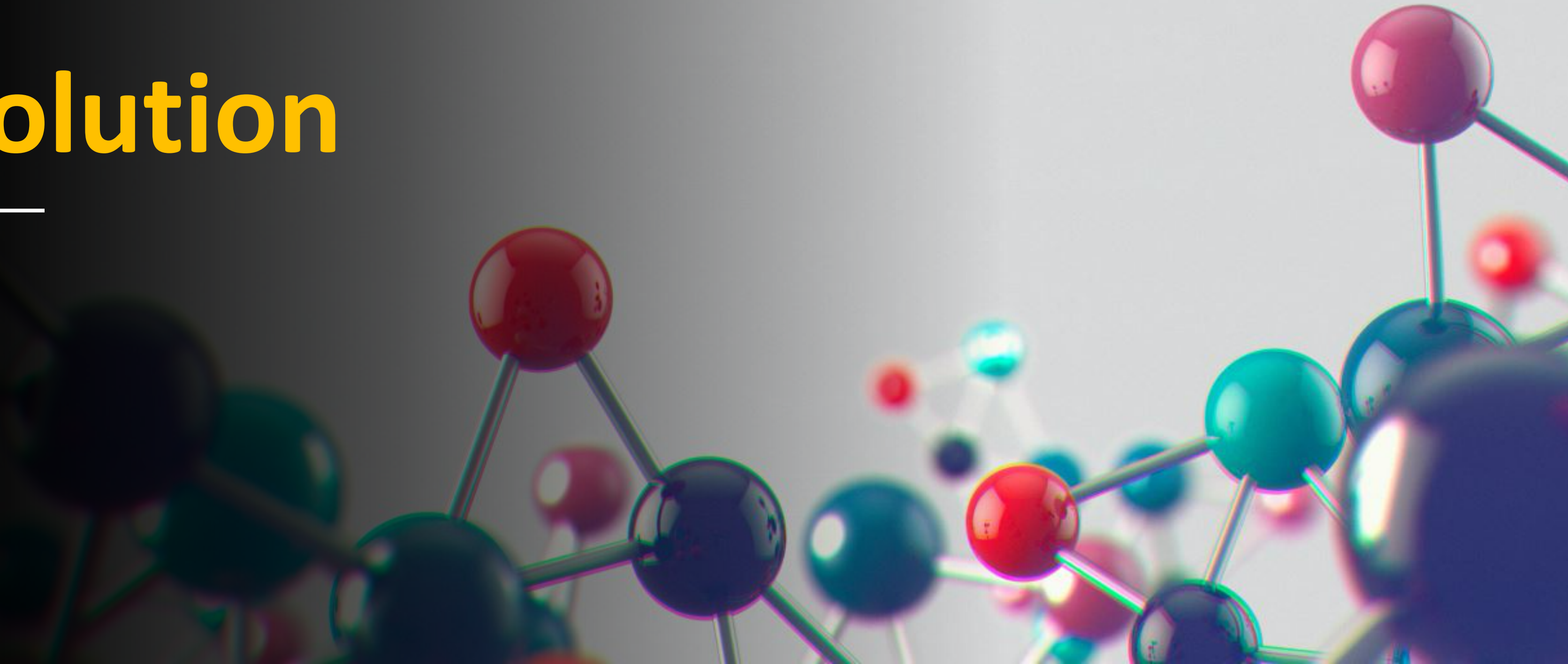
A2 Physical Chemistry

Enthalpies of Solution and Hydration

ChemistryTuition.Net



Definitions

The **Standard Enthalpy of Solution** is the enthalpy change which takes place when one mole of solute is completely dissolved in a solvent to form a solution of concentration 1 mol dm^{-3}, measured under standard conditions.

The **Standard Enthalpy of Hydration** is the enthalpy change when 1 mole of gaseous ions become hydrated (surrounded by water molecules), measured under standard conditions.

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

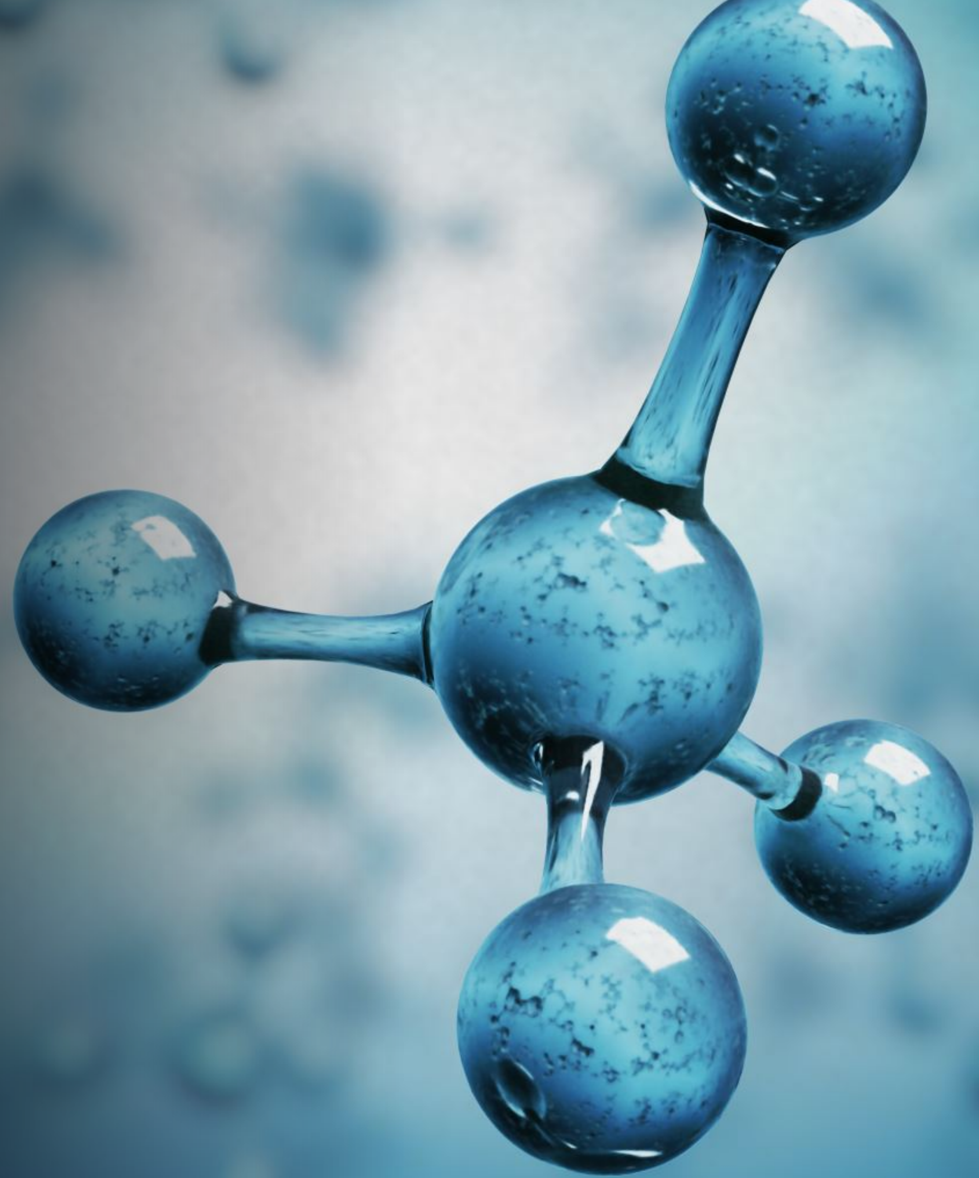
The **Standard Enthalpy of Solution** is the enthalpy change which takes place when one mole of solute is completely dissolved in a solvent to form a solution of concentration 1 mol dm^{-3}, measured under standard conditions.

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

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

The **Standard Enthalpy of Hydration** is the enthalpy change when 1 mole of gaseous ions become hydrated (surrounded by water molecules), measured under standard conditions.

