Does Life on Earth Violate the Second Law of Thermodynamics?

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The second law of thermodynamics (the law of increase of entropy) is sometimes used as an argument against evolution. Evolution, the argument goes, is a decrease of entropy, because it involves things getting more organized over time, while the second law says that things get more disordered over time. So evolution violates the second law.

There are many things wrong with this argument, and it has been discussed ad infinitum. However, these discussions never seem to involve any numerical calculations. This is unfortunate, since a very simple calculation shows that it is physically impossible for evolution to violate the second law of thermodynamics.

It is important to note that the earth is not an isolated system: it receives energy from the sun, and radiates energy back into space. The second law doesn't claim that the entropy of any part of a system increases: if it did, ice would never form and vapor would never condense, since both of those processes involve a decrease of entropy. Rather, the second law says that the total entropy of the whole system must increase. Any decrease of entropy (like the water freezing into ice cubes in your freezer) must be compensated by an increase in entropy elsewhere (the heat released into your kitchen by the refrigerator).

A slightly more sophisticated form of the anti-evolution argument recognizes that the earth is not an isolated system; it receives energy from the sun. But, the argument goes on, the sun's energy only increases disorder. It speeds the processes of breakdown and decay. Therefore, even with an energy source, evolution still violates the second law.

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For the earth, though, we have to take into account the change of entropy involved with both the absorption of energy from the sun and the radiation of energy into space. Think of the sun as a heat reservoir that maintains a constant temperature $T_1 = 6000K$. (I am using the absolute, or Kelvin, temperature scale.) That's the temperature of the radiating surface of the sun, and so it's the effective temperature of the energy we receive from the sun. When the earth absorbs some amount of heat, Q, from this reservoir, the reservoir loses entropy:

$$\Delta S_1 = \frac{-Q}{T_1}$$

On average, the earth's temperature is neither increasing nor decreasing. Therefore, in the same time that it absorbs heat energy Q from the sun's radiation, it must radiate the same amount of heat into space. This energy is radiated at a much lower temperature that is approximately equal to the average surface temperature of the earth, $T_2 = 280K$. We can think of space as a second heat reservoir that absorbs the heat Q and consequently undergoes an entropy increase

$$\Delta S_2 = \frac{+Q}{T_2}$$

Since T_1 is much larger than T_2, it is clear that the net entropy of the two reservoirs increases:

$$\Delta S_1 + \Delta S_2 = \frac{-Q}{T_1} + \frac{Q}{T_2} = Q\left(\frac{T_1 - T_2}{T_1 T_2}\right) > 0$$

Even if it is true that the processes of life on earth result in an entropy decrease of the earth, the second law of thermodynamics will not be violated unless that decrease is larger than the entropy increase of the two heat reservoirs. Any astronomy textbook will tell you that the earth absorbs 1.1×10^{17} Joules per second of power from the sun, so in one year we get $(1.1 \times 10^{17} \frac{J}{sec}) \times (365 \frac{days}{year}) \times (24 \frac{hrs}{day}) \times (60 \frac{min}{hr}) \times (60 \frac{sec}{min}) = 3.5 \times 10^{24}$ Joules of energy from the sun. This corresponds to an entropy increase in the heat reservoirs of

$$\Delta S_1 + \Delta S_2 = \frac{-3.5 \times 10^{24} J}{6000K} + \frac{3.5 \times 10^{24} J}{280K} = 1.2 \times 10^{22} \frac{J}{K}$$

Just how big is this increase? For comparison, let's calculate the entropy change needed to freeze the earth's oceans solid. The heat energy involved is

$$\begin{aligned} Q &= \text{(latent heat of fusion)} \times \text{(mass of ocean water)} \\ &= \left(3.3 \times 10^5 \frac{J}{kg}\right) \times \left(1.3 \times 10^{21} kg\right) = 4.3 \times 10^{26} J \end{aligned}$$

Water freezes at 273 K on the absolute scale, so the corresponding entropy change is

$$\Delta S_{ocean} = \frac{-Q}{T} = \frac{-4.3 \times 10^{26} J}{273K} = -1.6 \times 10^{24} \frac{J}{K}$$

Comparing with the entropy increase of the two heat reservoirs, we see that this is a factor of $(1.6 \times 10^{24} \frac{J}{K})/(1.2 \times 10^{22} \frac{J}{K}) \approx 130$ times larger. Remember, though, that the number for the heat reservoirs was for one year. Each year, more entropy is generated. The second law will only be violated if all the oceans freeze over in about 130 years or less.

Now, the mass of all the living organisms on earth, known as the *biomass*, is considerably less than the mass of the oceans (by a very generous estimate, about 10^{16} kilograms). If we perform a similar calculation using the earth's biomass instead of the mass of the oceans, we find that the second law of thermodynamics will only be violated if the entire biomass is somehow converted from a highly-disorganized state (say, a gas at 10,000 K) to a highly-organized state (say, absolute zero) in about a month or less.

Evolutionary processes take place over millions of years; clearly they cannot cause a violation of the second law.

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[1] This article is adapted from my notes for Mr. Tompkins Gets Serious: The Essential George Gamow.