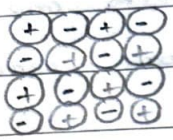
PAGE No.
DATE / /

A-level chemistry

⇒ Lattice energy

→ Ionic lattice structure



→ Important definitions

→ Lattice Energy - The enthalpy change when one mole of an ionic lattice is formed from its gaseous ions under standard conditions. Always Exothermic.

- Example. $Na^+(g) + Cl^-(g) \rightarrow NaCl(s)$

→ Atomisation - The enthalpy change when one mole of gaseous atoms is formed from its element in standard state, under standard conditions. Always Endothermic.

- Example $Na(s) \rightarrow Na(g)$
- Example $\frac{1}{2} Cl_2(g) \rightarrow Cl(g)$

→ Hydration - The enthalpy change when one mole of gaseous ions is dissolved in excess water to form an infinitely dilute solution. Always Exothermic

- Example $Na^+(g) \rightarrow Na^+(aq)$

→ The enthalpy change when one mole ionic solid is dissolved in excess water to form an infinitely dilute solution. Either endothermic or exothermic. Under standard conditions. This the standard enthalpy change of solution.

→ Example $NaCl(s) \rightarrow NaCl(aq)$

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

→ Important definitions (Continued)

→ First electron affinity - The enthalpy change when one mole of electrons is added to one mole of gaseous atoms to form one mole of gaseous 1^- ions under standard conditions. Generally Exothermic

→ Example - $O(g) + 1e^- \rightarrow O^-(g)$

→ Second electron affinity - The enthalpy change when one mole of electrons is added to one mole of ~~gaseous~~ gaseous 1^- ions to form one mole of 2^- ions under standard conditions. Endothermic

→ Example - $O^-(g) + 1e^- \rightarrow O^{2-}(g)$

→ Standard enthalpy change of formation - The enthalpy change when one mole of a compound is formed from its elements under standard conditions. The elements must be in their standard state.

→ Example: $Na(s) + \frac{1}{2}Cl_2(g) \rightarrow NaCl(s)$

⇒ Calculating lattice energy using born-haber cycles

$$Na(s) + \tfrac{1}{2}Cl_2(g) \xrightarrow{\Delta H_f^\circ} NaCl(s)$$

ΔH ↘ ↗ Lattice energy

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

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

First IE ↑ $Na(g) + Cl(g)$ — First EA ↓ $Na^+(g) + Cl^-(g)$

ΔH°_{atm} of Cl ↑ $Na(g) + \frac{1}{2}Cl_2(g)$

ΔH°_{atm} of Na ↑ $Na(s) + \frac{1}{2}Cl_2(g)$

ΔH°_f of NaCl ↓ LE ↓ $NaCl(s)$

$\Delta H_f^\circ = \Delta H + LE$

$LE = \Delta H_f^\circ - \Delta H$

$= \Delta H_f^\circ - (\Delta H^\circ_{atm}\,Na + \Delta H^\circ_{atm}\,Cl + \text{First IE} + \text{First EA})$

PAGE No.
DATE / /

Lattice → gaseous ions to compound

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

Atomisation → element in standard state to gaseous 1 mole.

