CH 19

Lattice energy/Born Haber cycle

Lattice energy: is the enthalpy change when 1 mole of ionic compound is formed from its gaseous ions

* lattice energy is always negative → energy released when bonds are formed.
* Stronger ionic bonds → higher lattice energy → more exothermic

* factors affecting lattice energy:

[1] charge on the ions → the higher charge → stronger attraction forces → higher amound of energy released → higher lattice → more exothermic.

e.g. Na^{+1} Cl^{-1}

Mg^{+2} O^{-2} → higher lattice energy

[2] Radius of the ion → larger radius → weaker attraction forces → less energy released → less lattice → less exothermic

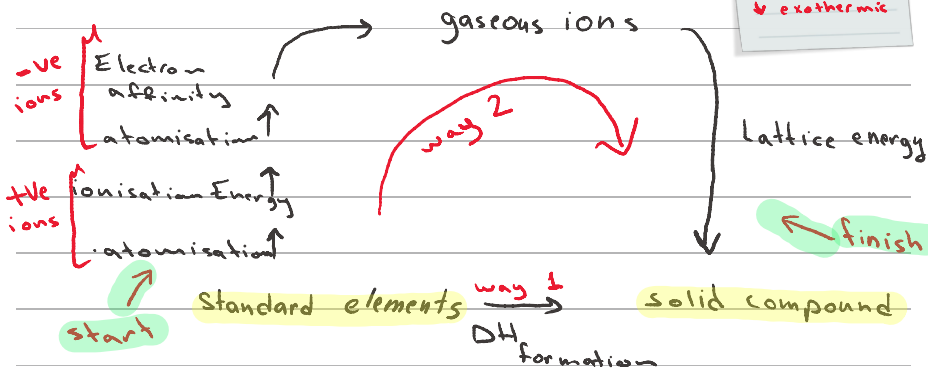
e.g. NaCl

NaI → lower lattice energy.

Born-Haber Cycle

arrow up ↑ endothermic

arrow down ↓ exothermic



way 1 = way 2

* **ΔH formation**: Enthalpy change when 1 mole of a compound is formed from its elements under standard conditions.

e.g. $Na_{(s)} + \frac{1}{2}Cl_2 \rightarrow NaCl_{(s)}$ ← 1 mole

* **ΔH atomisation**: Enthalpy change when one mole of gaseous atom is formed from its element under standard conditions

e.g. $Na_{(s)} \rightarrow Na_{(g)}$ (endothermic)

* ΔH Ionisation energy IE : Enthalpy change when 1 mole of ion is formed from 1 mole of gaseous atom by losing 1 mole of electrons under standard conditions

e.g. $Na_{(g)} \rightarrow Na^{+1}_{(g)} + 1\,e^-$ (endothermic)

* ΔH Electron affinity: ΔH_{EA_1} is the energy released when 1 mole of electrons is added to 1 mole of gaseous atom to form 1 mole of gaseous negative ion

remember:
breaking bonds ⇒ endothermic
making bonds ⇒ exothermic

e.g. $Cl_{(g)} + 1\,e^- \rightarrow Cl^-_{(g)}$ EA_1 is exothermic

EA_2 $Cl^-_{(g)} + 1\,e^- \rightarrow Cl^{-2}_{(g)}$ endothermic

because energy needed to overcome repulsion between −ve ion and incoming e^-

* factors affecting Electron affinity are same as factors affecting IE

1. nuclear charge → more protons → more attractive forces between p+ and incoming ē → EA increases

2. distance between nucleus and incoming ē → more radius → less attraction → EA decreases

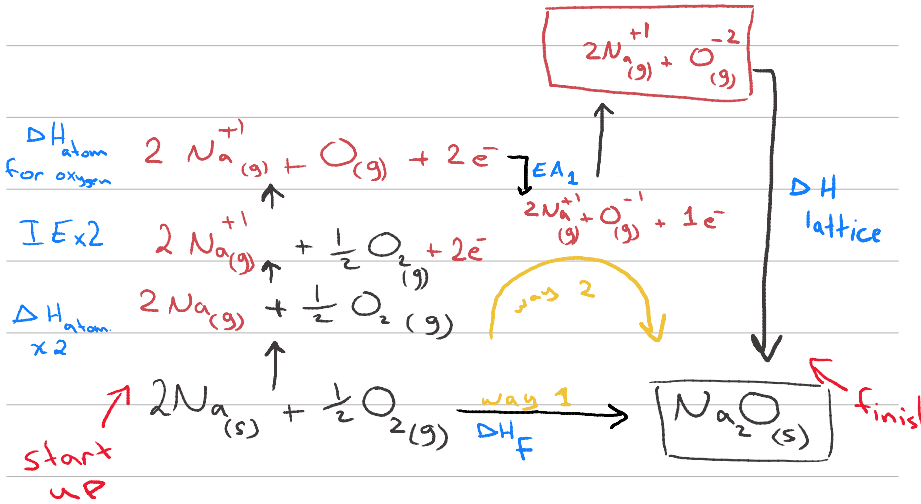
3. Shielding (increasing in inner shells) → less attraction between p+ and incoming ē → EA decreases

