

Enthalpy

$$q_p = \Delta H = \Delta E + P\Delta V$$
$$\Delta H = H_{\text{final}} - H_{\text{initial}} = H_{\text{products}} - H_{\text{reactants}}$$
$$w = -P\Delta V$$

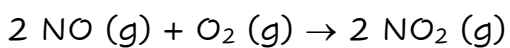
1. What conditions will the enthalpy change of a process or reaction be equal to the heat that is transferred into or out of the system?

If the process occurs under a constant pressure with only PV work. $\Delta H = q_p$

2. If a process is run under constant pressure and heat is released from the system, will the enthalpy of the system increase or decrease?

It will decrease because both q and ΔH are negative.

3. Consider the following balanced equation:



If the reaction were carried out in a constant volume container at constant temperature, would the amount of heat (absorbed or released) correspond to ΔH or ΔE ? Which quantity would be larger for this reaction?

At constant volume, ΔV is equal to zero and ΔE is equal to q_v . $H = E + PV$ and $\Delta H = \Delta E + \Delta(PV)$. An ideal gas at constant temperature and volume has $\Delta(PV) = V\Delta P = RT\Delta n$ where n is equal to the number of moles of gas. There are 3 moles of reactant gas and 2 moles of product gas and $\Delta n = 2 - 3 = -1$. $\Delta(PV)$ is negative. The negative $\Delta(PV)$ term means ΔH will be smaller than ΔE .

4. A gas is confined to a vessel under a constant pressure. The gas undergoes a chemical reaction and absorbs 785 J of heat from the surroundings. There are 625 J of work done on the gas from the surroundings. Calculate both ΔH and ΔE for this reaction.

$$\Delta E = q + w = 785 \text{ J} + 625 \text{ J} = 1.41 \times 10^3 \text{ J}$$

$$\text{At constant pressure, } q_p = \Delta H = 785 \text{ J} = 0.785 \text{ kJ}$$