

PROSPERITY ACADEMY

A2 CHEMISTRY 9701

Crash Course

RUHAB IQBAL

**CHEMICAL
ENERGETICS**

COMPLETE NOTES



0331 - 2863334

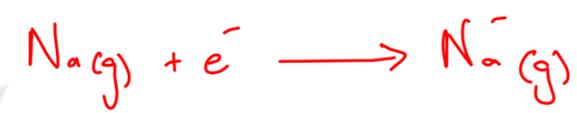


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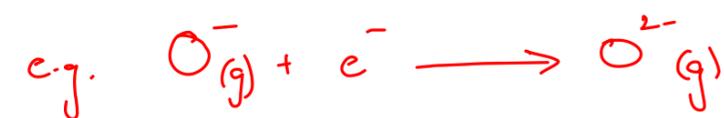


Electron affinities (EA):-

1st electron affinity:- Amount of energy released or absorbed when 1 mole of gaseous atoms gain 1 mole of electrons to become 1 mole of -1 charged gaseous anions



2nd electron affinity:- Amount of energy absorbed when 1 mole of gaseous -1 charged anions gain 1 mole of electrons to become 1 mole of -2 charged gaseous anions



Why are all electron affinities after 1st electron affinity positive (i.e. absorb energy/endothermic)?

Ans. Energy is needed to overcome the repulsion between the 2 negatively charged species.

Trends in electron affinities (Only discussing Group 16 and 17):-

| First electron affinity / kJ mol ⁻¹ | |
|--|----------|
| Group 16 | Group 17 |
| O: -141 | F: -328 |
| S: -200 | Cl: -345 |
| Se: -195 | Br: -325 |
| Te: -190 | I: -295 |

Across period, gets more exothermic

Down the group, gets less exothermic

less exothermic than expected as their radii are too small so already there is too much repulsion due to high electron density.

• Across period, 1st electron affinity gets more exothermic:-

- Nuclear charge increases
- Shielding remains constant
- Atomic radius decreases hence attraction between nucleus and valence shell increases
- more energy released when electron gained (stronger bond between nucleus and electron)

• Across group, 1st electron affinity gets less exothermic

- Nuclear charge and shielding both increase and counteract each other
- Atomic radius increases due to more number of shells so attraction between nucleus and valence shell decreases
- less energy released when electron gained (weak bond between nucleus and electron).

Born Haber cycles:- This is used to find out lattice energy using Hess' law

↳ The enthalpy change of reaction is independent of the path taken provided all measurements are made under the same conditions.

Q. Find the lattice energy of NaCl, given the following:-

$$\Delta H_f^\ominus \text{NaCl} = -411 \text{ kJ mol}^{-1}$$

$$\Delta H_{\text{at}}^\ominus \text{ of Na} = +107 \text{ kJ mol}^{-1}$$

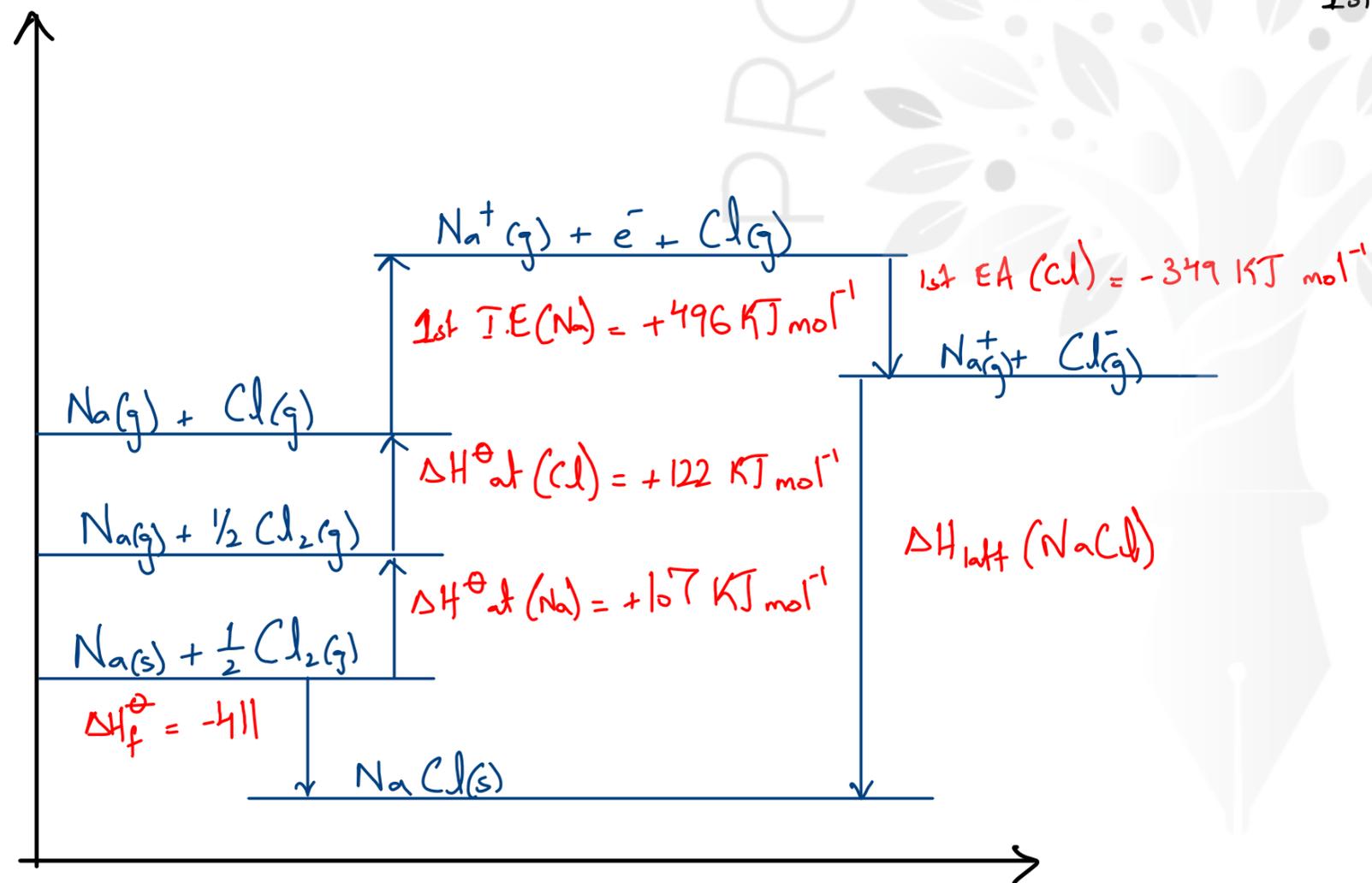
$$\Delta H_{\text{at}}^\ominus \text{ of Cl} = +122 \text{ kJ mol}^{-1}$$

