The entropy of a substance is given as

\[ S = k \ln \Omega \]  

(1)

where \( \Omega \) is the number of microstates accessible to the substance.

For a 3-d ideal gas, this is given by the Sackur-Tetrode formula:

\[ S = Nk \left[ \ln \left( \frac{V}{N} \left( \frac{4\pi mU}{3Nk^2} \right)^{3/2} \right) + \frac{5}{2} \right] \]  

(2)

where \( V \) is the volume, \( U \) is the energy, \( N \) is the number of molecules, \( m \) is the mass of a single molecule and \( h \) is Planck’s constant.

Although this formula looks a bit complicated, we can see that increasing any of \( V \), \( U \) or \( N \) increases the entropy. For an isothermal expansion, the gas expands quasistatically so that its temperature stays constant. This means that \( U = \frac{3}{2} NkT \) also stays constant, so that only the volume changes. Since the gas is doing work \( W \) by expanding, the energy for the work must be provided by an amount of heat \( Q \) input into the gas to maintain the temperature as constant. This heat is given by the formula

\[ Q = NkT \ln \frac{V_f}{V_i} \]  

(3)

where \( V_i \) and \( V_f \) are the initial and final volumes.

However, from (2), the change in entropy in a process where only the volume changes is

\[ \Delta S = S_f - S_i = Nk \ln \frac{V_f}{V_i} \]  

(4)

Combining these two equations gives

\[ \Delta S = \frac{Q}{T} \]  

(5)

This relation is valid for the case where the expanding gas does work, so that heat must be input to provide the energy for the work. In a free
expansion, the gas expands into a vacuum so does no work (well, technically, after some of the gas has entered the vacuum area, it’s no longer a vacuum so that some work is done, but we’ll assume the vacuum area is very large so we can neglect this). In this case, the internal energy $U$ still doesn’t change, since the gas neither absorbs any heat nor does any work, so $\Delta U = Q + W = 0$. However, the volume occupied by the gas does increase (and it’s the only thing that changes) so 4 is still valid, although 5 is not.

Another property of 2 is that if the energy $U$ drops low enough, the log term can decrease below $-\frac{5}{2}$ making $S$ negative. This isn’t possible, so the Sackur-Tetrode equation must break down at low energies. For a monatomic ideal gas, $U = \frac{3}{2}NkT$, so this implies that things go wrong for low temperatures. For example, suppose we have a mole of helium and cool it (assuming it remains a gas). Then the critical temperature is found from

$$\frac{-5}{2} = \ln\left(\frac{V}{N}\left(\frac{2\pi mkT}{h^2}\right)^{3/2}\right)$$

$$T_{\text{crit}} = \frac{h^2}{2\pi mk}\left(\frac{Ne^{-5/2}}{V}\right)^{2/3}$$

If we start at room temperature $T = 300 \text{ K}$ and atmospheric pressure $P = 1.01 \times 10^5 \text{ Pa}$, and can hold the density $N/V$ fixed, this will give an actual temperature at which the entropy becomes zero. The density is

$$\frac{N}{V} = \frac{P}{kT} = 2.44 \times 10^{25} \text{ m}^{-3}$$

The mass of a helium atom is $4 \times 10^{-3} \text{ kg mol}^{-1}/6.02 \times 10^{23}$, so plugging in the other values gives

$$T_{\text{crit}} = 0.012 \text{ K}$$

In fact, helium liquefies at around 4 K, so it appears that 2 might actually be valid for the region where helium remains a gas.

As a final example, we can observe that the entropy of an ideal gas is $Nk$ multiplied by a logarithm, and of an Einstein solid is also $Nk$ multiplied by a logarithm (because $\Omega \approx (qe/N)^N$ for high-temperature solids). For any macroscopic object, $N$ is a large number and the logarithm is much smaller, so for a rough order-of-magnitude estimate of the entropy, we can neglect the log term and take $S \sim Nk$. A few such estimates are:

For a 1 kg book, we can take it to be 1 kg of carbon, with a molar mass of $12 \times 10^{-3} \text{ kg mol}^{-1}$, so the entropy of a book is around
\[ S \sim \frac{6.02 \times 10^{23}}{12 \times 10^{-3}} (1) (1.38 \times 10^{-23}) = 692 \text{ J K}^{-1} \quad (10) \]

For a 400 kg moose, which we can approximate by 400 kg of water with molar mass of around \( 18 \times 10^{-3} \) kg \( \text{mol}^{-1} \), we have

\[ S \sim \frac{6.02 \times 10^{23}}{18 \times 10^{-3}} (400) (1.38 \times 10^{-23}) = 1.85 \times 10^5 \text{ J K}^{-1} \quad (11) \]

For the sun, we can take it to be \( 2 \times 10^{30} \) of ionized hydrogen (protons) with molar mass of \( 10^{-3} \) kg \( \text{mol}^{-1} \). The entropy is around

\[ S \sim \frac{6.02 \times 10^{23}}{10^{-3}} (2 \times 10^{30}) (1.38 \times 10^{-23}) = 1.66 \times 10^{34} \text{ J K}^{-1} \quad (12) \]

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