Standard Enthalpy of Solution and Hydration Calculations



Standard Enthalpy of Hydration $Na^+_{(g)}$

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

Standard Enthalpy of Hydration $Cl^-_{(g)}$

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

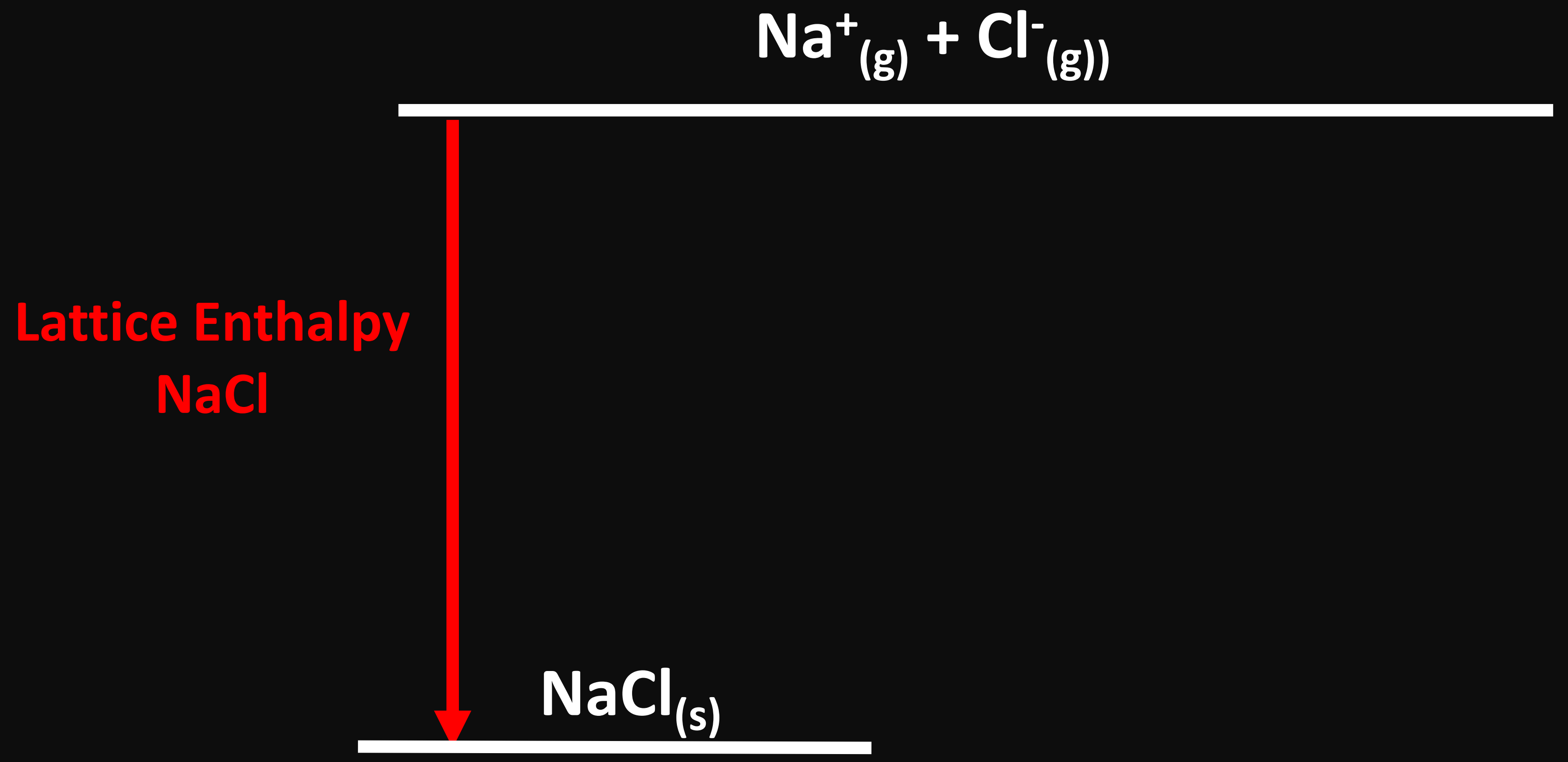
Standard Enthalpy of Solution

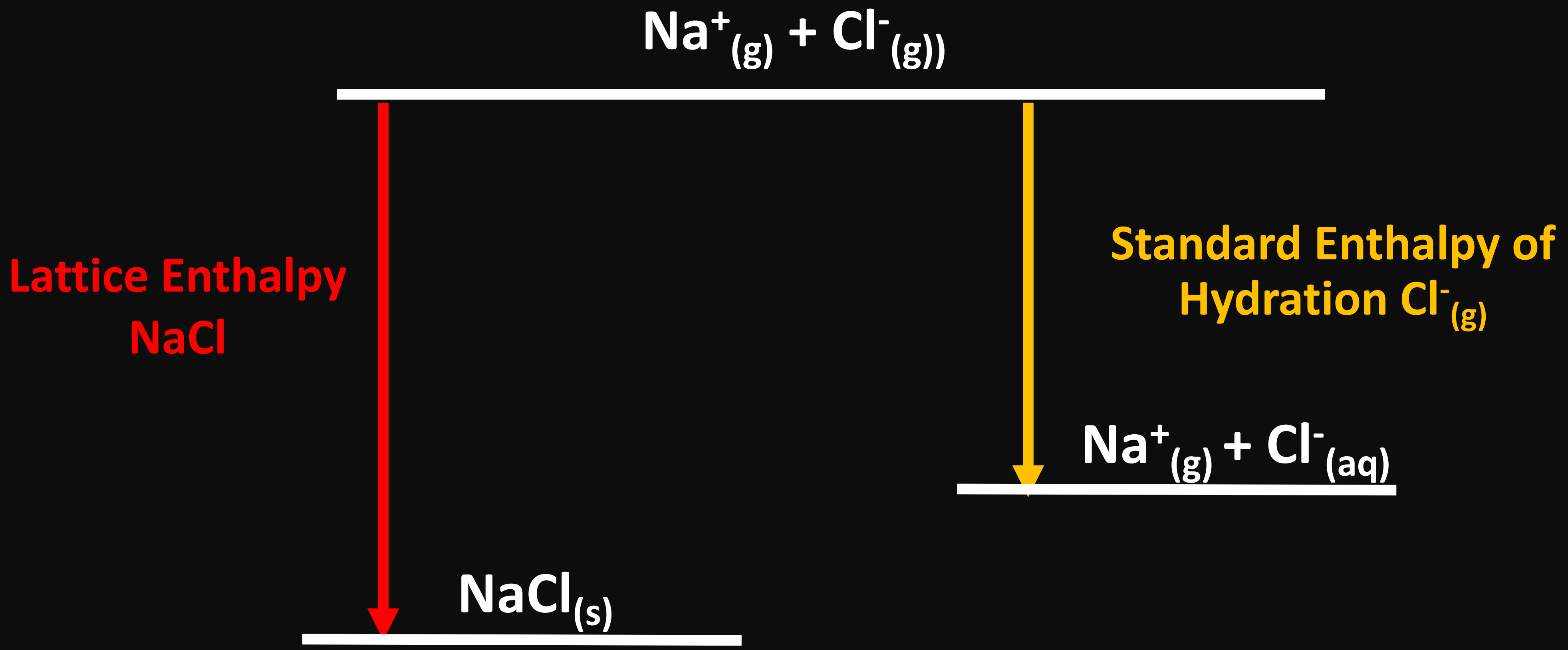
$NaCl_{(s)} + (aq) \rightarrow Na^+_{(aq)} + Cl^-_{(aq)}$