$$\text{1st I.E. of Na} = +496 \text{ kJ mol}^{-1}$$

$$\text{1st E.A. of Cl} = -349 \text{ kJ mol}^{-1}$$



energy / kJ mol^{-1}



$$-411 = +107 + 122 + 496 - 349 + \Delta H_{\text{latt}}$$

$$\Delta H_{\text{latt}} = -787 \text{ kJ mol}^{-1}$$

2 Most car air bags contain a capsule of sodium azide, NaN_3 . In a crash, the NaN_3 decomposes into its elements.

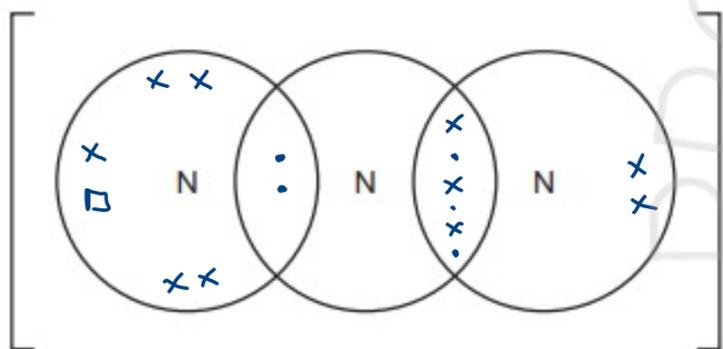
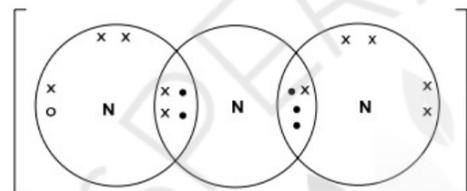
(a) Write an equation for the decomposition of NaN_3 .



(b) Complete the 'dot-and-cross' diagram for the azide ion, N_3^- .

Use the following key for the electrons.

- electrons from central nitrogen atom
- x electrons from the other two nitrogen atoms
- added electron(s) responsible for the overall negative charge



[3]

(c) Lattice energies are always negative showing that they represent exothermic changes.

(i) Explain what is meant by the term *lattice energy*.

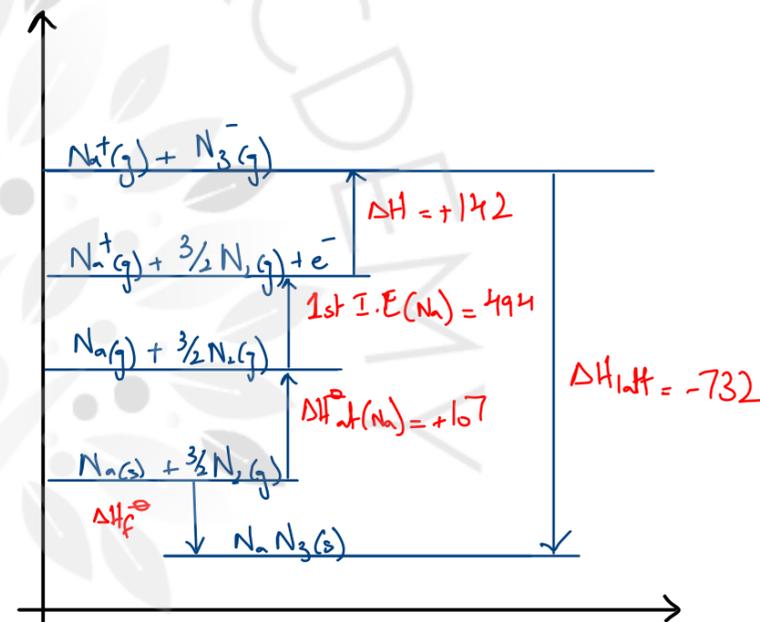
Amount of energy released when 1 mole of an ionic compound is formed from its gaseous ions under standard conditions [2]

(ii) Explain why lattice energy represents an exothermic change.

A strong ionic bond is made. Bond formation is always exothermic. [1]

(iii) Use the following data and any relevant data from the *Data Booklet* to calculate the standard enthalpy change of formation, ΔH_f^\ominus , of $\text{NaN}_3(\text{s})$. $\frac{1}{2} \text{N} \equiv \text{N} \rightarrow \text{N}(\text{g})$
Include a sign in your answer. Show all your working.

| | |
|---|----------------------------|
| lattice energy, $\Delta H_{\text{latt}}^\ominus$, of $\text{NaN}_3(\text{s})$ | -732 kJ mol^{-1} |
| standard enthalpy change of atomisation, $\Delta H_{\text{at}}^\ominus$, of $\text{Na}(\text{g})$ | $+107 \text{ kJ mol}^{-1}$ |
| standard enthalpy change, ΔH^\ominus , for $\frac{1}{2} \text{N}_2(\text{g}) + \text{e}^- \rightarrow \text{N}_3^-(\text{g})$ | $+142 \text{ kJ mol}^{-1}$ |



$$\Delta H_f^\ominus = 107 + 494 + 142 - 732 = 11 \text{ kJ mol}^{-1}$$

ΔH_f^\ominus of $\text{NaN}_3(\text{s}) = \dots + 11 \text{ kJ mol}^{-1}$ [3]

(iv) The lattice energy, $\Delta H_{\text{latt}}^\ominus$, of $\text{RbN}_3(\text{s})$ is -636 kJ mol^{-1} .

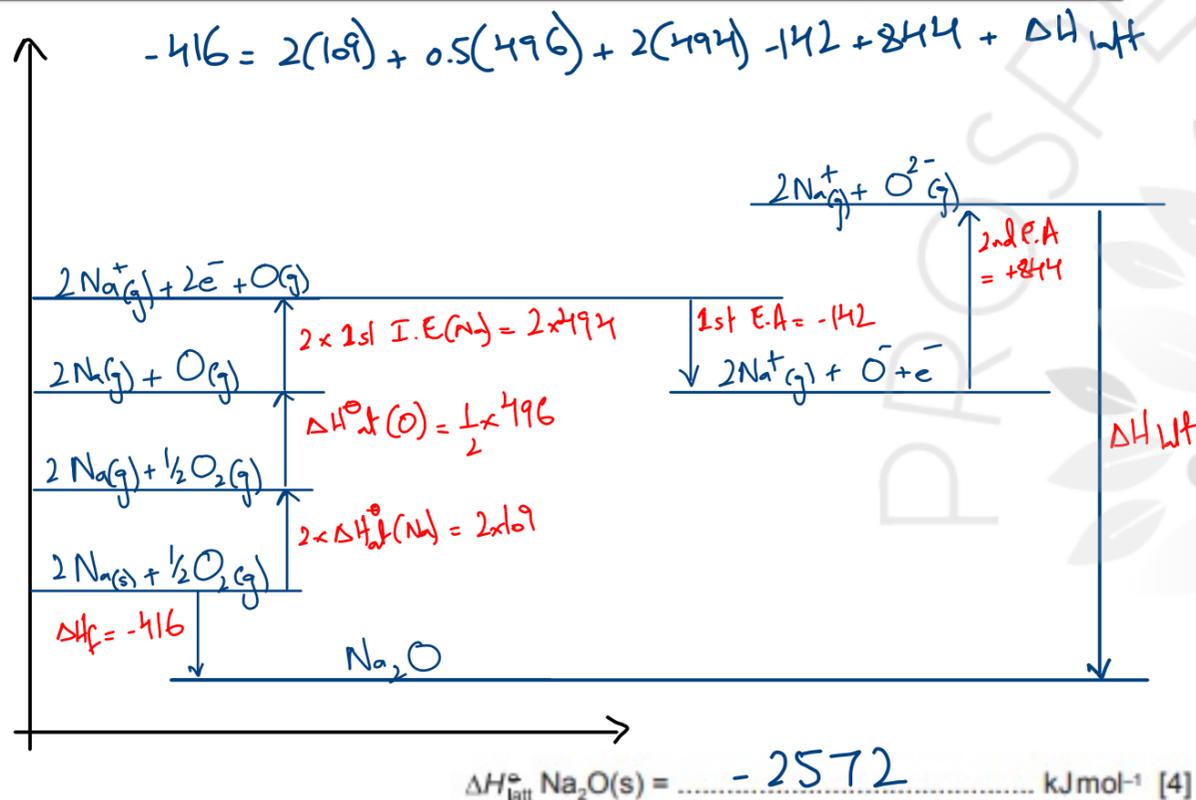
Suggest why the lattice energy of $\text{NaN}_3(\text{s})$, -732 kJ mol^{-1} , is more exothermic than that of $\text{RbN}_3(\text{s})$.

