

AP Chemistry - Thermodynamics - "Work" Problems

1. A gas, initially at a pressure of 10.0 atm and having a volume of 5.00 L, is allowed to expand at a constant temperature and against a constant pressure of 4.00 atm, until it reaches a volume of 12.5 L. Calculate the work done by the gas on the surroundings.
2. Calculate the work done on the system when 6.00 L of a gas is compressed to 1.00 L under a constant pressure of 2.00 atm.
3. A gas is allowed to expand at constant temperature from a volume of 10.0 L to 20.0 L against an external pressure of 1.00 atm. If the gas also absorbs 255 J of heat from the surroundings, calculate the q , w , and the internal energy change of the system.
4. Calculate the work done when 24.3 g of Mg are reacted with an excess of HCl acid. Assume an external pressure of 1.00 atm and a temperature of 20.0°C.

Energy and enthalpy... and misunderstandings...

Confused about the difference between enthalpy and energy? I understand your confusion and, a large part of the problem seems to be in the formal definitions of these terms. In this formal sense, a system does not *have* heat, nor does it *have* work. Heat and work are rather mechanisms by which energy is transferred into and out of a system. When work is done on a system, the system's energy increases by that amount (w). When the energy of a system increases due to the flow of energy across a temperature gradient, this is referred to as heat (q).

Energy is a state function, therefore the change in energy (ΔE) when a system goes from one state to another depends only on the initial and final states. However, both heat and work are path-dependent. Since $\Delta E = q + w$, for all paths between the same initial and final states, ΔE is the same and therefore $q = \Delta E - w$. That is, ***the more energy that is released as work, the less is released as heat*** (remember that the +/- signs of q and w refer to the system, so a system that does work on the surroundings has $w < 0$ and a reaction that releases heat to the surroundings has $q < 0$).

In concrete terms, imagine the combustion of a fixed amount of fuel (such as Methane) in Oxygen occurring under two different sets of conditions. If the combustion occurs in a steel bomb calorimeter of constant volume, no work can be done ($w = P\Delta V = 0$), so all of the energy is released as heat since, under these conditions $\Delta E = q$. On the other hand, if the same amount of fuel is burned in Oxygen in a cylinder fitted with a moveable piston, so that the pressure remains constant, then the volume will increase and push against a constant pressure, doing work ($w = P\Delta V > 0$). In this reaction, the same energy will be released by the reaction, but a smaller amount of the energy will be released as heat ($q = \Delta E - w$). Most paths are much more complicated, where the pressure might change dramatically during the reaction, so there are many possible values of q and w for the same initial and final states, but they all give the same value of ΔE .

Because most chemical reactions occur under conditions of (roughly) constant pressure, the work done under these "special" conditions can be determined and subtracted from the ΔE . Because this represents the amount of energy that is released as heat under these constant pressure conditions and is of interest to chemists, it is called "Enthalpy." In one sentence, the energy absorbed or released as heat is equal to the change in energy minus the amount that is gained or lost to work due to changes in volume, or $\Delta H = \Delta E - P\Delta V$. Obviously, this is only important where in reactions where there are large changes in volume, that is when gases are consumed or produced.