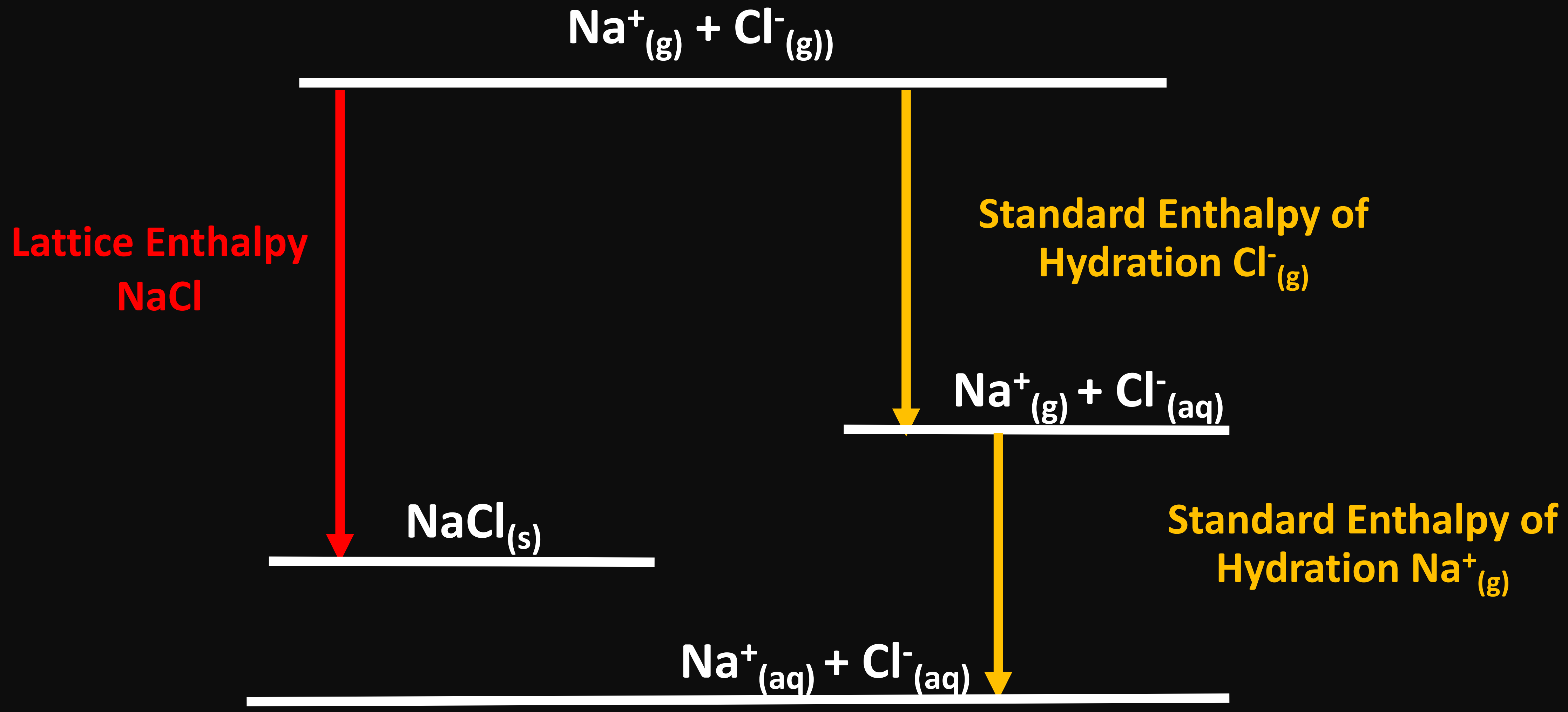
Lattice Enthalpy of $NaCl_{(s)}$

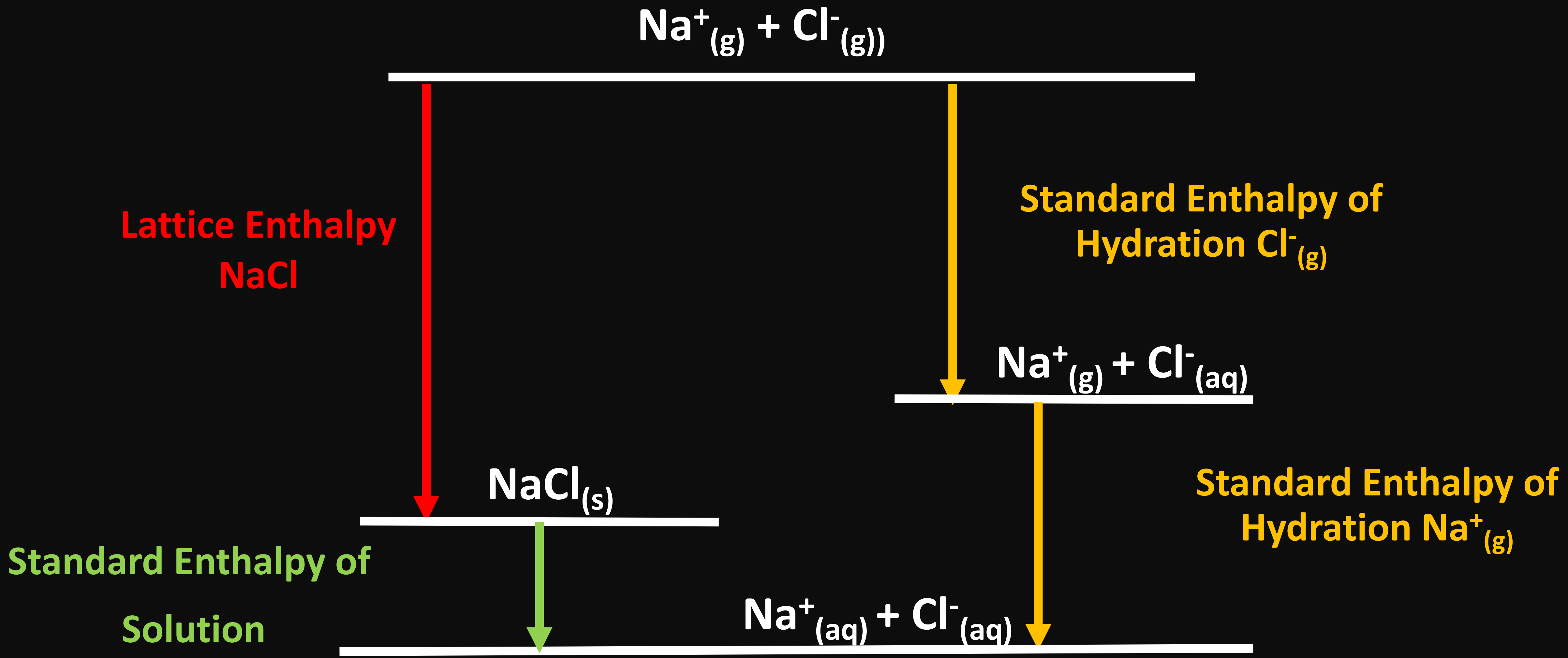
$Na^+_{(g)} + Cl^-_{(g)} \rightarrow NaCl_{(s)}$

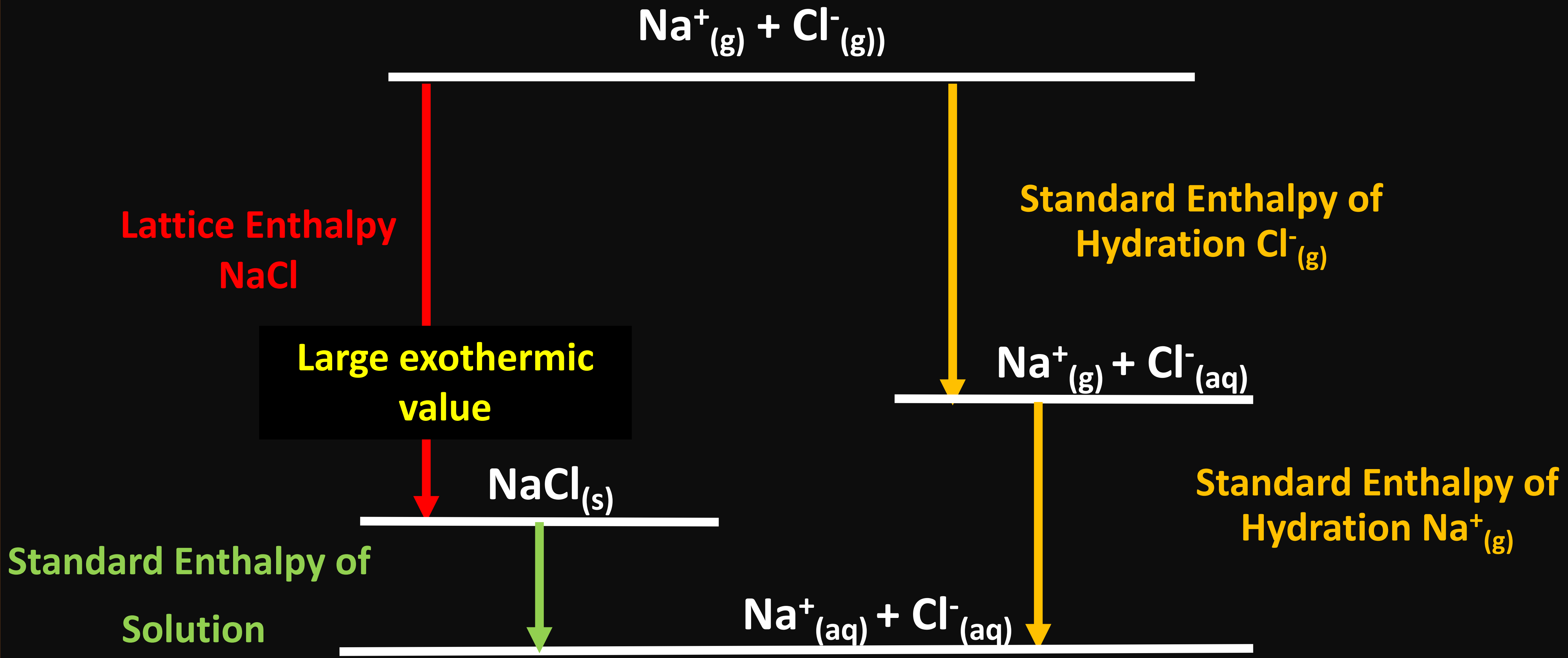
$Na^+_{(g)} + Cl^-_{(g)}$)

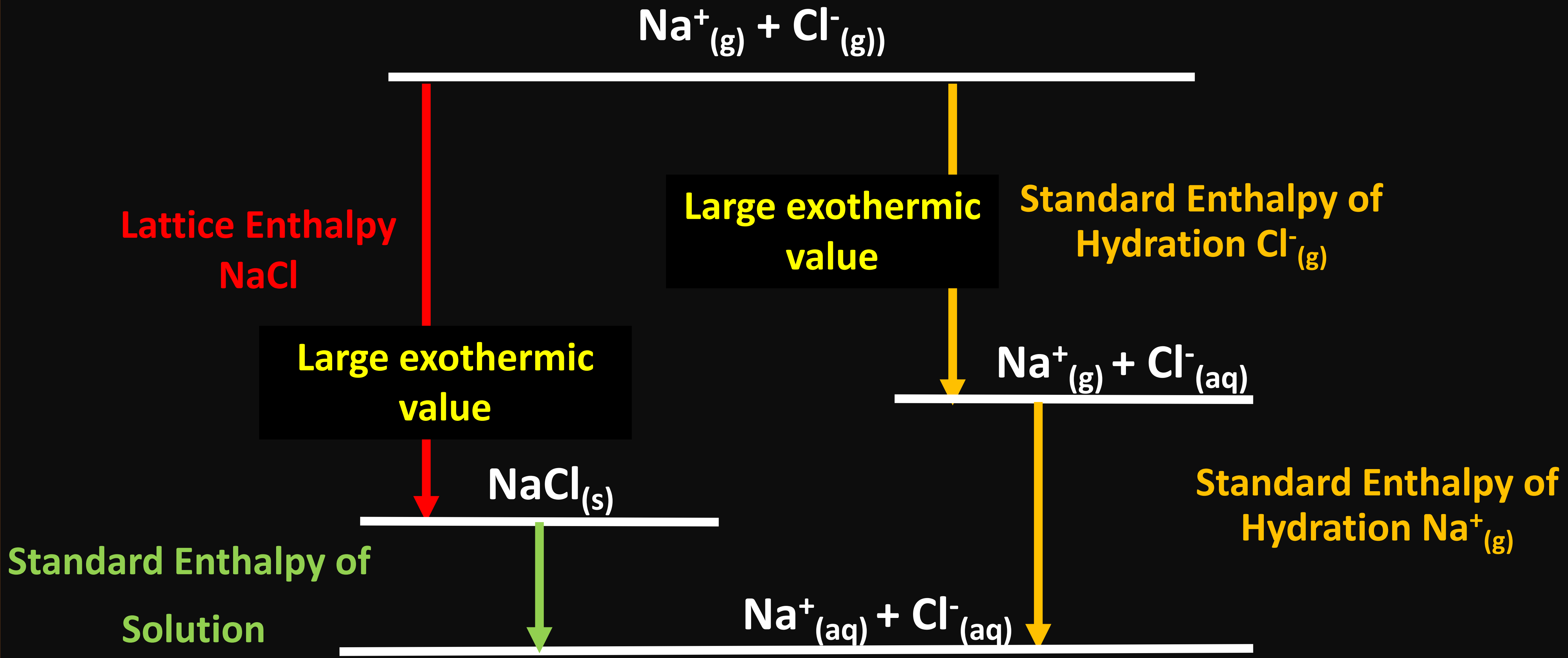

$Na^+_{(g)} + Cl^-_{(g))}$
Lattice Enthalpy
NaCl
$NaCl_{(s)}$


$Na^{+}_{(g)} + Cl^{-}_{(g))}$
Lattice Enthalpy
NaCl
$NaCl_{(s)}$
Standard Enthalpy of
Hydration $Cl^{-}_{(g)}$
$Na^{+}_{(g)} + Cl^{-}_{(aq)}$

