THE FIRST LAW OF THERMODYNAMICS.

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Theory recap

Last time recap. Last time we discussed work done by gas and its' graphical representation. We have learned that in a process where pressure p of the gas is constant and its volume increases by ΔV, the work W done by the gas is calculated as follows:

$$W = p\Delta V.$$

For cyclic processes total work done by the cycle is numerically equal to the area enclosed inside the cycle depicted on a $p - V$ diagram.

Quantity of heat. Imagine heating a gas by a burner. The burner supplies some thermal energy to the gas. This energy coming from the burner is called quantity of heat and denoted by Q. Like all other kinds of energy, quantity of heat is measured in Joules.

First law of thermodynamics. Where does this energy go to? There are two possible directions. In order to identify them, let's first imagine that the temperature of the gas stays constant. Temperature of the gas is the measure of its internal energy and as temperature does not change, internal energy stays constant. When the volume increases, gas does work because of the heat supplied. Energy is always conserved, and since internal energy does not accumulate in the system, work W is just equal to the supplied amount of heat in this process:

$$W = Q \quad \text{for} \quad T = const.$$

On the other hand, let us look what happens if we fix the position of a piston. Then gas does not do any work and clearly because of the burner its temperature increases. Back when discussing temperature we've learned that temperature is a measure of internal energy of a substance. When temperature increases, internal energy also grows. Let us denote the change in internal energy by ΔE. Since the volume is fixed, no work is done by the gas. Therefore due to energy conservation all of the supplied heat Q should be accumulated in the system in the form of the internal energy:

$$\Delta E = Q \quad \text{for} \quad V = const.$$

What happens if we combine these two cases? Imagine we allow gas to change both its volume and temperature, which results in increase in internal energy of the gas by ΔE and work W done by the gas. The energy balance in this case works as follows: energy Q comes into the system from the outside, but energy W is spent on doing the work. Therefore the accumulation of internal energy is the difference between the supplied and spent energy:

$$\Delta E = Q - W.$$

This is **the first law of thermodynamics**. Essentially, it is nothing more but energy conservation law. Another common form of writing the first law of thermodynamics is as follows:

$$Q = \Delta E + W,$$

which means that all the heat supplied to the system could go either into work done by the system or into the change of internal energy.

Homework

1. 5 moles of helium were heated at constant pressure by an alcohol burner, so that helium's temperature increased by 300° C. Such an increase in temperature corresponds to an increase in internal energy of 18700 J (we will learn to calculate it next time). Every gram of burned alcohol releases $26,000$ J of heat. Assume that due to heat losses only 30% of heat released by the burner actually reaches the helium. What mass of alcohol was used to fuel the burner during this process?