\Delta_f H$ / kJ mol^{-1}**	**S / J K^{-1}mol^{-1}**
$MgCO_3$	-1095.8	65.7
MgO	-601.6	27.0
CO_2	-393.5	213.8

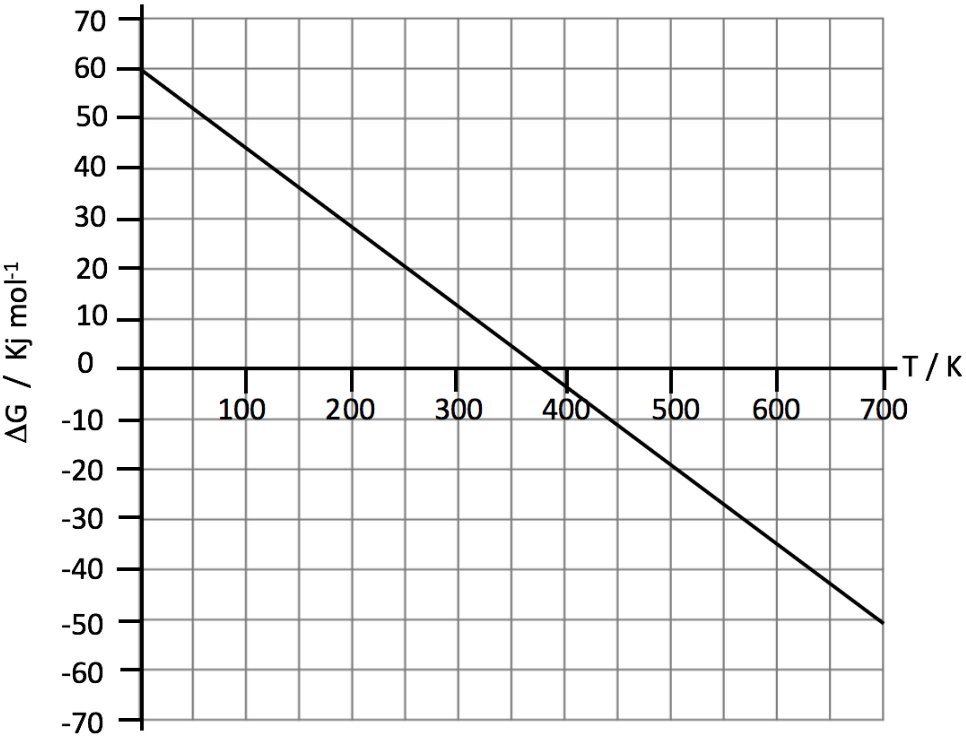
   c. Use the data in the table to calculate the entropy change:

   d. Use the data in the table to calculate Gibbs free energy and state whether the reaction is feasible at room temperature:

   e. Explain why this reaction is only feasible at high temperatures?

   f. Calculate the minimum temperature that this reaction will be feasible

5) The free energy graph below shows how free energy changes with temperature for a gas phase reaction:



a. Explain how this thermodynamic relationship obeys the maths straight line equation, $y = mx + c$

b. Use the graph to determine ΔH for this reaction:

c. Use the graph to determine ΔS for this reaction:

d. Explain why ΔG decreases as T increases:

e. What can you deduce about the feasibility of the reaction at temperatures below 350K?