The laws of thermodynamics

First and Second Laws of Thermodynamics, as they apply to biological systems.

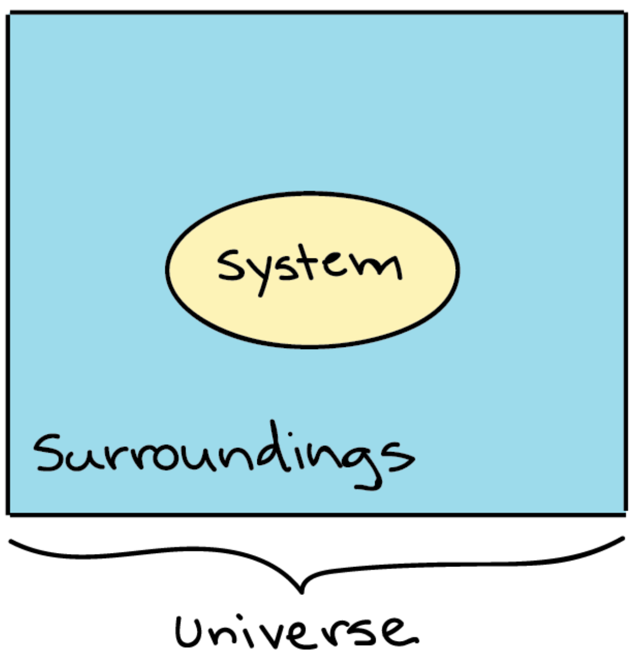
Introduction

What kind of system are you: open or closed? As it turns out, this is a physics question, not a philosophical one. You, like all living things, are an open system, meaning that you exchange both matter and energy with your environment. For instance, you take in chemical energy in the form of food, and do work on your surroundings in the form of moving, talking, walking, and breathing.

All of the exchanges of energy that take place inside of you (such as your many metabolic reactions), and between you and your surroundings, can be described by the same laws of physics as energy exchanges between hot and cold objects, or gas molecules, or anything else you might find in a physics textbook. Here, we'll look at two physical laws – the First and Second Laws of Thermodynamics – and see how they apply to biological systems like you.

Systems and surroundings

Thermodynamics in biology refers to the study of energy transfers that occur in molecules or collections of molecules. When we are discussing thermodynamics, the particular item or collection of items that we're interested in (which could be something as small as a cell, or as large as an ecosystem) is called the **system**, while everything that's not included in the system we've defined is called the **surroundings**.



For instance, if you were heating a pot of water on the stove, the system might include the stove, pot, and water, while the surroundings would be everything else: the rest of the kitchen, house, neighborhood, country, planet, galaxy, and universe. The decision of what to define as the system is arbitrary (up to the observer), and depending on what you wanted to study, you could equally well make just the water, or the entire house, part of the system. The system and the surroundings together make up the **universe**.

There are three types of systems in thermodynamics: open, closed, and isolated.

An **open system** can exchange both energy and matter with its surroundings. The stovetop example would be an open system, because heat and water vapor can be lost to the air.

A **closed system**, on the other hand, can exchange only energy with its surroundings, not matter. If we put a very tightly fitting lid on the pot from the previous example, it would approximate a closed system.

An **isolated system** is one that cannot exchange either matter or energy with its surroundings. A perfect isolated system is hard to come by, but an insulated drink cooler with a lid is conceptually similar to a true isolated system. The items inside can exchange energy with each other, which is why the drinks get cold and the ice melts a little, but they exchange very little energy (heat) with the outside environment.

If you've watched the video on the Second Law of Thermodynamics and entropy, you may have noticed that Sal refers to this same example, a cooler, as an approximation of a "closed" system. Here, it's described as an approximation of an "isolated" system. What's up with that?

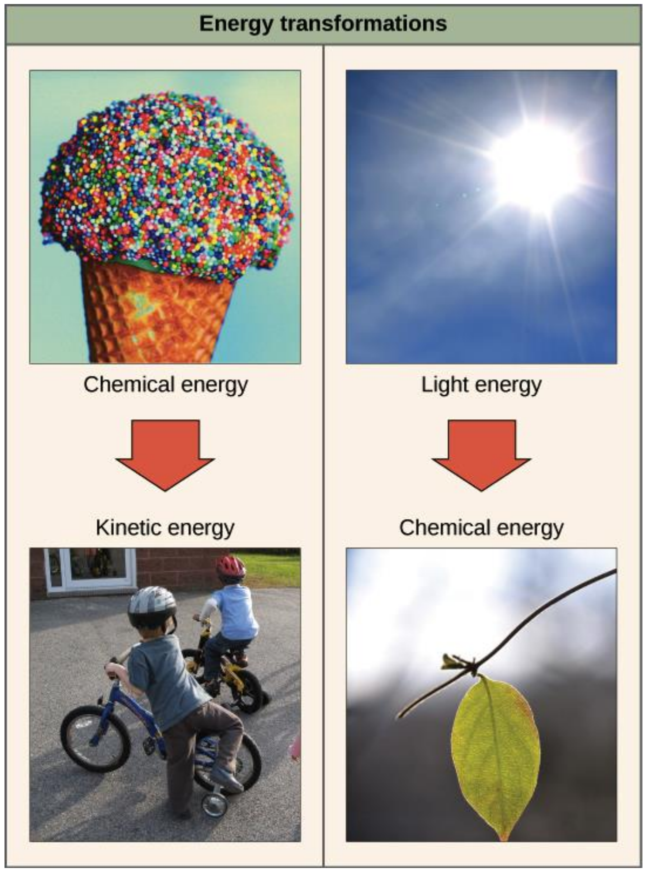
As it turns out, this is a case where two different branches of physics use terms in slightly different ways. A system that doesn't exchange either energy or matter with its environment is called an "isolated" system by physicists who study thermodynamics, but a "closed " system by physicists who study classical mechanics. In his video, Sal uses the classical mechanics definition, while this article uses the thermodynamics definition.