exception : Fluorine lower EA → F is very small with high electron density → repulsion between electrons in Fluorine and incoming ē → this decreases the attraction between nucleous and incoming ē → EA decreases

general trend

EA increases by going up-right periodic table.

→ let's construct Born-Haber cycle for Na_2O



way 1 = way 2

$\Delta H_F = (\Delta H_{atom\ Na} \times 2) + (IE_{Na} \times 2) + (\Delta H_{atom\ O_2}) + (EA_1\ O) + (EA_2\ O) + \Delta H_{latt}$

Enthalpy Changes in Solutions

* $\Delta H_{solution}$: is the enthalpy change when 1 mole of ionic compound is dissolved in water to form dilute solution under standard conditions.

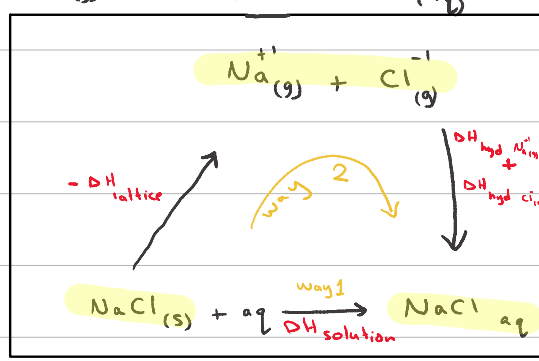
e.g. $MgCl_{2(s)} + aq \rightarrow MgCl_{2(aq)}$

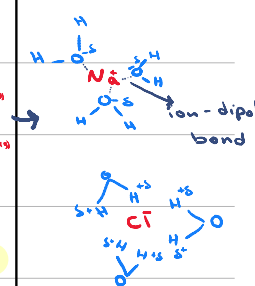
→ (can be exothermic or endothermic)

* $\Delta H_{hydration}$: is the enthalpy change when 1 mole of of gaseous ion dissolves in water to form hydrated ions under standard conditions → always exothermic (making)

e.g. $Mg^{+2}_{(g)} + aq \rightarrow Mg^{+2}_{(aq)}$

aq = water





way 1 = way 2

$\Delta H_{solu} = -\Delta H_{lattice} + \Delta H_{hyd}$

* factors affecting ΔH hydration

1. Charge on the ion → higher charge → stronger ion-polar attraction → higher $\Delta H_{hyd.}$ → more exothermic → more -ve.
2. size of ion → larger radius → less attraction → less ΔH_{hyd} → less exothermic → more +ve

* solubility in group II sulfates:

decreases down the Group. → $\Delta H_{soln} > 0$

why:

* $\Delta H_{soln} = \Delta H_{hyd.} - \Delta H_{Lattice}$ going down group
* both ΔH_{hyd} and ΔH lattice become less exothermic
* but $\Delta H_{hyd.}$ decreases faster than $\Delta H_{latt.}$
* → ΔH soln. becomes less -ve less exothermic

* solubility in group II hydroxides increases down the Group → $\Delta H_{soln.} < 0$

why * $\Delta H_{soln} = \Delta H_{hyd} - \Delta H_{lattice}$

* down group II both ΔH_{hyd} and $\Delta H_{latt.}$ become less exothermic
* but $\Delta H_{latt.}$ decreases faster * → ΔH soln. becomes more -ve.

Ion polarisation

happens when positive ion attracts the electrons in the negative ion and causes distortion of the electron clouds of the anion.

* factors affection inon polarisation:

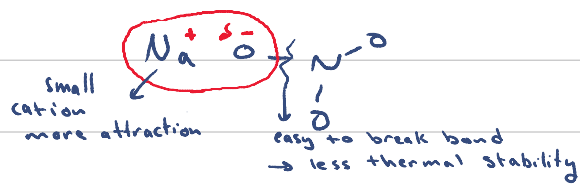
1. charge density of cation → higher charge and smaller radius → higher polarising power
2. size of anion

larger size of the anion → easier to be distorted → easier to be polarised.

e.g.

small ← $NaNO_3$

$CsNO_3$ → large



* mention polarising effect when asked about thermal stability