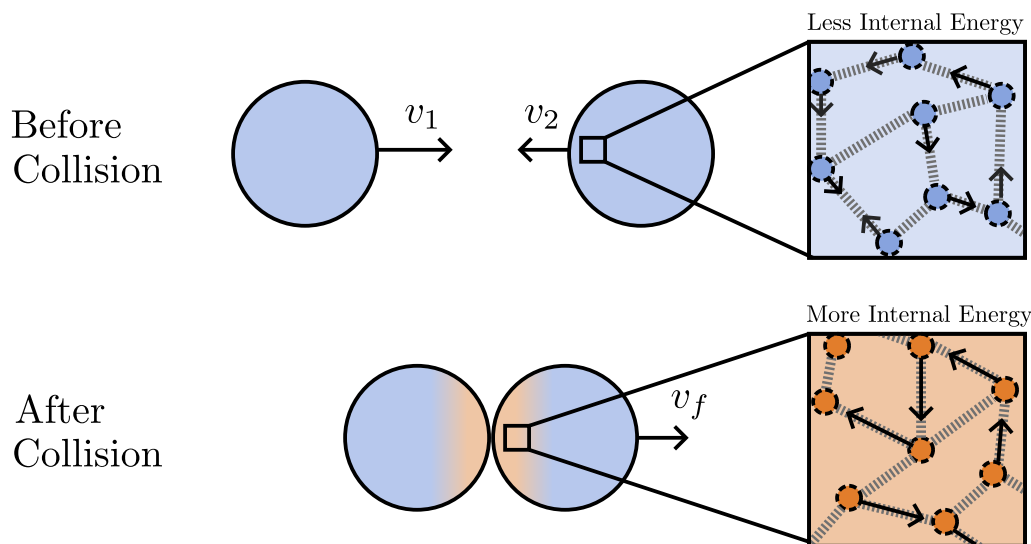# 1   Energy and the first law

Figure 1: Energy flow in an inelastic collision. The lost kinetic energy turns into **internal** energy. We divide internal energy in two parts: work to deform the balls, and heat energy.

Summary of concepts:

- Partitioning of energy in an inelastic collision: $\sum_i \frac{1}{2} m_i v_{i,initial}^2 + \sum_i U_{i,initial} = \sum_i \frac{1}{2} m_i v_{i,final}^2 + \sum_i U_{i,final}$

- In PHYS 213, we are mainly concerned with changes in internal energy. We ignore all the v's.

- Internal energy changes are broken into two parts $dU = dQ - pdV$. The second term is "work" and represents deformation of the object. The first term is everything else, which goes into the microscopic motion of particles.

- The first law of thermodynamics is energy conservation. Total change in internal energy, is work plus heat.

You might recall that in your classical mechanics courses, you were told about *inelastic* and *elastic* collisions. The difference between them is that *kinetic energy* is conserved in an elastic collision, while it is not in an inelastic collision. Where does the energy go in the inelastic collision, like the one in Figure 1? It goes into the microscopic motion of atoms; they actually move faster after the inelastic collision happens than they were moving previously. We perceive this as an increase in temperature.

Thermodynamics is about the flow of **thermal energy**. Thermal energy is the microscopic kinetic and potential energy of atoms that we can't observe very easily in our day to day lives. This motion is very similar to the mechanics that you've already studied in physics, but it involves the motion of $\sim 10^{23}$ atoms for a roughly human-sized object. One could try to keep track of the velocities and positions of all these atoms,[1] but most of us would quickly run out of patience trying to do that. We'd prefer to instead use *collective* descriptions of the atoms, like how much total energy is in their motion, and how compressed they are. In future chapters, we will introduce other collective variables, like temperature and entropy. These collective variables are called **macrostates**.

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[1]It turns out that you can do computer simulations on millions of atoms which actually do act very similarly to 10^{23} particles; this is an entire field of research!
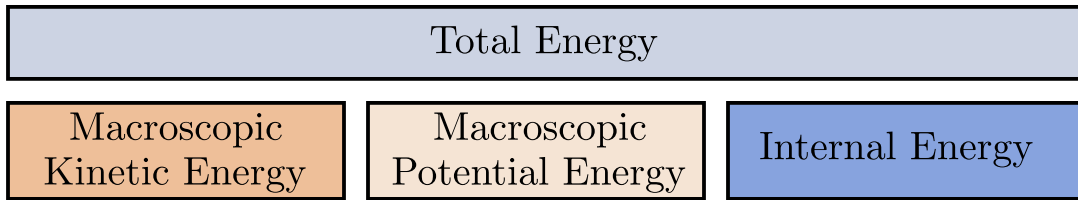
Figure 2: The total energy of a system is divided into macroscopic kinetic energy, macroscopic potential energy, and internal energy.
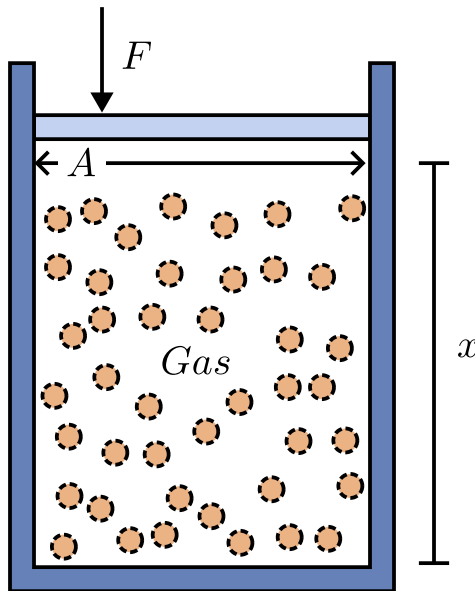

Figure 3: The relation of pdV to $Fd\ell$. The pressure is the force per area.