Na^+ has a smaller radius than Rb^+ so it forms a stronger ionic bond and more energy is released [1]

[Total: 11]

(d) Use the data below, and other suitable data from the *Data Booklet*, to calculate the lattice energy of sodium oxide, $\Delta H_{\text{latt}}^{\ominus} \text{Na}_2\text{O}(\text{s})$.

| energy change | value/ kJ mol^{-1} |
|---|-----------------------------|
| standard enthalpy change of formation of sodium oxide, $\Delta H_f^{\ominus} \text{Na}_2\text{O}(\text{s})$ | -416 |
| standard enthalpy change of atomisation of sodium, $\Delta H_{\text{at}}^{\ominus} \text{Na}(\text{s})$ | +109 |
| electron affinity of O(g) | -142 |
| electron affinity of O ⁻ (g) | +844 |



(e) State how $\Delta H_{\text{latt}}^{\ominus} \text{Na}_2\text{S}(\text{s})$ differs from $\Delta H_{\text{latt}}^{\ominus} \text{Na}_2\text{O}(\text{s})$. $\text{S}^{2-} \text{ radius} > \text{O}^{2-} \text{ radius}$
Indicate this by placing a tick (✓) in the appropriate box in the table.

| $\Delta H_{\text{latt}}^{\ominus} \text{Na}_2\text{S}(\text{s})$ is more exothermic than $\Delta H_{\text{latt}}^{\ominus} \text{Na}_2\text{O}(\text{s})$ | $\Delta H_{\text{latt}}^{\ominus} \text{Na}_2\text{S}(\text{s})$ is the same as $\Delta H_{\text{latt}}^{\ominus} \text{Na}_2\text{O}(\text{s})$ | $\Delta H_{\text{latt}}^{\ominus} \text{Na}_2\text{S}(\text{s})$ is less exothermic than $\Delta H_{\text{latt}}^{\ominus} \text{Na}_2\text{O}(\text{s})$ |
|---|--|---|
| | | ✓ |

Explain your answer.

The radius of S^{2-} is greater than O^{2-} so the ionic bond formed in Na_2S is weaker and so less energy is released.

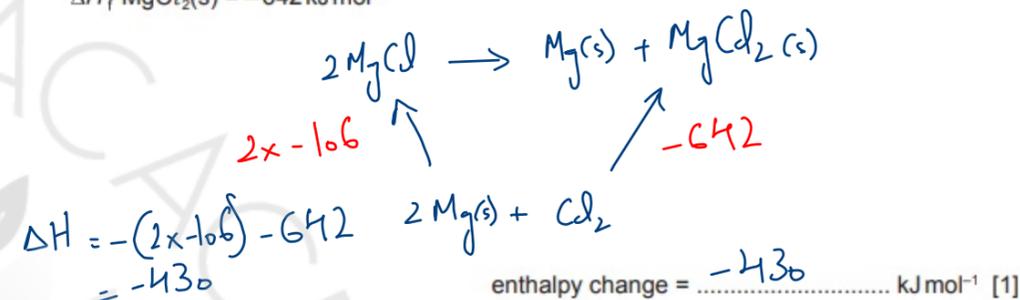
(b) Magnesium(I) chloride, MgCl , is an unstable compound and readily decomposes as shown.



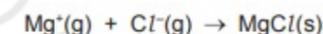
Use the following data to calculate the enthalpy change of this reaction.

$$\Delta H_f^{\ominus} \text{MgCl}(\text{s}) = -106 \text{ kJ mol}^{-1}$$

$$\Delta H_f^{\ominus} \text{MgCl}_2(\text{s}) = -642 \text{ kJ mol}^{-1}$$



(c) (i) The equation for which ΔH is the lattice energy for MgCl is shown.

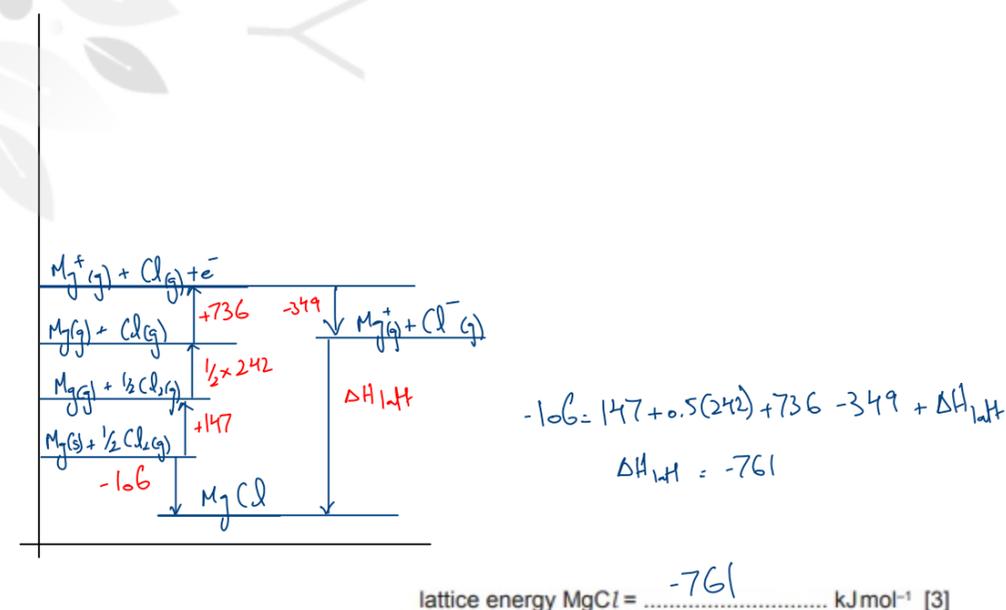


Use the equation, the following data, and relevant data from the *Data Booklet* to calculate a value for the lattice energy of MgCl . You might find it helpful to construct an energy cycle.

$$\text{electron affinity of Cl}(\text{g}) = -349 \text{ kJ mol}^{-1}$$

$$\text{enthalpy change of atomisation of Mg}(\text{s}) = +147 \text{ kJ mol}^{-1}$$

$$\text{enthalpy change of formation of MgCl}(\text{s}) = -106 \text{ kJ mol}^{-1}$$



(ii) Suggest how the lattice energies of MgCl_2 and NaCl will compare to that of MgCl . Explain your answers.

