Lesson 42b: First Law of Thermodynamics

The First Law of Thermodynamics can be stated in several ways.

- The easiest thing to say is that energy can not be created or destroyed, only change form or be transferred. This is sometimes called the **Law of Conservation of Energy**.
  - What we need to do at this point in the course is to focus on how this applies to gases in the kinetic theory of ideal gases.
- We will look at how adding **heat** and doing **work** to a gas can increase its **internal energy**, while removing **heat** and allowing the gas to do **work** will decrease its **internal energy**.

Internal Energy

Any gas will have a certain amount of **internal energy**.

You will see stuff about the internal energy of a **system** sometimes. Remember that a system is just a well defined area and all the things in it.

- This is due to things like the molecules of the gas moving around in the container, hitting each other, vibrating because of the energy they have, and so on.
  - All of these energies are happening at the microscopic level, for the molecules of the gas itself.
- Internal energy does ***not*** include anything that happens at the macroscopic (big) scale of things.
  - If you take a can of air freshener and throw it across the room you did ***not*** increase its internal energy. Throwing it gave the whole can (macroscopic) kinetic energy as it moved across the room, but had nothing to do with the energy of the gas molecules inside.
- We use the symbol **U** in formulas to show the **internal energy** of something, measured in Joules. We most often have to look at how the internal energy of a system has changed, so expect to see and use $\Delta U = U_f - U_i$ also.

Δ**U** is positive when the internal energy of the system has increased.

Δ**U** is negative when the internal energy of the system has decreased.

Heat

You might also remember doing calculations back in Science 10 to figure out how much **heat** was added to a substance to warm it up a few degrees using the formula $Q = mc\Delta T$.

- **Heat** is shown using the symbol **Q**.
  - Although it is measured in Joules, **heat** is not energy itself, it is the ***transfer* of energy**.
  - This is sort of like saying a engine in a car is not a force, but it exerts a force.
- If we add heat to a system (like a container of gas) we are really adding energy to its internal energy.

Q is positive when heat is added to the system.

Q is negative when heat is removed from the system.

Work

We can also perform **work** on a system, shown by the symbol **W** and measured in Joules.

- One example is when we push down on the piston of a cylinder of gas, reducing its volume.
  - If we smush the molecules together like this, then we are increasing the collisions that will happen in that smaller volume. Doing work on the gas increases the internal energy.
- It is also possible for the gas to do work, like if the gas pushes the cylinder back up, increasing its volume.
  - In this case the gas has done the work, so the internal energy of the gas decreased.

W is positive when work is done ***on*** the system.

W is negative when work is done ***by*** the system.

The First Law of Thermodynamics

Since we have three things that are all really just measurements of energy (they're all measured in Joules), it sounds like we should be able to write out an equation that says something about how energy is conserved in a system.

- We can not say anything about the initial internal energy of the system unless we are told specifically told about it.
- Instead, we will measure how much we **changed the internal energy** of the system by **adding/removing heat** and doing **work on/by** the gas.

$$\Delta U = Q + W$$

Example 1: We have a canister of gas that has a plunger on top. We add 325 J of heat to the canister while pushing the plunger down 1.50 m with 200 N of force. **Determine** how much the internal energy of the canister has changed.

$$\Delta U = Q + W$$
$$\Delta U = Q + Fd$$
$$\Delta U = 325 + 200(1.50)$$
$$\Delta U = 625\,J$$

Don't forget that W = Fd.

The internal energy of the canister has *increased* by 625 J.

Example 2: A cylinder loses 475 J of heat while expanding a piston 2.30 m with 100 N of force. **Determine** how much the internal energy of the canister has changed.

$$\Delta U = Q + W$$
$$\Delta U = Q + Fd$$
$$\Delta U = -475 + -[100(2.30)]$$
$$\Delta U = -705\,J$$

The negative sign in front of the brackets does not suggest a direction to the force or displacement. Instead, it is because the gas expanded the cylinder, meaning the work was done ***by*** the gas.

The internal energy of the canister has *decreased* by 705 J.