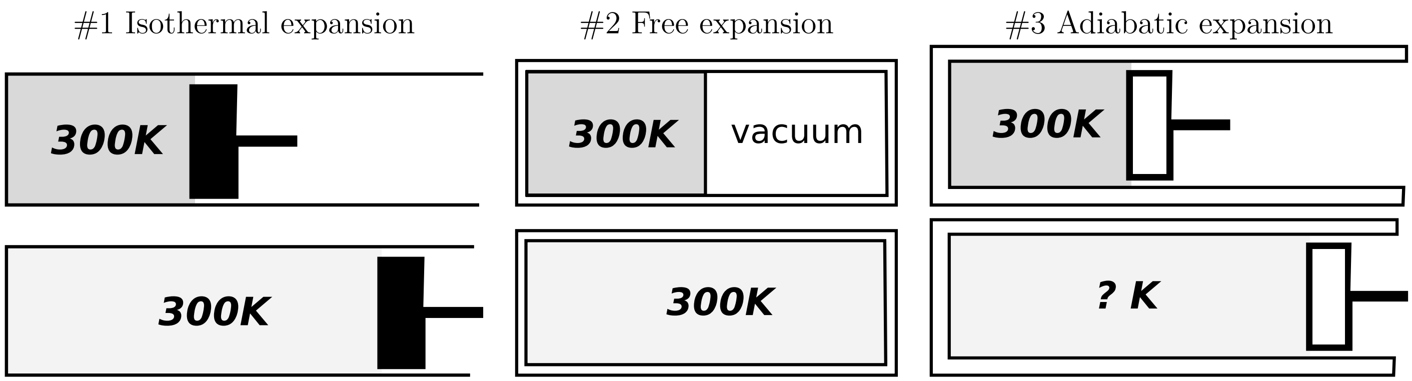
Consider the three processes described below.

Process #1 Five moles of an ideal gas are initially confined in a one-liter cylinder with a movable piston, at a temperature of 300 K. Slowly the gas expands against the movable piston, while the cylinder is in contact with a thermal reservoir at 300 K. The temperature of the gas remains constant at 300 K while the volume increases to two liters.

Process #2 A thin plastic sheet divides an insulated two-liter container in half. Five moles of the same ideal gas are confined to one half of the container, at a temperature of 300 K. The other half of the container is a vacuum. The plastic divider is suddenly removed and the gas expands to fill the container. No work is done on or by the gas. The final temperature of the gas is also 300 K.

Process #3 The same cylinder as in process #1 is thermally insulated and then allowed to slowly expand, starting at 300 K, to twice its original size (two liters).



1. Consider the change in the entropy of the gas for process 1, 2 and 3. We'll call these changes ΔS_1, ΔS_2 and ΔS_3. Are these changes positive, negative, or zero? Please explain your reasoning.

Solution

Process #1 As the gas is expanded isothermally, the gas does work on the piston, and thus loses energy due to work. To stay at the same temperature, the gas must gain energy by heating, which means that since $dS = đQ/T$ its entropy must increase.

Process #2 In the free expansion, the system has the same initial and final state as for the previous case, and so its entropy must increase. *Warning: This is not a quasistatic process as the gas rushes to fill the vacuum, so we cannot make direct use of* $dS = đQ/T$*!!!*

Process #3 In this case, there is no energy flow do to heating because we insulated the system. Since it is a slow process, we can conclude that the entropy of the system remains unchanged. It's OK that $\Delta S_1 \neq \Delta S_3$ because the final state of process 1 is different from the final state of process 3 (different T).

2. Is ΔS_1 greater than, less than, or equal to ΔS_2? How do each of these compare with ΔS_3? Please explain.

Solution $\Delta S_1 = \Delta S_2$ because these processes have the same initial state and final state. In contrast, $\Delta S_3 = 0$.

3. Consider the change in entropy of the surroundings for process 1, 2 and 3. We'll call these changes $\Delta S_{\text{surr-1}}$, $\Delta S_{\text{surr-2}}$ and $\Delta S_{\text{surr-3}}$. Are these changes positive, negative, or zero? Please explain.

Solution

Process #1 The surroundings loses energy by heating the gas and must therefore see its entropy decrease.

Process #2 The surroundings are not affected by what the gas is doing in its box, so the entropy of the surroundings remains fixed.

Process #3 The entropy of the surroundings remains unchanged for the same reason the entropy of the system remained unchanged.

Solution The bit of this activity that most students find challenging and confusing is that in the second process the total entropy of system plus surroundings *increases*. I frequently get the question, "but where did the entropy come from?" My answer is that entropy is created whenever there is an irreversible process. This is the essence of the Second Law: you can create entropy, but you can never destroy entropy.