MgCl_2 and MgCl ... The lattice energy of MgCl_2 will be more exothermic as Mg has 2+ charge in MgCl_2 while only 1+ charge in MgCl .

NaCl and MgCl ... The lattice energy of NaCl will be more exothermic as Na^+ has a smaller ionic radius than Mg^+ (Na loses shell, Mg does not) [3]

Enthalpy change of hydration:- Energy released when 1 mol of gaseous ion dissolves in sufficient water to form a very dilute solution at standard conditions



Enthalpy change of solution:- Energy released or absorbed when 1 mole of solute dissolves in a solvent to form an infinitely dilute solution.
↳ A solution that produces no further enthalpy change upon addition of more solute.



Factors affecting Hydration enthalpy:- Always exothermic as intermolecular bonds are made when dissolving.

$\Delta H_{\text{hyd}}^{\ominus} \propto \frac{\text{charge}}{\text{radius}}$

- Greater charge means stronger bonds between polar water and ion (more energy released)
- smaller radius means stronger bonds between polar water and ion (more energy released)

Dissolving can be thought of as:-

- 1) Breaking an ionic compound into its gaseous ions (Basically reverse of Lattice energy: $\text{NaCl}(\text{s}) \longrightarrow \text{Na}^+(\text{g}) + \text{Cl}^-(\text{g})$)
- 2) Hydration of the ions (Basically standard enthalpy change of hydration)

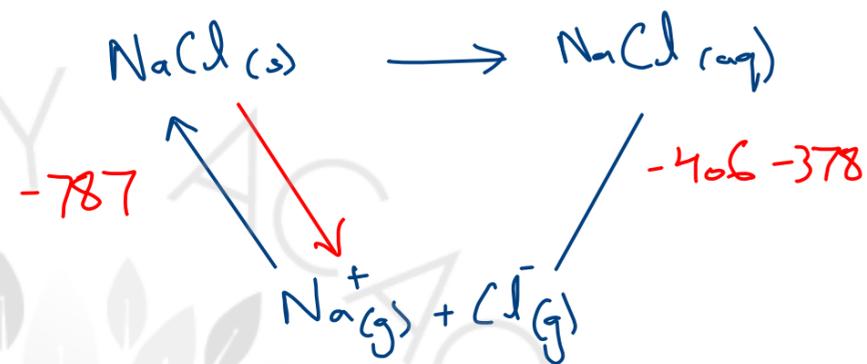
$$\therefore \Delta H_{\text{sol}}^{\ominus} = \sum \Delta H_{\text{hyd}}^{\ominus} - \Delta H_{\text{latt}}^{\ominus}$$

Q. Draw an enthalpy cycle to find $\Delta H^\ominus_{\text{sol}}$ of NaCl, given:

- $\Delta H^\ominus_{\text{latt}}(\text{NaCl}) = -787 \text{ kJ mol}^{-1}$

- $\Delta H^\ominus_{\text{hyd}}(\text{Na}^+) = -406 \text{ kJ mol}^{-1}$

- $\Delta H^\ominus_{\text{hyd}}(\text{Cl}^-) = -378 \text{ kJ mol}^{-1}$



$$\Delta H_{\text{sol}} = 787 - 406 - 378 = +3 \text{ kJ mol}^{-1}$$

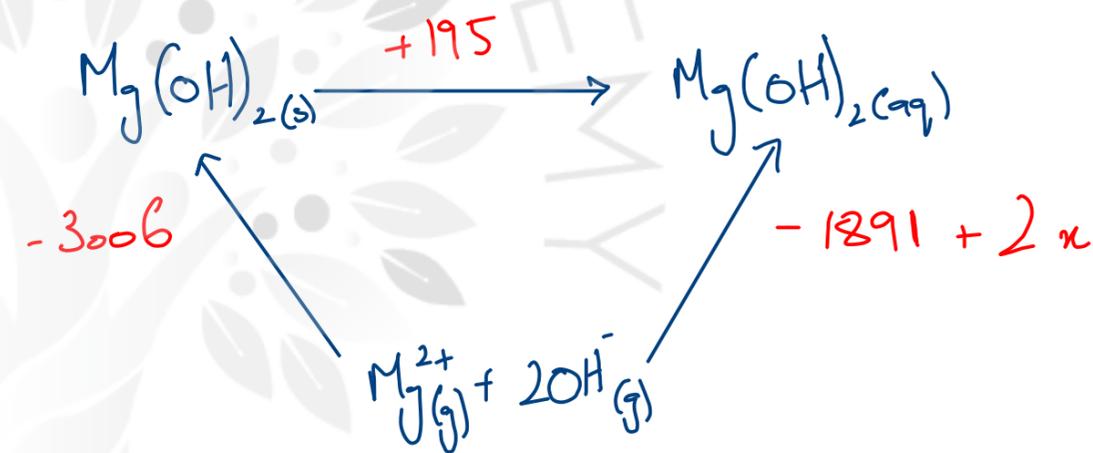
Q. Given the following information, calculate the $\Delta H^\ominus_{\text{hyd}}(\text{OH}^-)$:-

- $\Delta H^\ominus_{\text{latt}} \text{Mg(OH)}_2 = -3006 \text{ kJ mol}^{-1}$

- $\Delta H^\ominus_{\text{hyd}}(\text{Mg}^{2+}) = -1891 \text{ kJ mol}^{-1}$

- $\Delta H^\ominus_{\text{sol}} \text{Mg(OH)}_2 = +195$

↳ this is why Mg(OH)_2 is insoluble, it requires too much energy to dissolve



$$+195 = +3006 - 1891 + 2x$$

$$x = -460 \text{ kJ mol}^{-1}$$

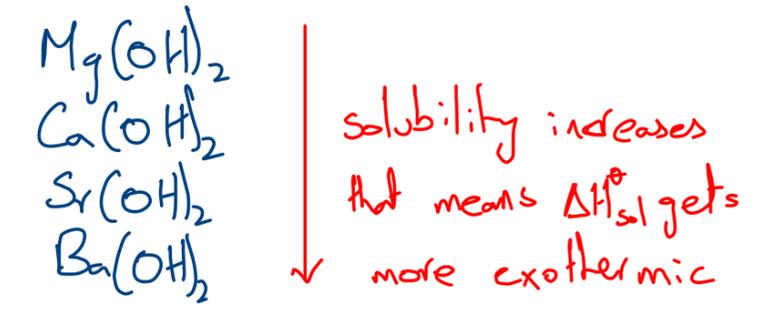
Solubility of Group 2 hydroxides and Sulfates:-