1.1   Internal energy

Internal energy is all the energy that is not macroscopic kinetic and potential energy, which is present in all the microscopic motions inside a given object. We typically write internal energy as U. A rough idea of the relationship between all the different energies is present in Figure 2. In reality, the internal energy is actually much larger than the macroscopic potential and kinetic energy. Energy can move between the different classifications. In this course, we will focus on the behavior of just the internal energy.

It is often very difficult to compute or measure the total value of the internal energy U. However, as you may recall from classical mechanics, all that matters for physics is not the total value of the energy, but the *change* in energy. So in thermodynamics, we typically compute only the change in internal energy during a process.

Example: Suppose that two balls collide inelastically on a flat plane (so potential energy is zero). The total kinetic energy before the collision is 40 J, while after the collision, the balls continue on but with only 20 J of kinetic energy. The internal energy of the balls increases by 20 J, the remaining energy; that is, $\Delta U = U_f - U_i = 20$ J. During this process, we know the change in U but not its total value.

1.2   Work

The balls in the inelastic collision may become damaged or deformed. It takes some energy to do this deformation, which is work. In other systems such as gas in a balloon, it takes energy to compress the

balloon. As we learned in classical mechanics, the work we do **on** an object is

$$W_{on} = \int_{x_i}^{x_f} F dx, \tag{1}$$

where we have applied force F over a distance from an initial position x_i to a final position x_f. In thermodynamics we will consider work on the object as changing its total internal energy.

To understand work a little more concretely, consider a piston (Figure 3) experiment. We will use these often in the course, so it's worth getting used to the setup. The chamber is filled with a gas and sealed, so no gas can get in or out. The sides and bottom are fixed in place. The top of the piston is allowed to move up and down. We will identify the position of the top of the piston as x from Equation 1.

Now suppose we apply a force F to the top of the piston. Using Equation 1, the work on the gas is

$$W_{on} = \int_{x_i}^{x_f} F dx \tag{2}$$

$$= \int_{x_i}^{x_f} \frac{F}{A} A dx \tag{3}$$

$$= \int_{V_i}^{V_f} \frac{F}{A} dV \tag{4}$$

$$, \tag{5}$$

where in the last step we used "u-substitution" to replace our integration variable with volume $V = Ax$. Finally, we substitute pressure $p = -F/A$ (Pressure always pushes outward!) to get our final result

$$W_{on} = - \int_{V_i}^{V_f} p dV. \tag{6}$$

In thermodynamics we typically think of collective variables volume and pressure, rather than force on a given point and specific dimensions of the object. Writing work in this form allows us to use those collective variables.

Work on versus work by. In the previous paragraph, we imagined deforming a gas by doing work on it. If the gas were expanding, we might prefer to say that the gas itself is *doing* work. In thermodynamics, that is referred to as work **by** the gas. Very simply, $W_{on} = -W_{by}$ from Newton's third law, but it's very important to keep them straight. A simple mnemonic is that if the volume of the object is increasing, it is doing work and so W_{by} is positive, and vice versa if the volume is decreasing.

1.3 Heat

Heat energy is all the energy that enters the object that is not macroscopic kinetic energy (which for the most part we ignore in thermodynamics) or work. We often call this Q. For the most part this is the microscopic vibrations of bonds between atoms, and in the case of gases, the velocity of molecules/atoms. Later on in this course, we will explicitly count up the heat energy for an idealized gas.

Note that both heat and work are defined only for changes over a process. There is no such thing as "total heat" or "total work" independent of a process. On the other hand, in principle there is a total internal energy of a given object.

1.4 Differentials and the first law

The first law of thermodynamics really just follows from the definition of heat and work. It says that

$$\Delta U = W_{on} + Q. \tag{7}$$

In other words, the change in internal energy during a process is the work done on it plus the heat energy that flows in. This is true by definition of heat energy, since we defined heat energy as all the energy that wasn't covered by work! However, the law is not usually written quite like the above; it is usually written in terms of differentials.

To write the first law in terms of differentials, note that

$$U_f - U_i = \int_{U_i}^{U_f} dU = \int_0^Q dQ - \int_{V_i}^{V_f} pdV. \tag{8}$$

This is a very useful equation that you will use throughout the course. Note that the integrals are important here; any dependence of the pressure on the volume must be included. In thermodynamics, often this equation will be written as

$$dU = dQ - pdV, \tag{9}$$

where it is implied that the equation must be integrated over a process to get a change in internal energy.