

Born-Haber Cycles part 1

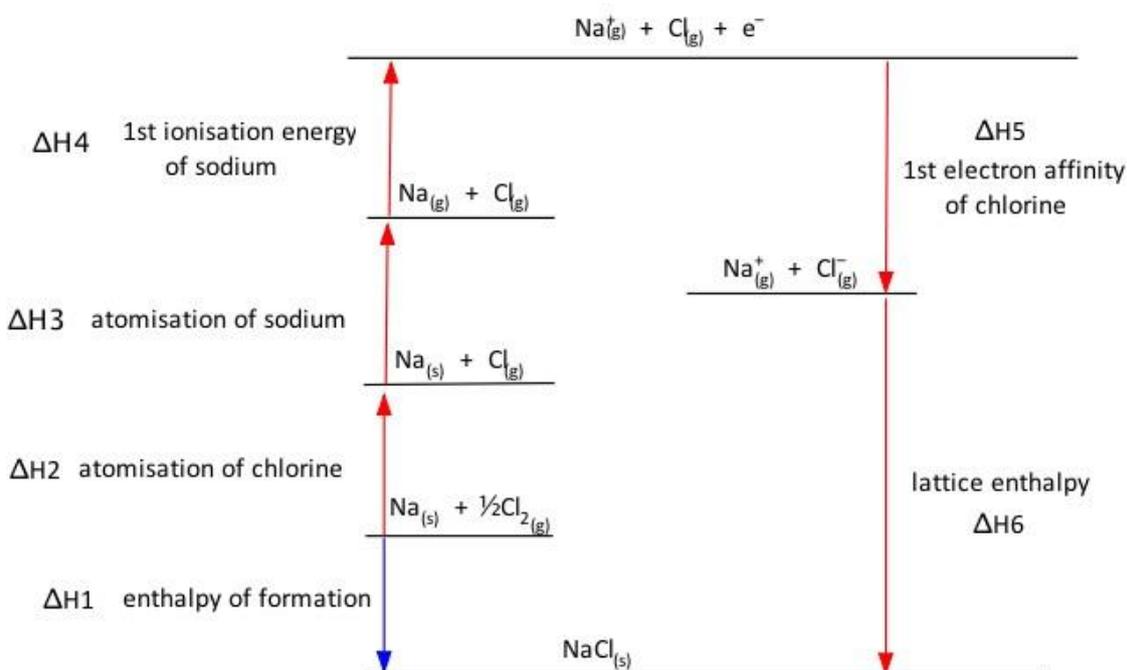
A Born-Haber cycle can be thought of as an extended [Hess cycle](#). But most students find Born-Haber much easier 😊. The good thing is that a Born-Haber cycle is consistent; every cycle has the same **five enthalpy** changes so you don't need to worry if the arrows point up or down. Result!

Born-Haber cycles look only at **ionic** compounds.

The questions are normally in three parts:

- identifying enthalpy changes/amending a cycle
 - calculating an enthalpy change
 - theory questions i.e. why two ionic compounds have different experimental and theoretical lattice enthalpy values
- ✓ the main thing to be able to do is to draw a Born-Haber cycle or at least be able to amend a partially complete cycle.

Below is the Born-Haber cycle for the formation of NaCl. We will discuss it in detail below, this is just to introduce what the cycle looks like:



- ✓ the direction of the arrows in the cycle tell you if the enthalpy change is exothermic or endothermic i.e. if the arrow points up it is endothermic and down is exothermic.

Enthalpy Changes

There are 5 types of enthalpy change in any Born Haber cycle: **ionisation, atomisation, electron affinity, lattice enthalpy and formation enthalpy.**

- ✓ the specification says **only ionisation, lattice and formation** definitions are required. But you still need to be able to recognise and write equations for atomisation and electron affinity.

Ionisation (endothermic)

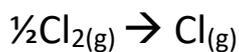
You should know this one already!