- if ΔH_{sol}^\ominus is +ve then dissolving requires energy / is endothermic which is unfavourable
- if ΔH_{sol}^\ominus is -ve then dissolving requires energy / is exothermic which is favourable

$$+ \Delta H_{sol}^\ominus = -\sum \Delta H_{hyd}^\ominus + \Delta H_{latt}^\ominus$$

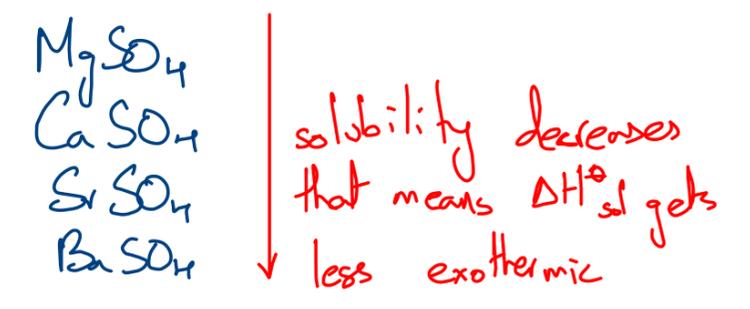
$$- \Delta H_{sol}^\ominus = -\sum \Delta H_{hyd}^\ominus + \Delta H_{latt}^\ominus$$

Group 2 hydroxides:-



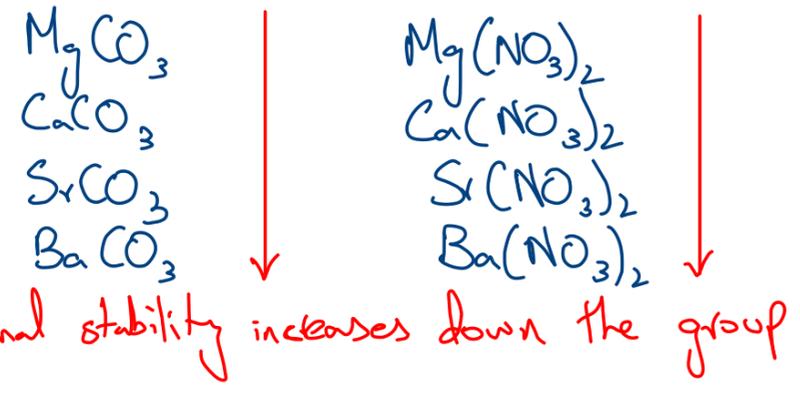
- Group 2 hydroxides become more soluble down the group as
- cationic radius increases
 - the lattice energy decreases much more in magnitude than the hydration enthalpy
 - making ΔH_{sol}^\ominus more exothermic and favouring dissolving

Group 2 Sulfates:-

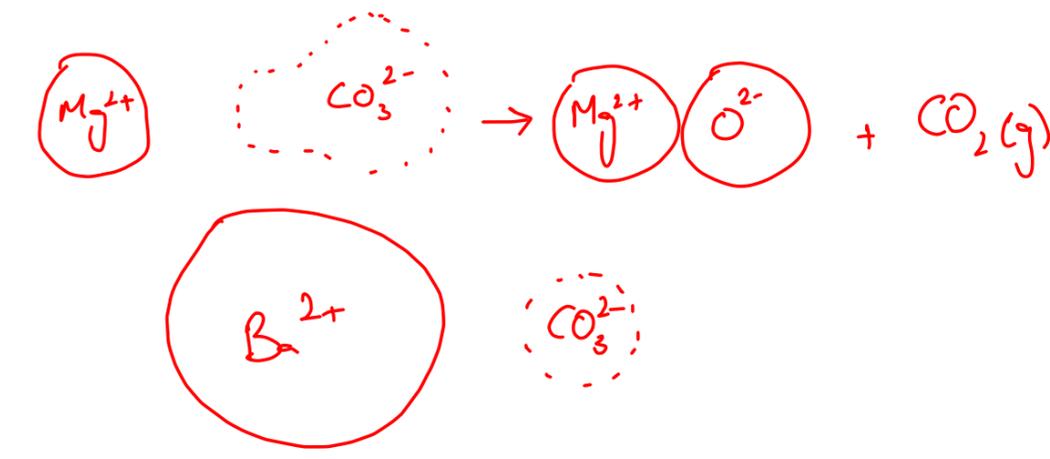


- Group 2 sulfates become less soluble down the group as
- cationic radius increases
 - the hydration enthalpy decreases much more in magnitude than the lattice energy
 - making ΔH_{sol}^\ominus less exothermic and unfavouring dissolving

Thermal Stability of Group 2 Carbonates and nitrates



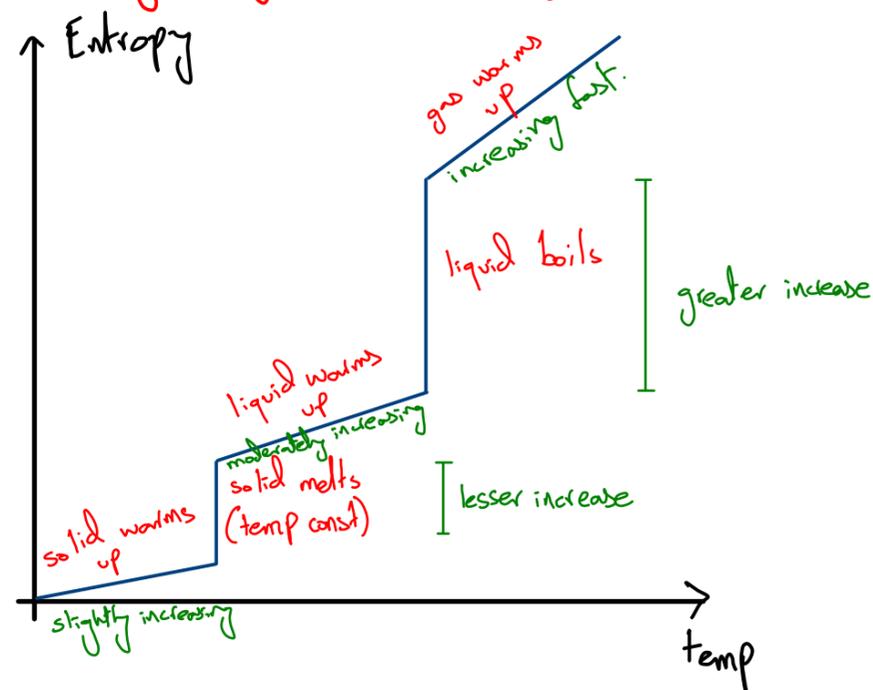
- Down the group, thermal stability increases as:-
- cationic radius increases
 - polarising power of cation decreases
 - It polarises CO_3^{2-} and NO_3^- ion less strongly
 - therefore compound stays stable and does not decompose



Entropy:- Entropy is the number of possible arrangements of the particles and their energy in a given system.