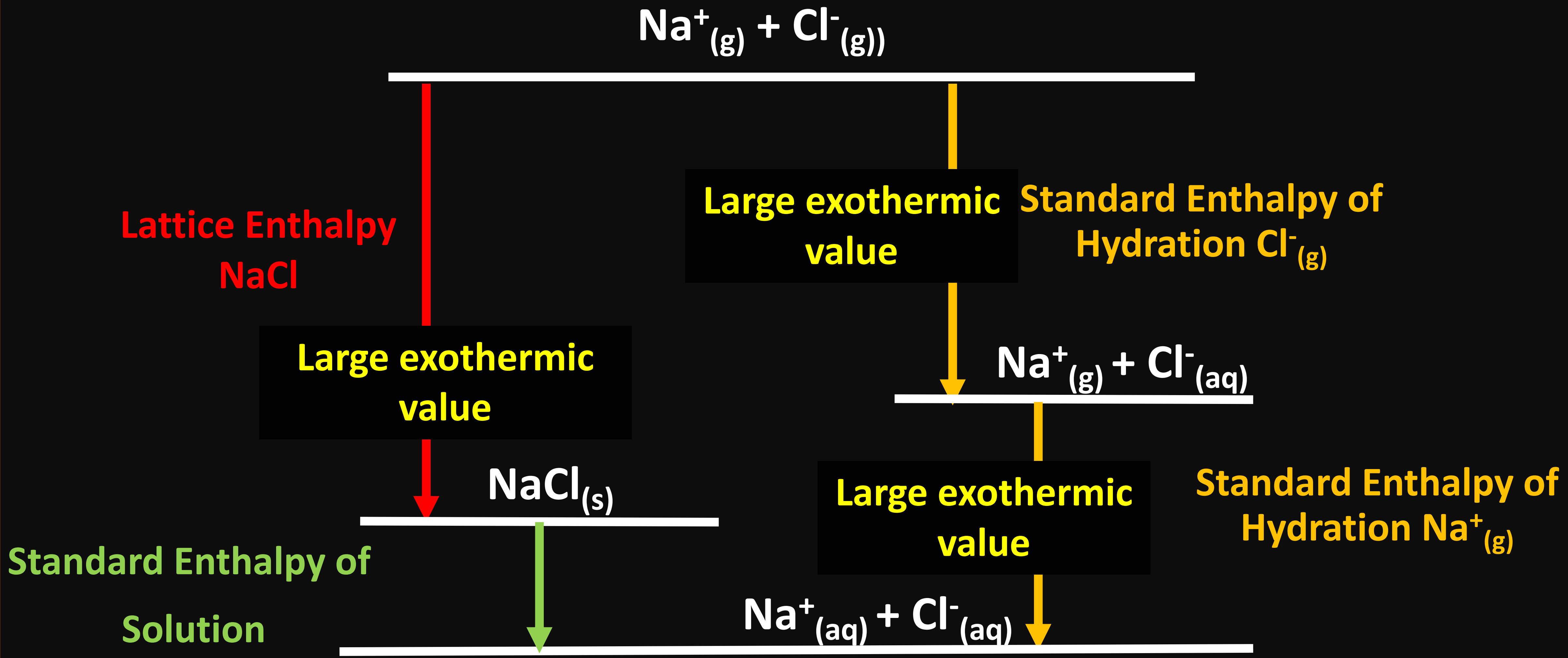

$Na^+_{(g)} + Cl^-_{(g))}$
Lattice Enthalpy NaCl
Standard Enthalpy of Hydration $Cl^-_{(g)}$
$Na^+_{(g)} + Cl^-_{(aq)}$
$NaCl_{(s)}$
Standard Enthalpy of Hydration $Na^+_{(g)}$
$Na^+_{(aq)} + Cl^-_{(aq)}$

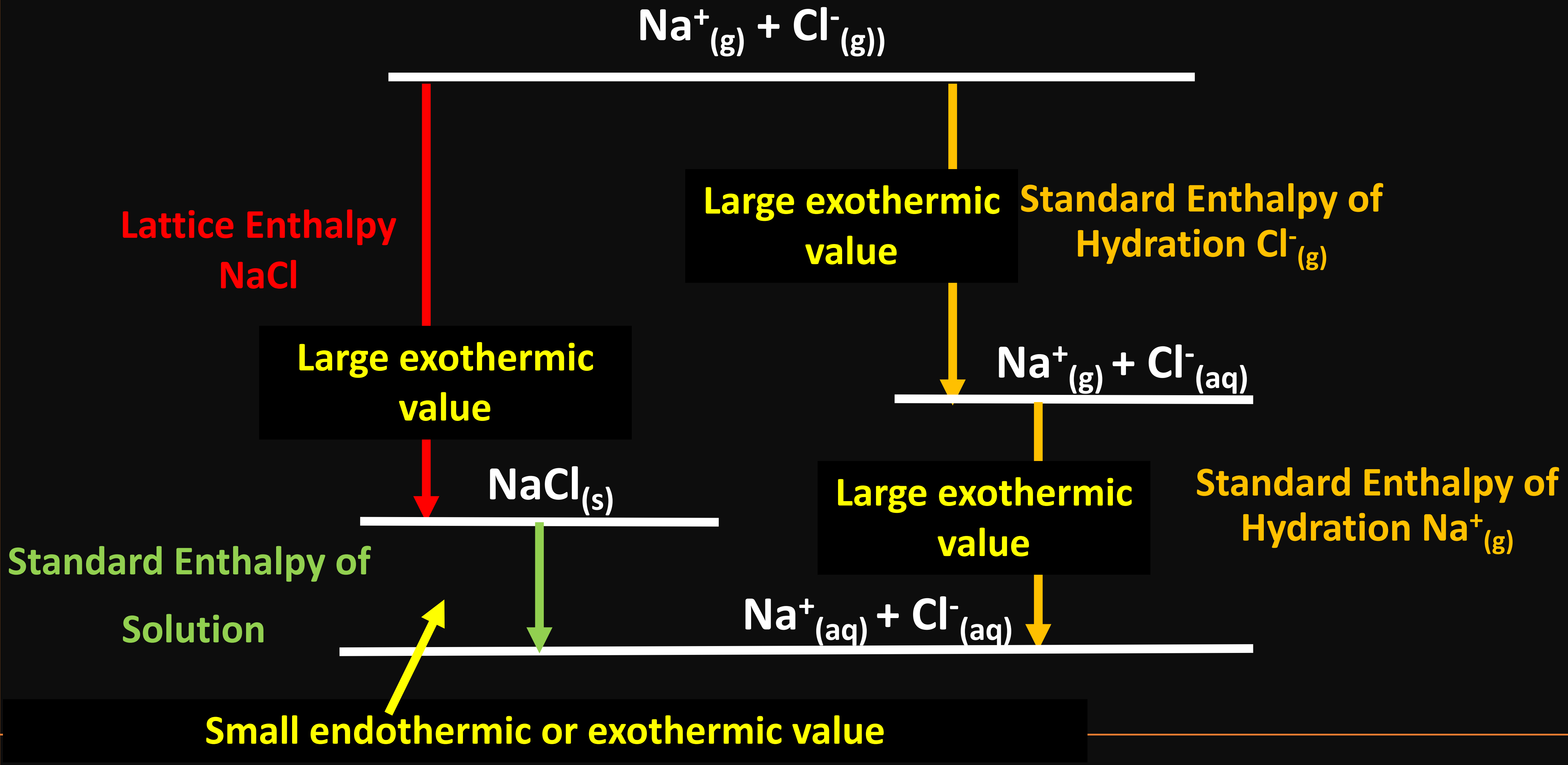

$Na^+_{(g)} + Cl^-_{(g))}$
Lattice Enthalpy
NaCl
Standard Enthalpy of
Hydration $Cl^-_{(g)}$
$Na^+_{(g)} + Cl^-_{(aq)}$
$NaCl_{(s)}$
Standard Enthalpy of
Hydration $Na^+_{(g)}$
Standard Enthalpy of
Solution
$Na^+_{(aq)} + Cl^-_{(aq)}$


$Na^+_{(g)} + Cl^-_{(g))}$
Standard Enthalpy of Hydration $Cl^-_{(g)}$
Lattice Enthalpy NaCl
Large exothermic value
$Na^+_{(g)} + Cl^-_{(aq)}$
$NaCl_{(s)}$
Standard Enthalpy of Hydration $Na^+_{(g)}$
Standard Enthalpy of Solution
$Na^+_{(aq)} + Cl^-_{(aq)}$