“Enthalpy change when one mole of gaseous ions are formed from 1 mole of gaseous atoms”



Atomisation (endothermic)

You probably haven't seen this one before. This involves the formation of **one atom**. Remember that...just one atom. The clue is in the name! It looks a bit strange seeing chlorine written as Cl rather than Cl₂ but that's what happens.



Electron Affinity

You can think of this as the **opposite of ionisation** energy. It is only for **non-metals**. We are **adding** an electron to a non-metal to form a negative ion.



If you are forming a 2- ion e.g. O²⁻, you need to do a **second** electron affinity.

The first electron affinity is **exothermic** but the second electron affinity is **endothermic**.

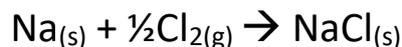
When adding a second electron, it is being added to a **negative ion**. Can you see the problem with that? Electron repulsion! Therefore it takes energy to add that second electron resulting in an endothermic process.

With the first electron affinity, the electron is being added to a **neutral** species so there is no problem.

Formation Enthalpy (exothermic)

You should know this one from Hess's law.

"Enthalpy change when one mole of a compound is formed from the elements in their standard states"



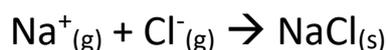
Lattice Enthalpy (exothermic)

This goes back to ionic compounds that exist as ionic lattices.

Lattice enthalpy is a measure of ionic bond strength.

A positive metal ion and a negative non-metal ion to form a lattice.

"Enthalpy change when one mole of a solid compound is formed from the gaseous ions under standard conditions"



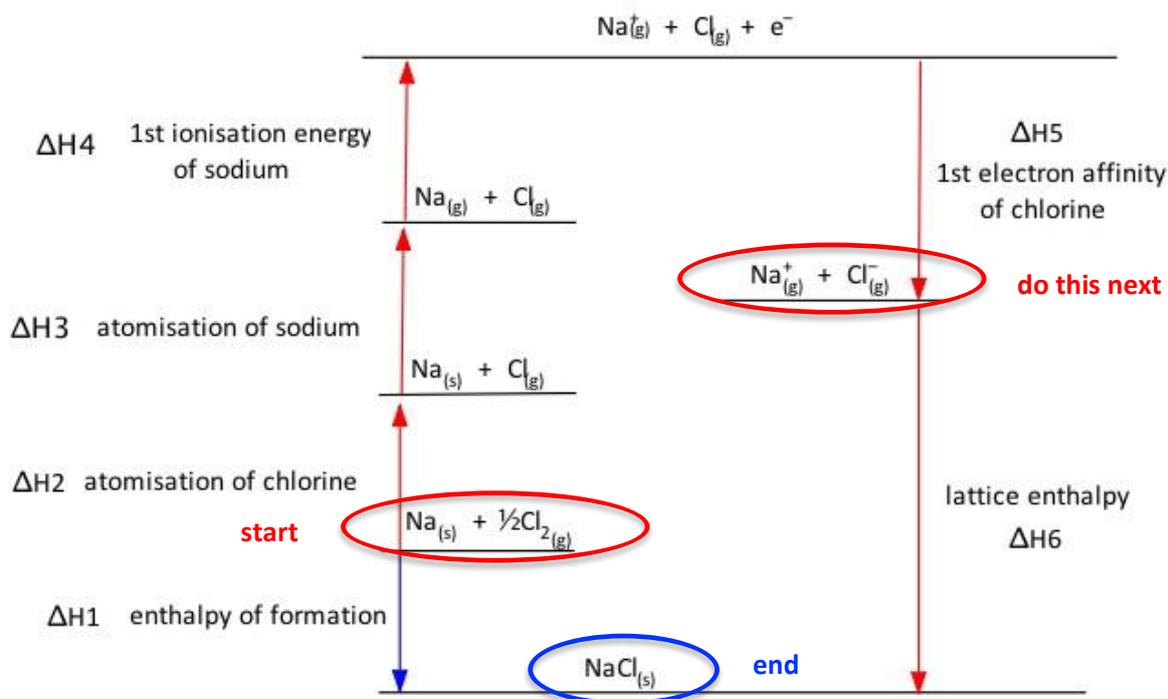
- ✓ students often get the enthalpy of formation and lattice enthalpy mixed up. You need to get these definitions clear as soon as possible, even if they are just in your own words.