$Na(s) \rightarrow Na(g)$

$\frac{1}{2}Cl_2(g) \rightarrow Cl(g)$

Hydration → gaseous ions to aqueous ions

$Na^+(g) \rightarrow Na^+(aq)$

Solution → ionic solid to aqueous

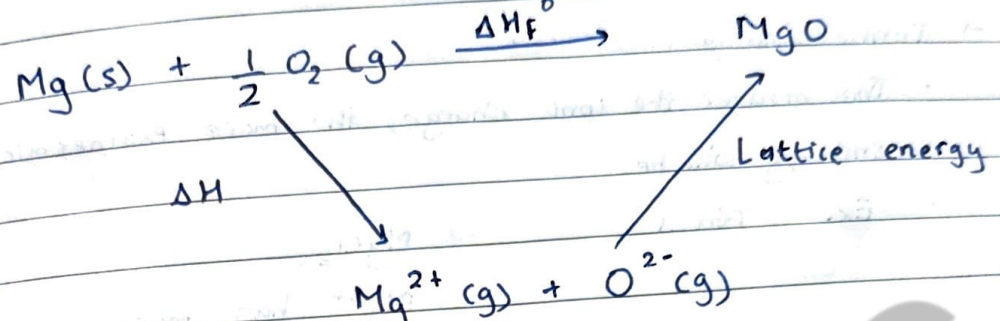
$NaCl(s) \rightarrow NaCl(aq)$

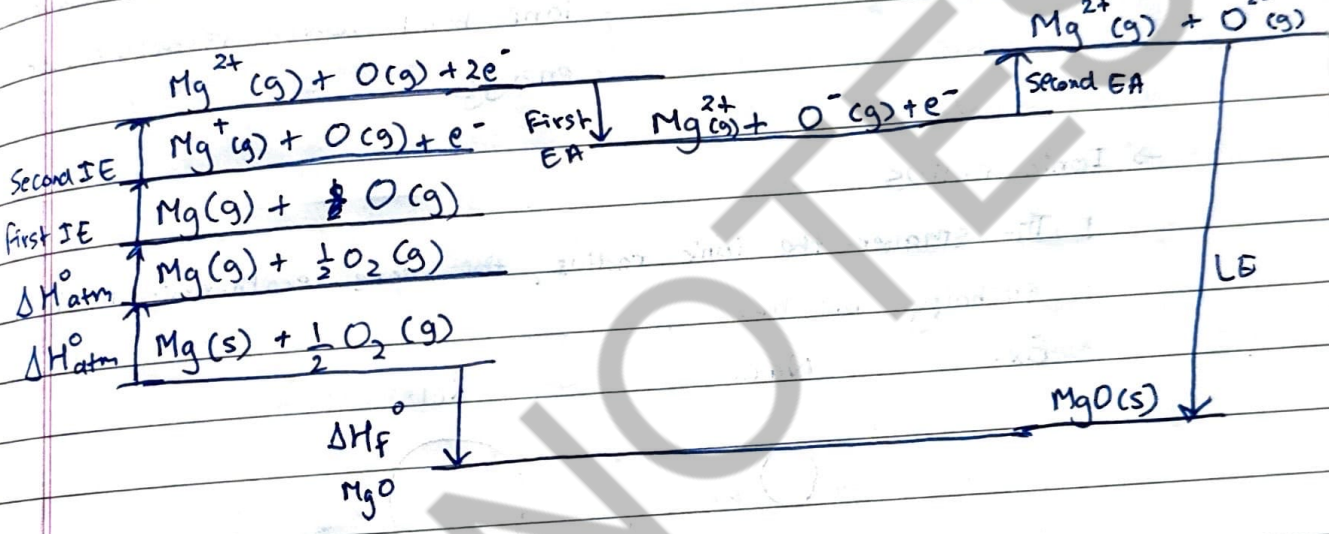
$NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$

Formation → elements in standard state to compound

$Na(s) + \frac{1}{2}Cl_2(g) \rightarrow NaCl(s)$

Lattice energy of MgO using Born-Haber





⇒ Diatomic molecules

ΔH°_{atm}	Bond energy	$\Delta H^\circ_{atm} = \frac{1}{2}$ BE
Cl_2 (g)	Cl_2 (g) ⟷ 2Cl (g) BE	$\frac{1}{2} Cl_2(g) \rightarrow Cl(g)\ \Delta H^\circ_{atm}$
O_2 (g)		
F_2 (g)		
N_2 (g)		
H_2 (g)		

PAGE No.
DATE / /

⇒ Factors affecting the values of lattice energy

→ Ionic Charge

- The greater the ionic charge, the more exothermic the lattice enthalpy will be.
- Ex. NaCl $MgCl_2$

Mg^{2+} has more charge so attraction is stronger, so it forms a stronger ionic bond which releases more energy.

→ Ionic radius

- The smaller the ionic radius, the more exothermic the lattice enthalpy will be.
- Ex. NaCl KCl

less volume
more charge density

more volume
less charge density

So more charge density means more ionic bond which means more energy is released.

⇒ Ion Polarisation

(+)(-) → Pure ionic → more exothermic

but in reality the force of the positive ion distorts the negative ion

(+)<(-) → this gives the bond some covalent properties
→ this weakens the ionic bonds.
→ No ionic compound is pure ionic.
→ less exothermic

PAGE No.
DATE / /

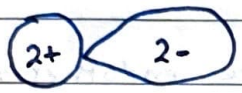
⇒ Charge density

→ measure of electric charge per unit volume.

→ if the ionic radius is small, volume is less, if volume is less charge density is high.

→ Example MgO

Mg^{2+} O^{2-}

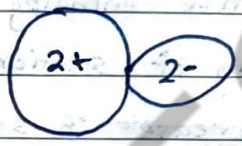


as it has high charge density it distorts the negative ion more.

as this is highly polarised / distorted the bond becomes weaker and thermal stability decreases

CaO

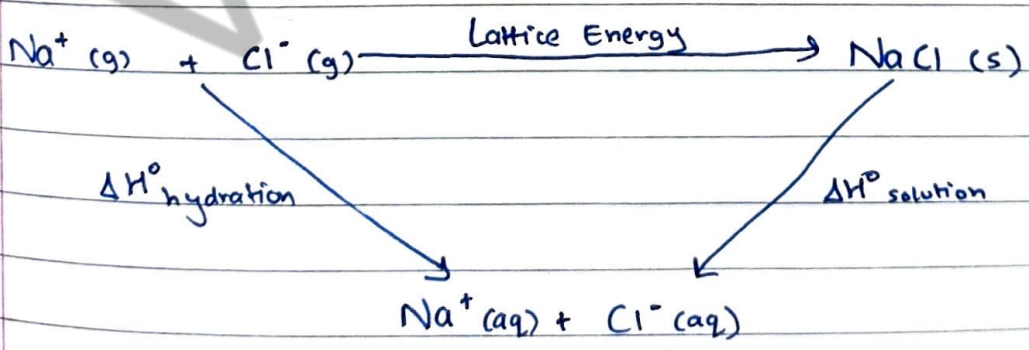
Ca^{2+} O^{2-}



as it has low charge density it distorts / polarises the negative ion less, compared to Mg^{2+}.

as this is not very polarised the bond does not become very weak. Hence it has a higher thermal stability compared to Mg.

⇒ Relation between lattice enthalpy, enthalpy of hydration, and enthalpy of solution



$\Delta H^\circ_{hydration} = LE + \Delta H^\circ_{solution}$

$\Delta H^\circ_{sol} = \Delta H^\circ_{hyd} - LE$

$= (\Delta H^\circ_{hyd\ Na} + \Delta H^\circ_{hyd\ Cl}) - LE$