You, like other organisms, are an open system. Whether you think about it or not, you are constantly exchanging energy and matter with your surroundings. For instance, suppose that you eat a carrot, or lift a bag of laundry onto a table, or simply breathe out and release carbon dioxide into the atmosphere. In each case, you are exchanging energy and matter with your environment.

Exchanges of energy that take place in living creatures must follow the laws of physics. In this regard, they are no different from energy transfers in, say, an electrical circuit. Let's take a closer look at how the **laws of thermodynamics** (physical rules of energy transfer) apply to living beings like yourself.

The First Law of Thermodynamics

The first law of thermodynamics thinks big: it deals with the total amount of energy in the universe, and in particular, it states that this total amount does not change. Put another way, the **First Law of Thermodynamics** states that energy cannot be created or destroyed. It can only change form or be transferred from one object to another.



This law may seem kind of abstract, but if we start to look at examples, we'll find that transfers and transformations of energy take place around us all the time. For example:

- Light bulbs transform electrical energy into light energy (radiant energy).
- One pool ball hits another, transferring kinetic energy and making the second ball move.
- Plants convert the energy of sunlight (radiant energy) into chemical energy stored in organic molecules.
- You are transforming chemical energy from your last snack into kinetic energy as you walk, breathe, and move your finger to scroll up and down this page.

Importantly, none of these transfers is completely efficient. Instead, in each scenario, some of the starting energy is released as thermal energy. When it's moving from one object to another, thermal energy is called by the more familiar name of **heat**. It's obvious that glowing light bulbs generate heat in addition to light, but moving pool balls do too (thanks to friction), as do the inefficient chemical energy transfers of plant and animal metabolism. To see why this heat generation is important, stay tuned for the Second Law of Thermodynamics.

The Second Law of Thermodynamics

At first glance, the first law of thermodynamics may seem like great news. If energy is never created or destroyed, that means that energy can just be recycled over and over again, right?

Well…yes and no. Energy cannot be created or destroyed, but it can change from more-useful forms into less-useful forms. As it turns out, in every real-world energy transfer or transformation, some amount of energy is converted to a form that's unusable (unavailable to do work). In most cases, this unusable energy takes the form of heat.

Although heat can in fact do work under the right circumstances, it can never be turned into other (work-performing) types of energy with 100% efficiency. So, every time an energy transfer happens, some amount of useful energy will move from the useful to the useless category.

Heat increases the randomness of the universe

If heat is not doing work, then what exactly does it do? Heat that doesn't do work goes towards increasing the randomness (disorder) of the universe. That may seem like a big logic jump, so let's take a step back and see how it can be the case.

When you have two objects (say, two blocks of the same metal) at different temperatures, your system is relatively organized: the molecules are partitioned by speed, with those in the cooler object moving slowly and those in the hotter object moving quickly. If heat flows from the hotter object into the cooler object (as it will spontaneously), the molecules of the hot object slow down, and the molecules of the cool object speed up, until all the molecules are moving at the same average speed. Now, rather than having a partition of between fast and slow molecules, we simply have one big pool of molecules going about the same speed – a less ordered situation than our starting point.

The system will tend to move towards this more disordered configuration simply because it's statistically much more likely than the temperature-separated configuration (i.e., there are many more possible states corresponding to the disordered configuration). You can explore this concept further in the videos in this tutorial, or in this straightforward physics video.

Entropy and the Second Law of Thermodynamics

The degree of randomness or disorder in a system is called its **entropy**. Since we know that every energy transfer results in the conversion of some energy to an unusable form (such as heat), and since heat that does not do work goes to increase the randomness of the universe, we can state a biology-relevant version of the **Second Law of Thermodynamics**: every energy transfer that takes place will increase the entropy of the universe and reduce the amount of usable energy available to do work (or, in the most extreme case, leave the overall entropy unchanged). In other words, any process, such as a chemical reaction or set of connected reactions, will proceed in a direction that increases the overall entropy of the universe.

This is actually a pretty crazy concept, if you extend it to its logical conclusion: that someday, all of the usable energy in the universe may have been converted to unusable energy (heat). This state has been suggested as one possible fate of the universe, and it's sometimes called the "heat death of the universe." However, it's not clear to physicists that this is actually what will happen. And even if it does, it won't be for another 10105610^{10^{56}}10105610, start superscript, 10, start superscript, 56, end superscript, end superscript years or so – so not something to keep you awake at night!

To sum up, the First Law of Thermodynamics tells us about conservation of energy among processes, while the Second Law of Thermodynamics talks about the directionality of the processes, that is, from lower to higher entropy (in the universe overall).