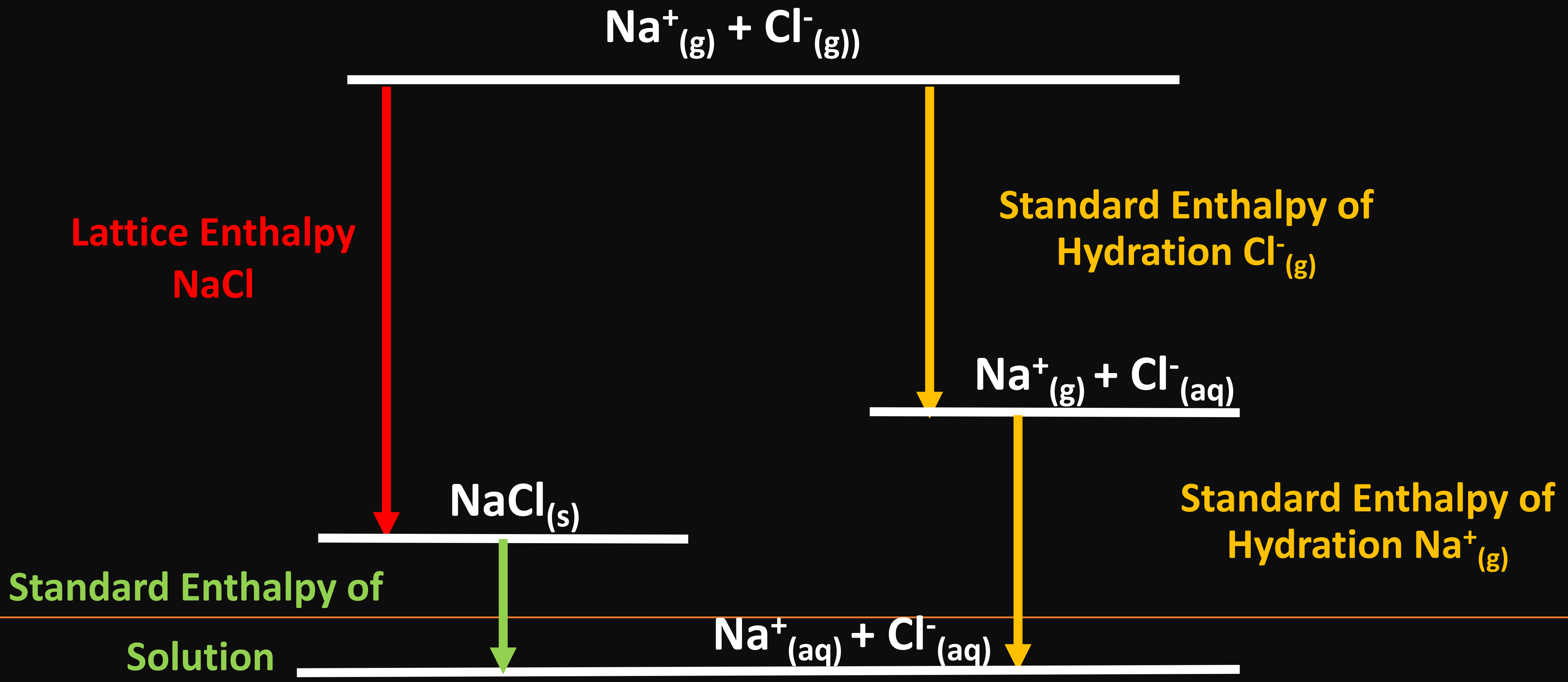
$Na^+_{(g)} + Cl^-_{(g))}$
Lattice Enthalpy NaCl
Large exothermic value
Large exothermic value
Standard Enthalpy of Hydration $Cl^-_{(g)}$
$Na^+_{(g)} + Cl^-_{(aq)}$
$NaCl_{(s)}$
Standard Enthalpy of Hydration $Na^+_{(g)}$
Standard Enthalpy of Solution
$Na^+_{(aq)} + Cl^-_{(aq)}$


$Na^+_{(g)} + Cl^-_{(g))}$
Lattice Enthalpy NaCl
Large exothermic value
Standard Enthalpy of Hydration $Cl^-_{(g)}$
Large exothermic value
$Na^+_{(g)} + Cl^-_{(aq)}$
Large exothermic value
$NaCl_{(s)}$
Standard Enthalpy of Hydration $Na^+_{(g)}$
Standard Enthalpy of Solution
$Na^+_{(aq)} + Cl^-_{(aq)}$

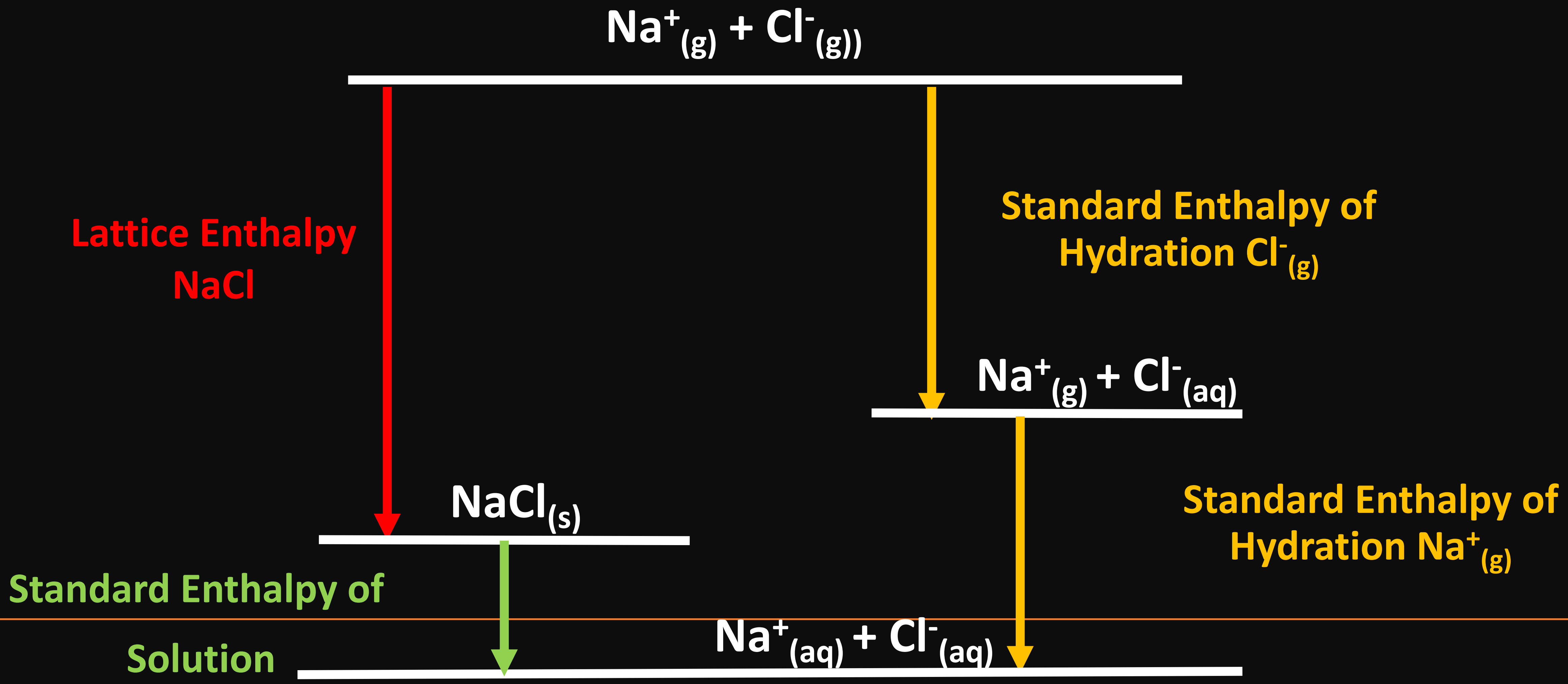


$Na^+_{(g)} + Cl^-_{(g))}$
Lattice Enthalpy NaCl
Large exothermic value
Large exothermic value
Standard Enthalpy of Hydration $Cl^-_{(g)}$
$Na^+_{(g)} + Cl^-_{(aq)}$
$NaCl_{(s)}$
Large exothermic value
Standard Enthalpy of Hydration $Na^+_{(g)}$
Standard Enthalpy of Solution
$Na^+_{(aq)} + Cl^-_{(aq)}$
Small endothermic or exothermic value

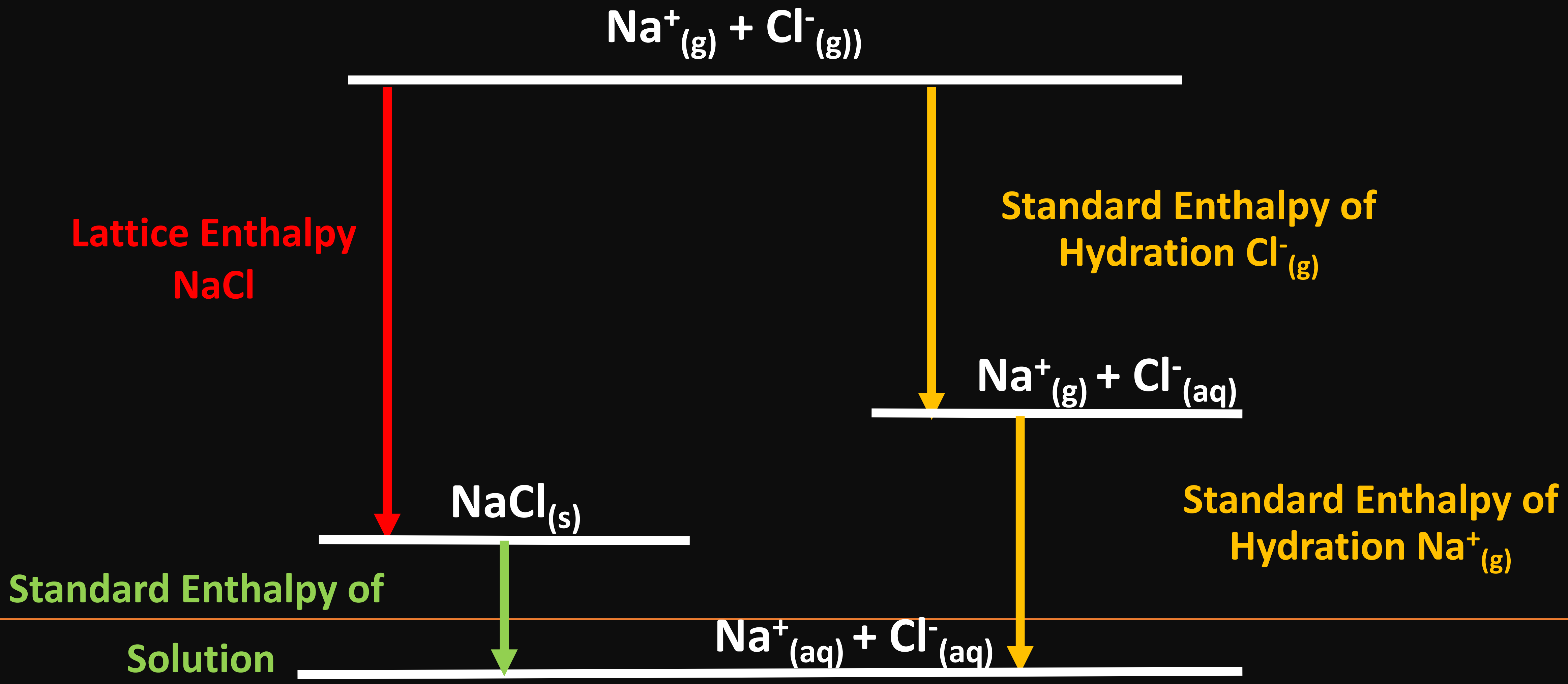
Standard Enthalpy of Solution depends upon the relative values of Lattice Enthalpy vs hydration enthalpies.