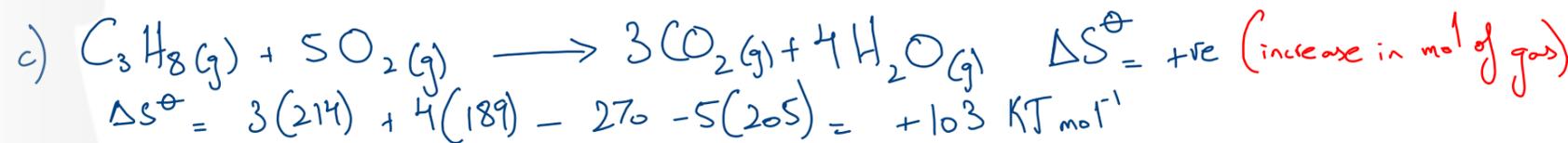
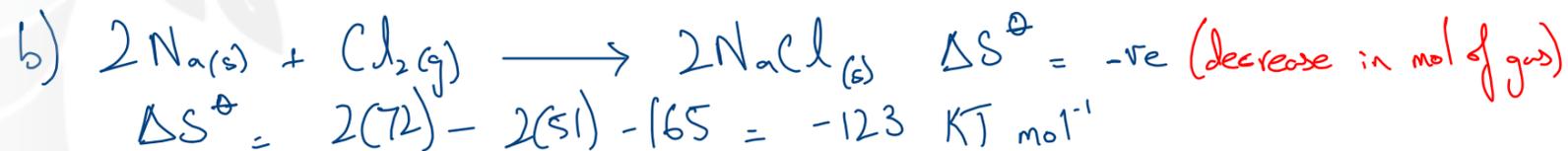
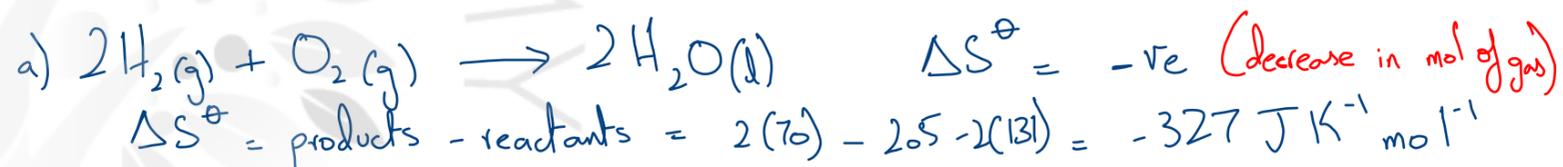
- Entropy increases with temperature as temp increases Kinetic energy of particles.
- Entropy increases with increasing number of particles as this increases the number of ways particles can be arranged in.
- Entropy is greatest in gases then liquids then solids

$$\Delta S^\ominus = \text{products} - \text{reactants}$$

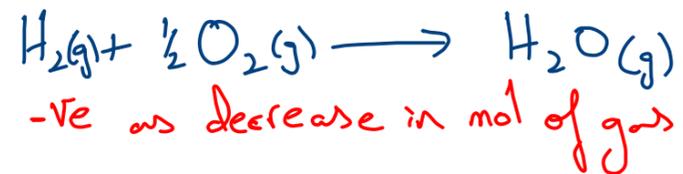
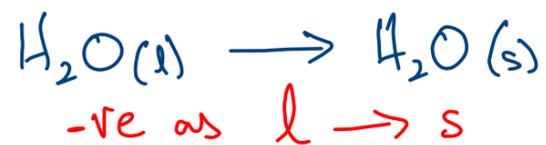
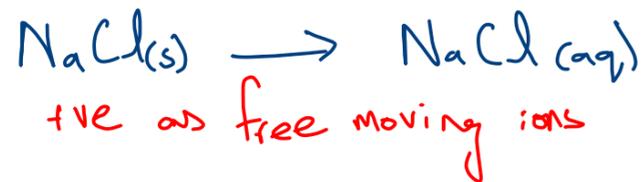


Q. Predict the change in entropy and then verify by calculation

| Substance | Entropy (S^\ominus) / $\text{JK}^{-1} \text{mol}^{-1}$ |
|----------------------------------|--|
| $\text{H}_2(\text{g})$ | 131 |
| $\text{O}_2(\text{g})$ | 205 |
| $\text{H}_2\text{O}(\text{l})$ | 70 |
| $\text{H}_2\text{O}(\text{g})$ | 189 |
| $\text{Na}(\text{s})$ | 51 |
| $\text{Cl}_2(\text{g})$ | 165 |
| $\text{NaCl}(\text{s})$ | 72 |
| $\text{NH}_3(\text{g})$ | 192 |
| $\text{HCl}(\text{g})$ | 187 |
| $\text{NH}_4\text{Cl}(\text{s})$ | 95 |
| $\text{C}_3\text{H}_8(\text{g})$ | 270 |
| $\text{CO}_2(\text{g})$ | 214 |



Predict the ΔS^\ominus :-



Gibbs Free Energy change, ΔG^\ominus :-

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

\downarrow change in Gibbs free energy
 \downarrow enthalpy change of reaction in kJ mol^{-1}
 \downarrow T in K
 \downarrow change in entropy in $\text{kJ mol}^{-1} \text{K}^{-1}$

- Tells us if a reaction is feasible or not
- if $\Delta G^\ominus = -ve \rightarrow$ reaction feasible / spontaneous
- if $\Delta G^\ominus = +ve \rightarrow$ reaction not feasible / spontaneous

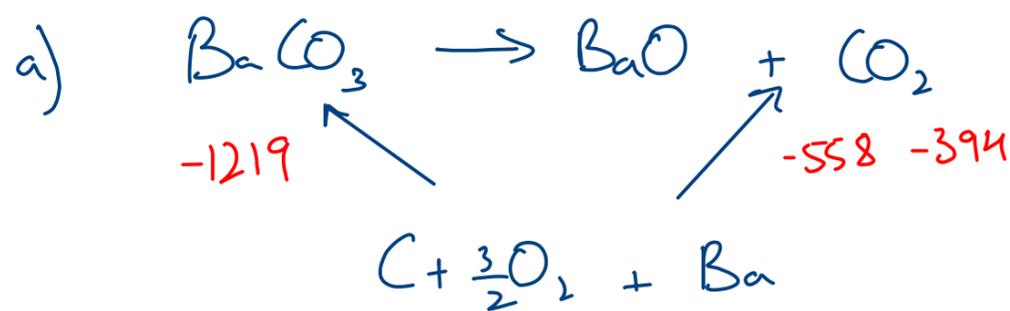
- C + 273.15 = K - Convert S^\ominus to $\text{kJ mol}^{-1} \text{K}^{-1}$ by $\times 10^{-3}$

Q. The decomposition of barium carbonate is given by: - $\text{BaCO}_3(\text{s}) \rightarrow \text{BaO}(\text{s}) + \text{CO}_2(\text{g})$