The Cycle

I am going to show you how to draw a Born-Haber cycle. You are unlikely to have to draw a complete cycle but if you can, you will be an absolute ninja master at this and won't have any problem with any of the questions.

Think of the "goal" of the cycle as forming the ionic compound e.g. NaCl. And we have **two routes** to get to NaCl (like Hess's law).

- We can get there via the short route: enthalpy of formation: $\text{Na}_{(s)} + \frac{1}{2}\text{Cl}_{2(g)} \rightarrow \text{NaCl}_{(s)}$ (shown by the blue arrow in the diagram below).
- Or...we can go via looonnggg route which involves all the other enthalpy changes ending with lattice enthalpy: $\text{Na}^+_{(g)} + \text{Cl}^-_{(g)} \rightarrow \text{NaCl}_{(s)}$ (shown by the red arrows in the diagram below).



Start: the easiest place to start is from $\text{Na}_{(s)} + \frac{1}{2}\text{Cl}_{2(g)}$ (the elements in their standard states).

End: the “end” of the cycle is the formation of the ionic compound NaCl. So start by drawing the enthalpy of formation with a **downward arrow**.

The second thing I do is the lattice enthalpy, which also makes NaCl. So I need the **gaseous ions**: $\text{Na}_{(g)}^+$ and $\text{Cl}_{(g)}^-$. This now gives me a target. I know I need to go from the elements to the gaseous ions which involves atomisation, ionisation and electron affinity.

Left hand side of the diagram

After we have started the cycle as discussed above, on the left hand side we must **atomise** the metal and non-metal and also **ionise** the metal (arrows going upwards).

On this side, we have a bit of choice with the order of the steps. You can atomise the metal or non-metal first, it doesn't matter. You could even do the ionisation of the metal before atomisation of the non-metal. But of course, you must atomise metal before you can ionise it.

Right hand side of the diagram

You can't change the order on the right hand side. You must start with **electron affinity** of the non-metal with the arrow going **downwards**. And finally, the lattice enthalpy also pointing downwards.

✓ A point to note in all these steps: **all the equations must balance**. If you do an ionisation step, for example: $\text{Na}_{(g)} + \text{Cl}_{(g)} + e^- \dots$ you must have the electron to balance the equation.

Calculations

The calculations are really easy but students get this wrong a lot! They could ask you to work out a value for *any* of the steps in the cycle, although frequently it is the lattice enthalpy.

The secret to this is consistency and not trying to rearrange too quickly.

Using Hess's law theory, we add up all the arrows going in the **same direction** i.e. all the red arrows are equal to the blue arrow, which gives us:

$$\Delta H_2 + \Delta H_3 + \Delta H_4 + \Delta H_5 + \Delta H_6 = \Delta H_1$$

You must write it out as above every single time...no matter what you are trying to calculate. I see students try to rearrange too quickly or others do something really bizarre like reversing the sign of all the enthalpy changes. From experience this just introduces problems. Keep it simple!

You then just put the numbers into above equation **before** you rearrange it. For example, to calculate ΔH_6 using the following values (all kJ mol^{-1}) for the NaCl cycle and putting the numbers into the above equation:

$$\Delta H_1 = -411$$

$$\Delta H_2 = +122$$

$$\Delta H_3 = +107$$

$$\Delta H_4 = +496$$

$$\Delta H_5 = -349$$

$$-411 = 122 + 107 + 496 - 349 + \Delta H_6$$

$$-411 = +376 + \Delta H_6$$

$$\Delta H_6 = -787$$

- ✓ The rearrangement takes place at the *end*. It is a simple subtraction to get the answer. I haven't changed the signs on any of the values either.