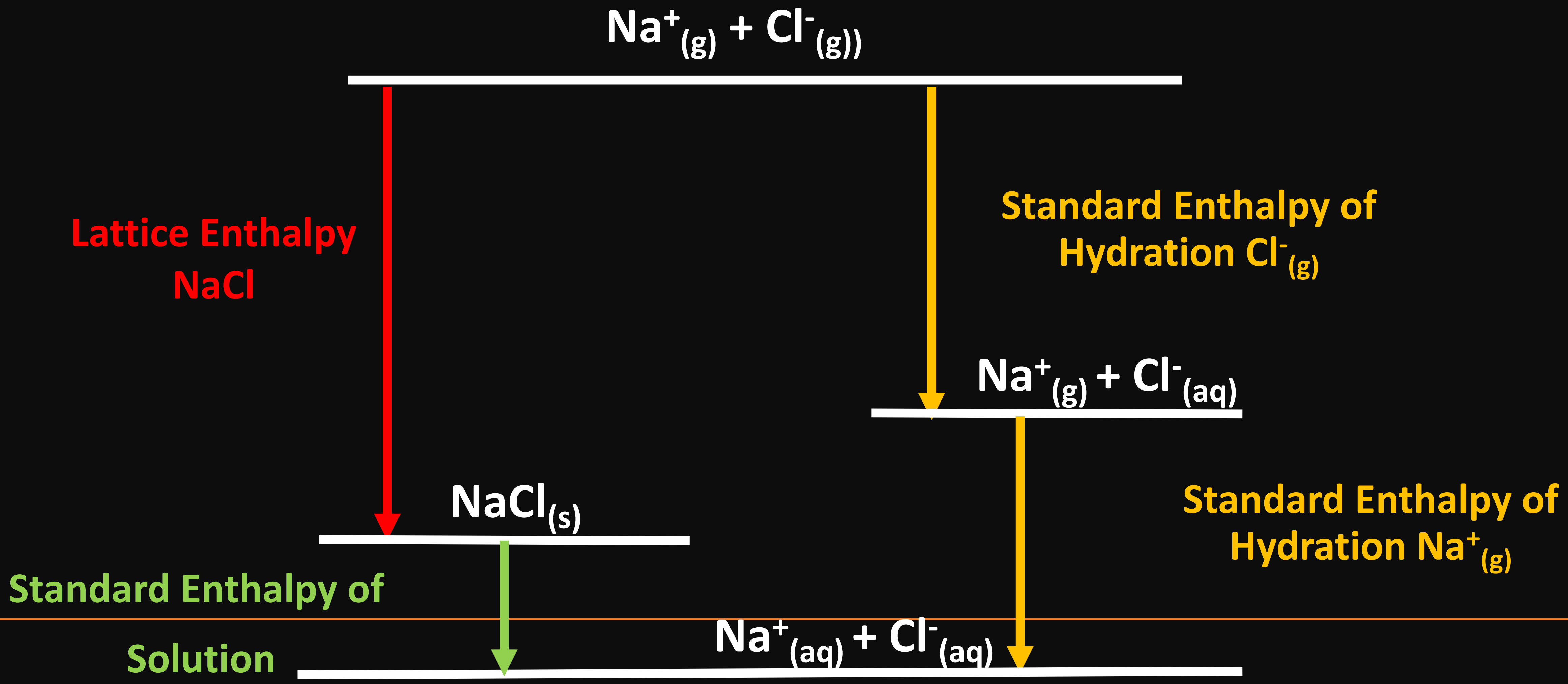
Enthalpy of Solution = Hydration Enthalpies - Lattice Enthalpy



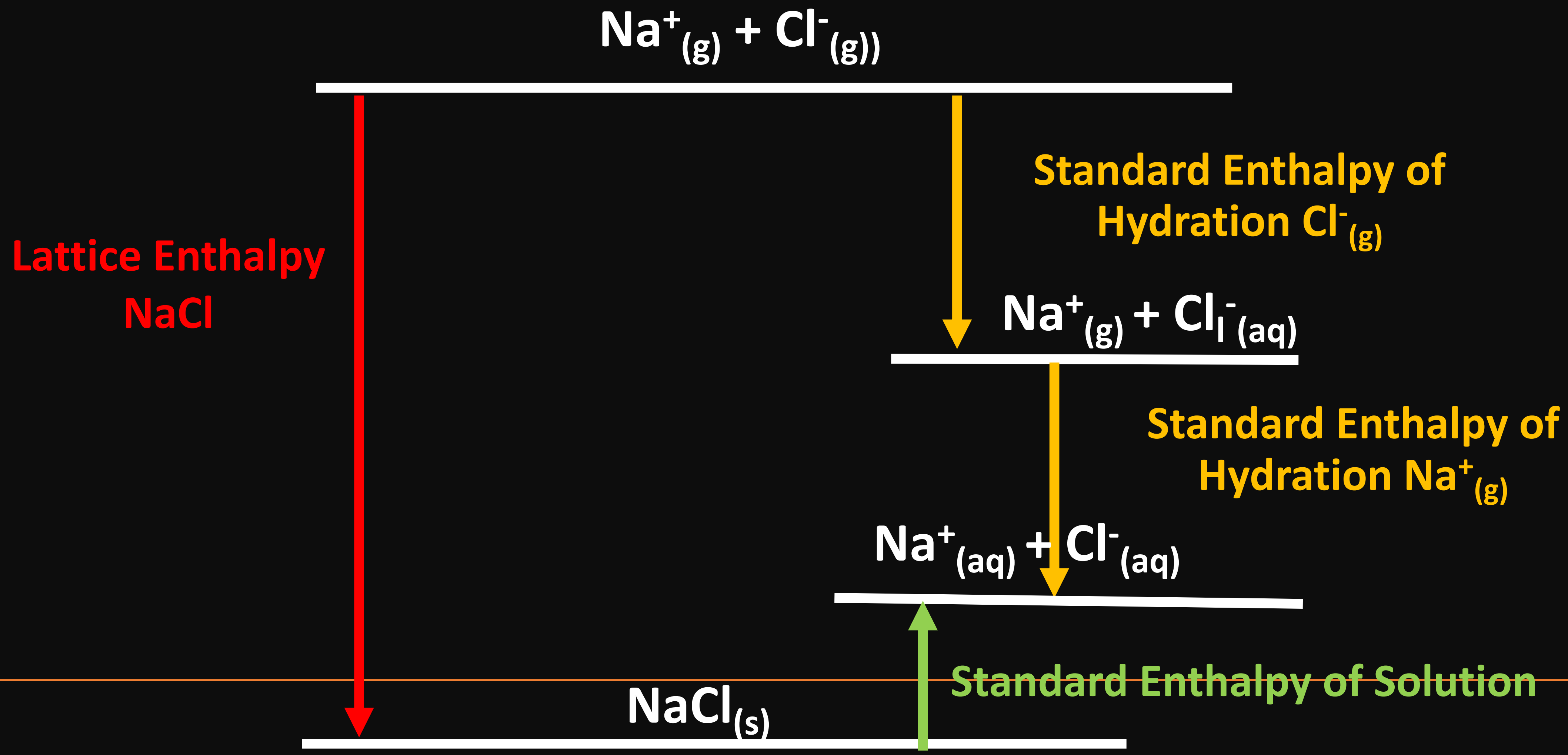
If the two hydration enthalpies are more exothermic than the lattice enthalpy, then the enthalpy of solution will be exothermic