The following data is also given: -

| | |
|--|--|
| $\Delta H_f^\ominus (\text{BaCO}_3(\text{s})) = -1219 \text{ kJ mol}^{-1}$ | $S^\ominus (\text{BaCO}_3(\text{s})) = +112 \text{ J K}^{-1} \text{ mol}^{-1}$ |
| $\Delta H_f^\ominus (\text{BaO}(\text{s})) = -558 \text{ kJ mol}^{-1}$ | $S^\ominus (\text{BaO}(\text{s})) = +70 \text{ J K}^{-1} \text{ mol}^{-1}$ |
| $\Delta H_f^\ominus (\text{CO}_2(\text{g})) = -394 \text{ kJ mol}^{-1}$ | $S^\ominus (\text{CO}_2(\text{g})) = +214 \text{ J K}^{-1} \text{ mol}^{-1}$ |

- Calculate the values of ΔH^\ominus , ΔS^\ominus and ΔG^\ominus for the forward reaction at 25°C
- State with a reason whether the reaction is spontaneous at 25°C
- Calculate the temperature at when the reaction first becomes spontaneous.



$$\Delta H = +1219 - 558 - 394 = +267 \text{ kJ mol}^{-1}$$

$$\Delta S^\ominus = \text{products} - \text{reactants} = 70 + 214 - 112 = +172 \text{ J K}^{-1} \text{ mol}^{-1}$$

$$\Delta G^\ominus = \Delta H^\ominus - T \Delta S^\ominus = 267 - (25 + 273.15)(172 \times 10^{-3}) = 215.7 \text{ kJ mol}^{-1} \approx 216 \text{ kJ mol}^{-1}$$

b) Not spontaneous as $\Delta G^\ominus > 0$

c) $\Delta G^\ominus = \Delta H^\ominus - T \Delta S^\ominus$

$$0 > 267 - T(172 \times 10^{-3})$$

$$T > \frac{267}{172 \times 10^{-3}} = 1552.3 \text{ K}$$

Spontaneous if $T > 1550 \text{ K}$

Important things to know regarding ΔG^\ominus :-

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

| ΔH^\ominus | ΔS^\ominus | ΔG^\ominus |
|--------------------|--------------------|--------------------|
| -ve | +ve | Always -ve |
| -ve | -ve | -ve / +ve |
| +ve | +ve | -ve / +ve |
| +ve | -ve | Always +ve |

Feasibility of Reaction

Always feasible

feasible only if T is low enough

feasible only if T is high enough

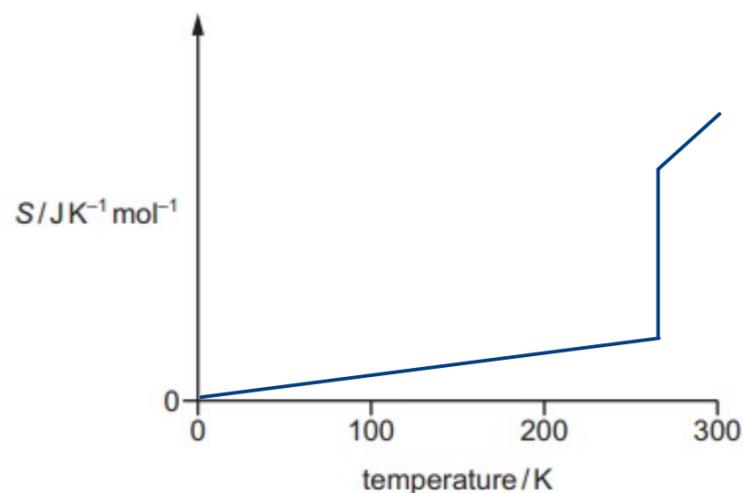
Never feasible

- A feasible/spontaneous reaction is very likely to take place but its not necessary it always takes place, it may have a high activation energy for e.g. the burning of methane

8 Entropy is a measure of the disorder of a system.

(a) Assume the entropy, S, for H_2O is zero at 0K.

Sketch a graph on the axes to show how the entropy changes for H_2O between 0K and 300K.



[2]

(b) Place one tick (✓) in each row of the table to show the sign of the entropy changes, ΔS .

| | ΔS is negative | ΔS is positive |
|---------------------------|------------------------|------------------------|
| solid dissolving in water | | ✓ |
| water boiling to steam | | ✓ |

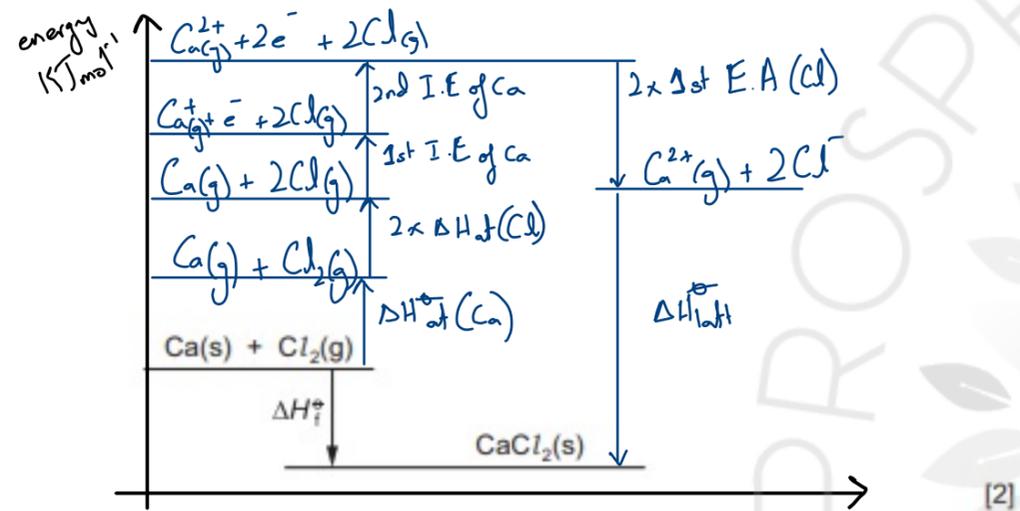
[1]

2 (a) Calcium metal reacts with chlorine gas to form calcium chloride, CaCl_2 .

(i) Write an equation, including state symbols, to represent the lattice energy of calcium chloride, CaCl_2 .



(ii) Complete a fully labelled Born-Haber cycle that could be used to calculate the lattice energy, $\Delta H_{\text{latt}}^{\ominus}$, for calcium chloride.



(iii) Use your answer to (ii) and the following data, together with relevant data from the Data Booklet, to calculate a value for $\Delta H_{\text{latt}}^{\ominus}$ for calcium chloride.

| | |
|---|----------------------------|
| standard enthalpy change of formation of $\text{CaCl}_2(\text{s})$, ΔH_f^{\ominus} | -796 kJ mol^{-1} |
| standard enthalpy change of atomisation of $\text{Ca}(\text{s})$, $\Delta H_{\text{at}}^{\ominus}$ | $+178 \text{ kJ mol}^{-1}$ |
| electron affinity of chlorine atoms | -349 kJ mol^{-1} |

$$-796 = 178 + 2\left(\frac{1}{2} \times 242\right) + 590 + 1150 + 2(-349) + \Delta H_{\text{latt}}^{\ominus}$$

$$\Delta H_{\text{latt}}^{\ominus} = -2258$$

$$\Delta H_{\text{latt}}^{\ominus} = -2258 \text{ kJ mol}^{-1} \quad [3]$$

(b) Entropy is a measure of the disorder of a system.

