

Answer on Question #55845 - Chemistry - General chemistry

Question:

What would be the three equations that cancel correctly for Hess Law for the following reaction for Magnesium oxide that you would have $\text{Mg(s)} + 1/2 \text{O}_2(\text{g}) = \text{MgO(s)}$? Would the change in heat be a negative or a positive? what would be the final net reaction? I am confused would you then have four equations

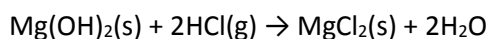
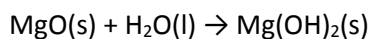
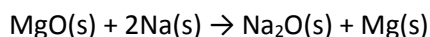
Answer:

For MgO the enthalpy formation is -601.83 kJ/mol . Since the corresponding values for Mg(s) and O₂(g) are zero the heat for this reaction is defined:

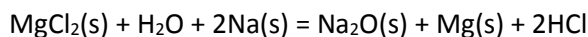
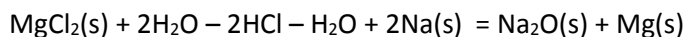
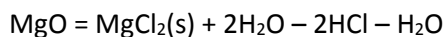
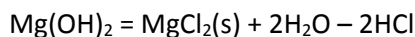
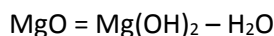
$$\Delta H = \Delta H(\text{MgO}) - \Delta H(\text{O}_2) - \Delta H(\text{Mg}) = -601.83 \text{ kJ/mol} - 0 - 0 = -601.83 \text{ kJ/mol}$$

This is exothermic process and that releases heat.

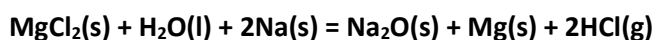
For instance, these three reactions are:



The final net reaction is found using the mentioned above reactions:



Thus, the final net reaction:



The enthalpy change for this process is determined:

Taking into account that:

$$\Delta H(\text{MgCl}_2)_{(\text{s})} = -641.8 \text{ kJ/mol}$$

$$\Delta H(\text{H}_2\text{O})_{(\text{s})} = -285.8 \text{ kJ/mol}$$

$$\Delta H(\text{Na}_2\text{O})_{(\text{s})} = -414.2 \text{ kJ/mol}$$

$$\Delta H(\text{HCl})_{\text{g}} = -92.30 \text{ kJ/mol}$$

The total heat equals:

$$\Delta H = \Delta H(\text{Na}_2\text{O}) + \Delta H(\text{Mg}) + 2 \Delta H(\text{HCl}) - \Delta H(\text{MgCl}_2) - \Delta H(\text{H}_2\text{O}) - 2 \Delta H(\text{Na}) = -414.2 + 0 + (-184.6) + 641.8 + 285.8 - 0 = +328.8 \text{ kJ/mol}$$