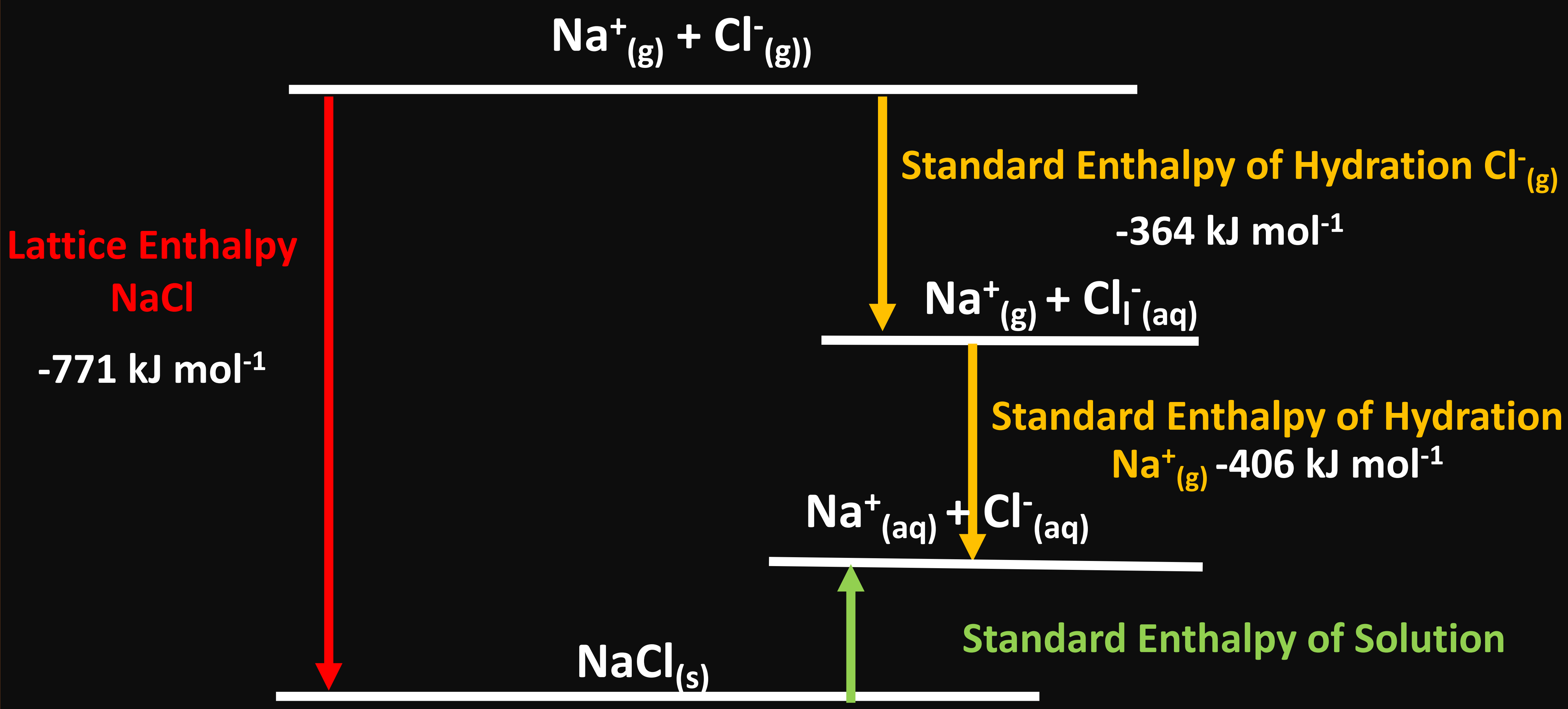
If the lattice enthalpy is more exothermic than the two hydration enthalpies



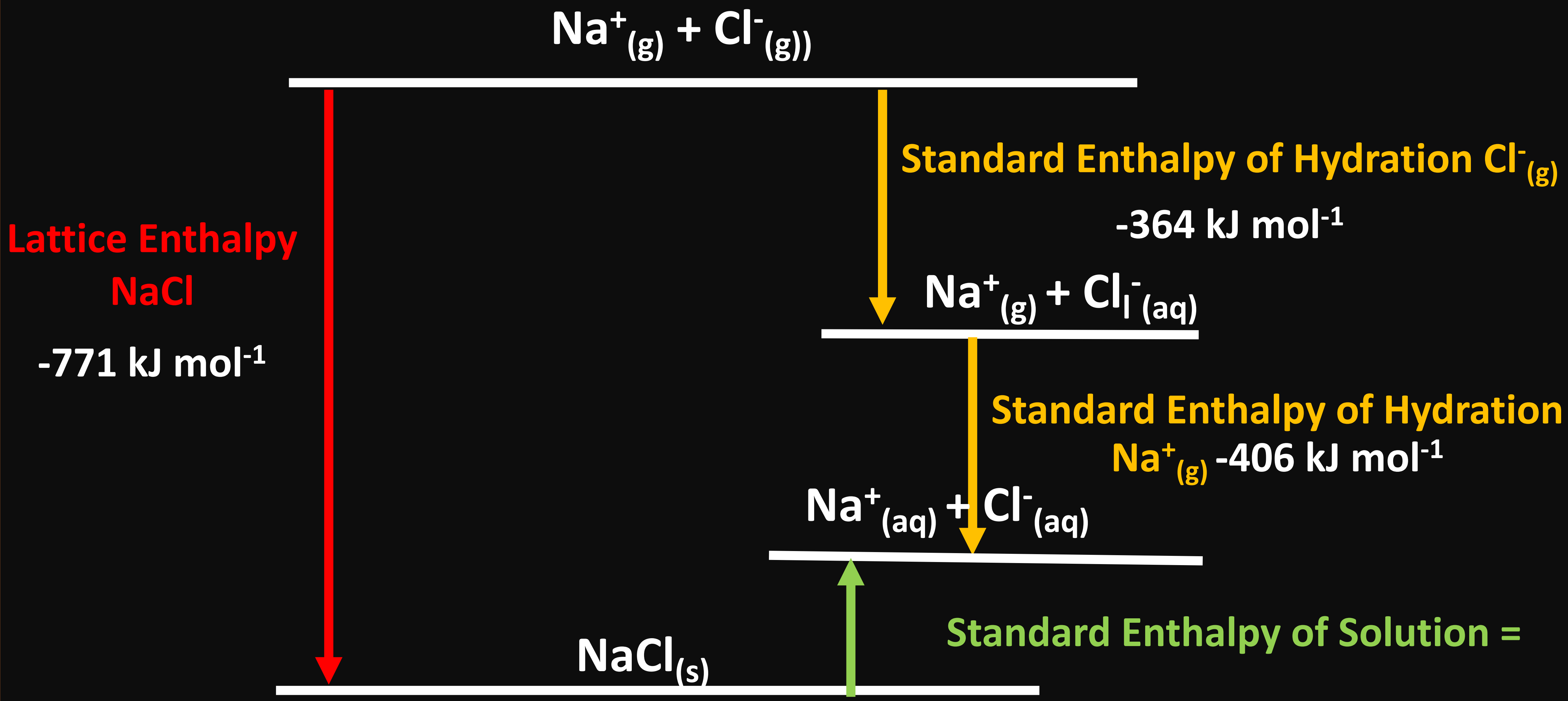
If the lattice enthalpy are more exothermic than the two hydration enthalpies, then the enthalpy of solution will be endothermic.