Describe and explain what happens to the entropy of a gas when the temperature is increased.

The kinetic energy of the particles increases and so the entropy of a system also increases as this increases the disorder in a system.

[2]

(c) The table shows four reactions.

(i) For each reaction, predict the sign of the entropy change, ΔS^{\ominus} . If you predict no entropy change, write 'no change' in the table below. The first one has been done for you.

| reaction | sign of ΔS^{\ominus} |
|--|------------------------------|
| $\text{CO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$ | negative |
| $\text{Mg}(\text{s}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{MgO}(\text{s})$ | negative |
| $\text{CuSO}_4(\text{s}) + 5\text{H}_2\text{O}(\text{l}) \rightarrow \text{CuSO}_4 \cdot 5\text{H}_2\text{O}(\text{s})$ | negative |
| $\text{NaHCO}_3(\text{s}) + \text{H}^+(\text{aq}) \rightarrow \text{Na}^+(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$ | positive |

[2]

(ii) Explain why the entropy change for the first process is negative.

Moles of gas decreased

[1]

(d) Calculate the standard entropy change, ΔS^{\ominus} , for this reaction.



Standard entropies, S^{\ominus} , in $\text{JK}^{-1} \text{mol}^{-1}$ are given.

| | | |
|------------------------|------------------------|-------------------------|
| $\text{N}_2(\text{g})$ | $\text{H}_2(\text{g})$ | $\text{NH}_3(\text{g})$ |
| +192 | +131 | +193 |

$$\Delta S^{\ominus} = 2(193) - 3(131) - 192 = -199$$

$$\Delta S^{\ominus} = -199 \text{ JK}^{-1} \text{mol}^{-1} \quad [2]$$

(e) Whether or not a chemical reaction is spontaneous (feasible) can be deduced by calculating the change in free energy, ΔG^{\ominus} , at a given temperature.



$$\Delta H^{\ominus} = +117 \text{ kJ mol}^{-1}$$

$$\Delta S^{\ominus} = +175 \text{ JK}^{-1} \text{mol}^{-1}$$

(i) Calculate the value of ΔG^{\ominus} at 298 K for the above reaction.

$$\Delta G^{\ominus} = 117 - 298(175 \times 10^{-3})$$

$$= +64.85 = 64.9 \text{ kJ mol}^{-1}$$

[2]

(ii) Use your answer to (i) to explain whether or not this reaction is spontaneous at 298 K.

Not spontaneous as ΔG^{\ominus} is positive

[1]

(c) (i) Silicon tetrachloride can be prepared according to reaction 1.



| | |
|---|--|
| standard entropy of silicon, $S^\ominus \text{ Si(s)}$ | $18.7 \text{ JK}^{-1} \text{ mol}^{-1}$ |
| standard entropy of silicon tetrachloride, $S^\ominus \text{ SiCl}_4(\text{l})$ | $239.0 \text{ JK}^{-1} \text{ mol}^{-1}$ |

Calculate the standard entropy of chlorine, $S^\ominus \text{ Cl}_2(\text{g})$. Show all your working.

$$\Delta S^\ominus = \text{products} - \text{reactants}$$

$$-225.7 = 239 - (18.7 + 2x)$$

$$x = 223$$

$$S^\ominus \text{ Cl}_2(\text{g}) = 223 \text{ JK}^{-1} \text{ mol}^{-1} \quad [2]$$

(ii) Explain why the entropy change for reaction 1 is negative.

decrease in moles of gas

[1]

(d) The standard enthalpy change of formation of silicon tetrachloride, $\Delta H_f^\ominus \text{ SiCl}_4(\text{l})$, is -640 kJ mol^{-1} .

Reaction 1 is spontaneous at lower temperatures, but it is not spontaneous at very high temperatures.

Calculate the temperature above which reaction 1 is not spontaneous.

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

$$0 < -640 - T(-225.7 \times 10^{-3})$$

$$T > \frac{+640}{225.7 \times 10^{-3}} \Rightarrow T > 2835.6$$

temperature = 2840 K [2]

[Total: 13]

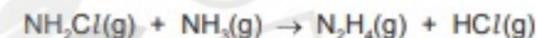
(c) Some values for standard enthalpy changes of formation, ΔH_f^\ominus , and standard entropies, S^\ominus , are given in the table.

| | $\Delta H_f^\ominus / \text{kJ mol}^{-1}$ | $S^\ominus / \text{JK}^{-1} \text{ mol}^{-1}$ |
|----------------------------------|---|---|
| $\text{NH}_2\text{Cl(g)}$ | +80.1 | +241 |
| $\text{NH}_3(\text{g})$ | -45.9 | +198 |
| $\text{N}_2\text{H}_4(\text{g})$ | +95.4 | +237 |
| HCl(g) | -92.3 | +187 |

(i) Define the meaning of the term *entropy*.

It is the measure of number of possible arrangements of particles and their energies in a given system. [1]

Hydrazine, N_2H_4 , can be produced from chloramine and ammonia as shown.



(ii) Calculate the standard entropy change, ΔS^\ominus , for this reaction.

$$\Delta S = 187 + 237 - (241 + 198)$$

$$= -15$$

$$\Delta S^\ominus = -15 \text{ JK}^{-1} \text{ mol}^{-1} \quad [1]$$

(iii) Calculate the standard enthalpy change, ΔH^\ominus , for this reaction.

$$\text{NH}_2\text{Cl(g)} + \text{NH}_3(\text{g}) \rightarrow \text{N}_2\text{H}_4(\text{g}) + \text{HCl(g)}$$

$$+80.1 - 45.9 \quad +95.4 - 92.3$$

$$\frac{5}{2} \text{H}_2 + \text{N}_2 + \frac{1}{2} \text{Cl}_2$$

$$\Delta H_f^\ominus = -80.1 + 45.9 + 95.4 - 92.3$$

$$\Delta H^\ominus = -31.1 \text{ kJ mol}^{-1} \quad [1]$$

(iv) Calculate the standard Gibbs free energy change, ΔG^\ominus , for this reaction at 298 K.

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

$$= -31.1 - 298(-15 \times 10^{-3})$$

$$= -26.63$$

$$\Delta G^\ominus = -26.6 \text{ kJ mol}^{-1} \quad [2]$$

(v) Explain, with reference to ΔG^\ominus , why this reaction becomes less feasible at higher temperatures.

As temperature increases the product $-T \Delta S^\ominus$ becomes a bigger positive value which will lead to a positive ΔG^\ominus making reaction less feasible [1]