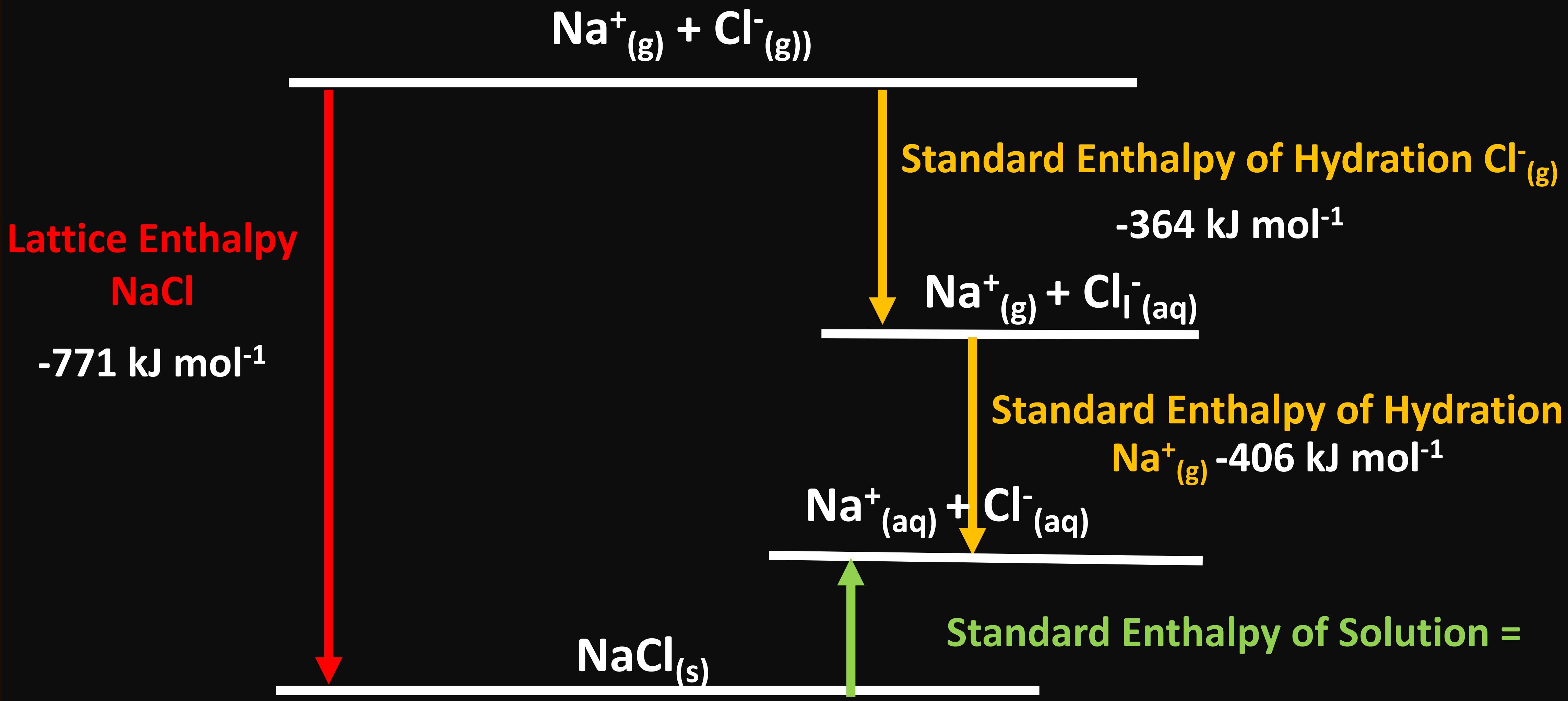
This is in fact the case for NaCl



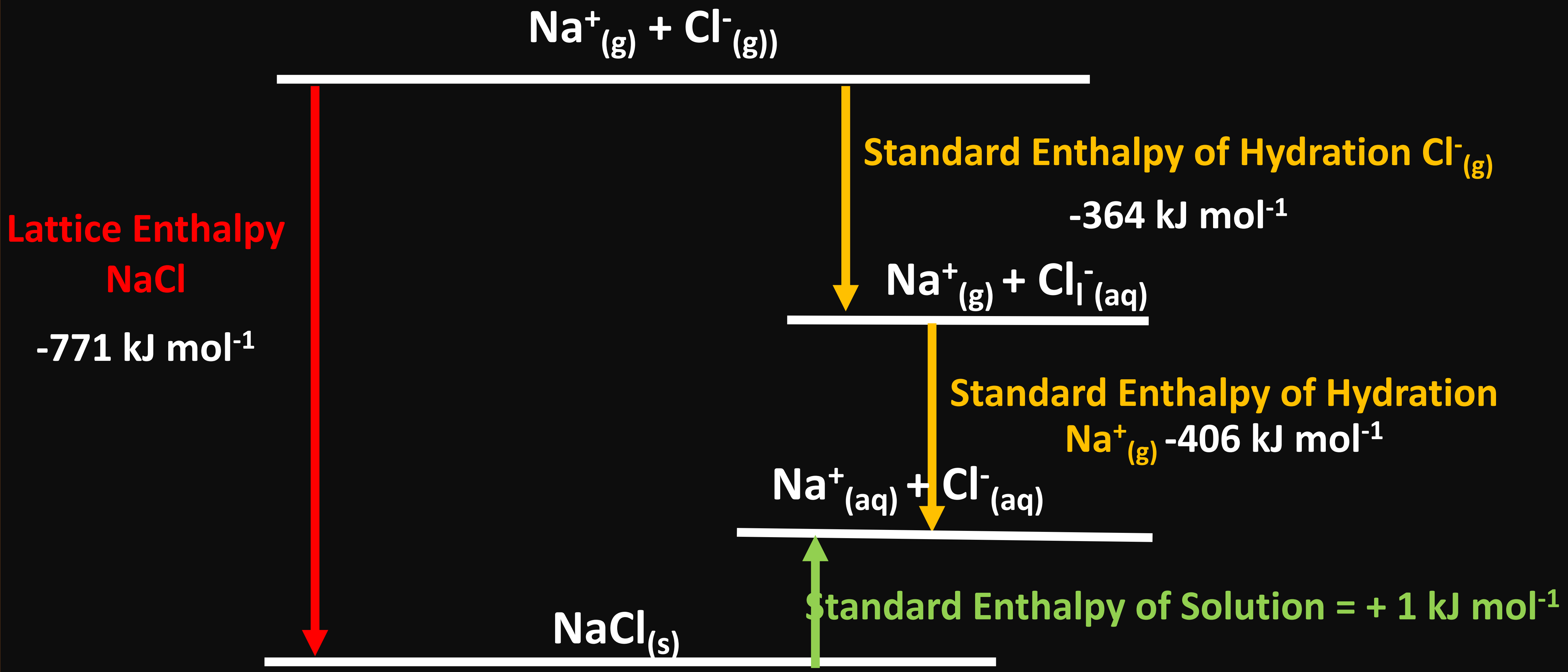
Solution Enthalpy = (-364 – 406) – (-771)



Solution Enthalpy = -770 + 771



Solution Enthalpy = +1 kJ mol^{-1}



Enthalpy of Solution = Hydration Enthalpies - Lattice Enthalpy

Determined by strength of the attraction between ions and water molecules

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Determined by strength of the attraction between the oppositely charged ions

Determined by strength of the attraction between ions and water molecules

Enthalpy of Solution = Hydration Enthalpies - Lattice Enthalpy

Determined by strength of the attraction between the oppositely charged ions

If the strength of attractions between the ions and water > ionic bond strength

Enthalpy of Solution is exothermic

Determined by strength of the attraction between ions and water molecules

Enthalpy of Solution = Hydration Enthalpies - Lattice Enthalpy

Determined by strength of the attraction between the oppositely charged ions

If ionic bond strength > the strength of attractions between the ions and water

Enthalpy of Solution is endothermic

Factors affecting Lattice Enthalpy and Hydration Energy

Stronger lattices (with more exothermic lattice enthalpies) occur with

- Smaller ions with smaller ionic radius which can pack more closely together giving stronger attractive forces

- Ions with a double charge (2+ or 2-) have much stronger attractive forces than single charges

As the ionic charge increases and ionic radius decreases, the enthalpy change of hydration becomes more exothermic because:

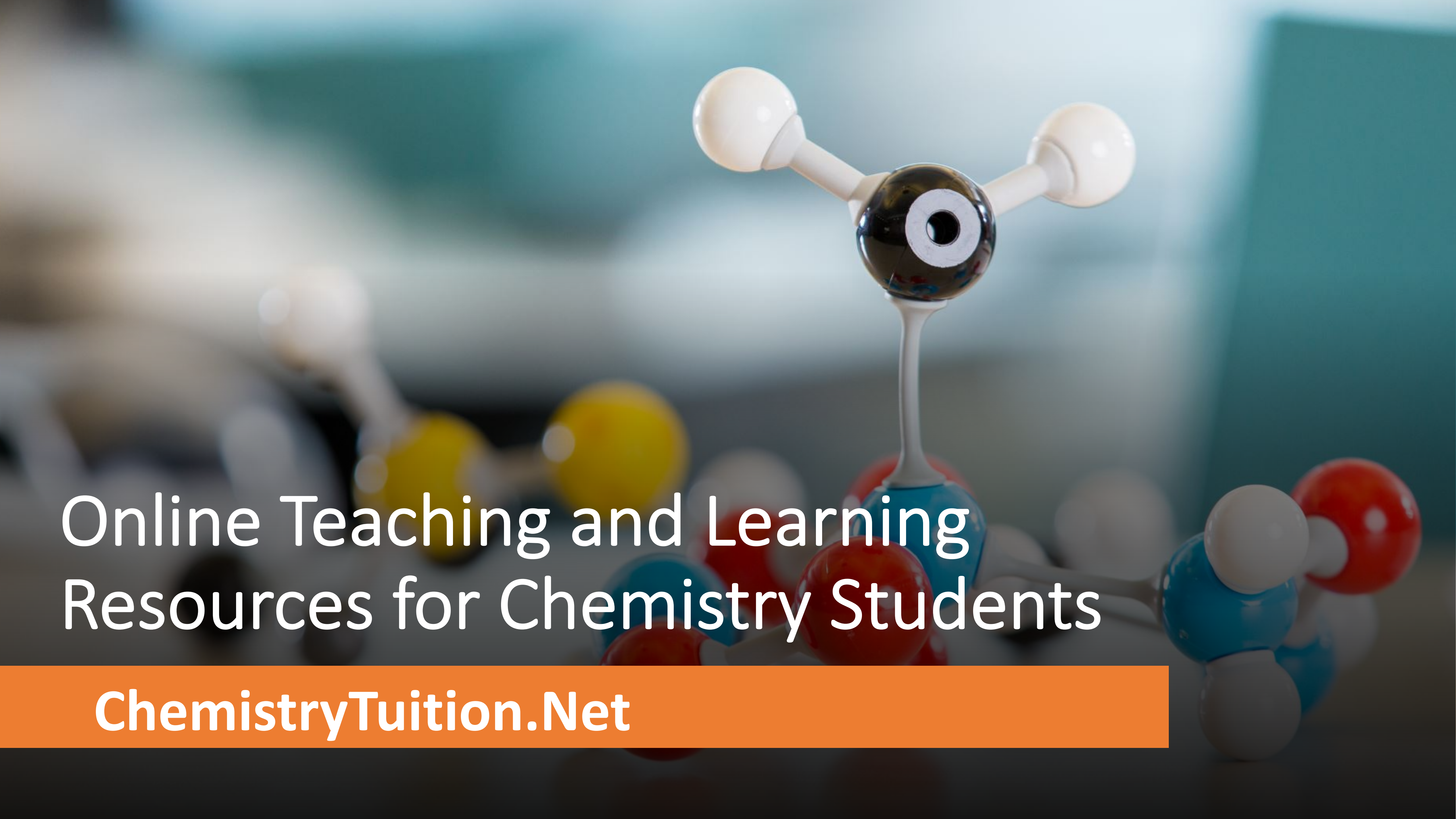
- Increasing the ionic charge and decreasing the ionic radius both INCREASE the charge density of the ion.
- As the ion's charge density increases, this leads to a stronger attraction for the water molecules

Determined by strength of the attraction between ions and water molecules

Enthalpy of Solution = Hydration Enthalpies - Lattice Enthalpy

Determined by strength of the attraction between the oppositely charged ions

Since both the hydration and lattice enthalpies become more exothermic as the ions become smaller and higher charged, it is difficult to predict whether the solution enthalpy will be endothermic or exothermic.



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