Entropy

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12:39 PM

disordered transmission of energy

Entropy

Entropy is the random or disordered motion of molecules. The absolute entropy of a particular object is not measureable, only the change in entropy can be calculated.

In formula change in entropy is represented by ΔS and has units of Joules / Kelvin (or cal / K).

** $\Delta S = Q / T$

Note that Q may be calculated in any thermodynamic manner ($Q = mc\Delta T$, mL, or $\Delta U = Q + W$)

Using eqn. ** is simple if T is constant. But in cases where $\Delta T \neq 0$ the calculation can still be performed using average temperatures if ΔT is small

Example 1 change in entropy during a phase change:

A 1.25 kg mass of lead at 327 °C is allowed to melt, calculate the change in entropy.

- $\Delta S = \dfrac{Q}{T}$ ← $Q = mL_f = 1.25 \times 0.25 \times 10^5 = 3.125 \times 10^4$ J
  ← $T = 327 + 273 = 600$ K

- $\Delta S = 52$ J / K

$\Delta E = 0 = mc\Delta T_{30} + mc\Delta T_{20}$
$0 = .4(4186)(T_f - 30) + .3(4186)(T_f - 20)$
$T_f = 25.7°C$

Example 2 change in entropy involving temperature change

A 400 gram sample of water at 30 °C is mixed with a 300 gram sample at 20 °C. Calculate the change in entropy of the 400 g sample, the 300 g sample and the entire system.

First calculate the final temperature of the entire system:

- $\Delta Q = 0 = mc\Delta T_{30} + mc\Delta T_{20}$
  $0 = .400\,(4186)(T_f - 30) + .300\,(4186)(T_f - 20)$
  $0 = 1674.4T_f - 50232 + 1255.8T_f - 25116$
  $0 = 2930.2T_f - 75348$
  $T_f = 25.7$ °C $= 298.7$ K

Using the final temperature determine the change in entropy of each sample in the experiment:

30→25.7 20→25.7

- $\Delta S = \Delta S_{hot} + \Delta S_{cold}$
- $\Delta S = \dfrac{mc\Delta T_{hot}}{T_{avehot}} + \dfrac{mc\Delta T_{cold}}{T_{avecold}} = \dfrac{[0.400\,(4186)(25.7 - 30)}{[(25.7 + 30)/2]+273} + \dfrac{0.300(4186)(25.7 - 20)]}{[(25.7 + 20)/2] + 273} = \dfrac{-7199.92}{300.85} + \dfrac{7158.06}{295.85}$

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Note that while the hot body has ΔS negative the cold body has a larger ΔS positive resulting a positive

- $\Delta S = -23.9 + 24.2 = 0.30$ J /K

Note that while the hot body has ΔS negative the cold body has a larger ΔS positive resulting a positive ΔS for the entire experiment.

Notice that the change in entropy of a system must always be positive (or 0 in the ideal case). The total entropy then must increase in an isolated system. The universe may be thought of as an isolated system and therefore any process which occurs must increase the entropy of the universe. This is another expression of the 2nd Law of Thermodynamics. The concept of entropy was initially put forward by Clausius in the 1800's, today it represents amount of disorder. By order physicists refer to structured state, disorder then is the lack of structure. Random molecular motion would be a measure of disorder, the shattered cup from the example on page 451 shows a change toward disorder illustrating entropy from Clausius' statement. The heat energy in a cold reservoir of an engine represents entropy in its molecular motion illustrating the Kelvin / Planck statement of the 2nd law. Thus the 2nd law can be written as: "the total entropy of any system (the universe) must increase during any process", or in terms of disorder "any system will move toward greater disorder due to natural processes".

Heat Death

Heat cannot die, it simply spreads over so large a volume that it cannot be detected. If all processes lead to some heat energy being generated in random molecular motion then eventually all processes will terminate in heat energy. At this point all energies will have reached their lowest possible level and no physical, chemical or nuclear process should be able to occur following this.

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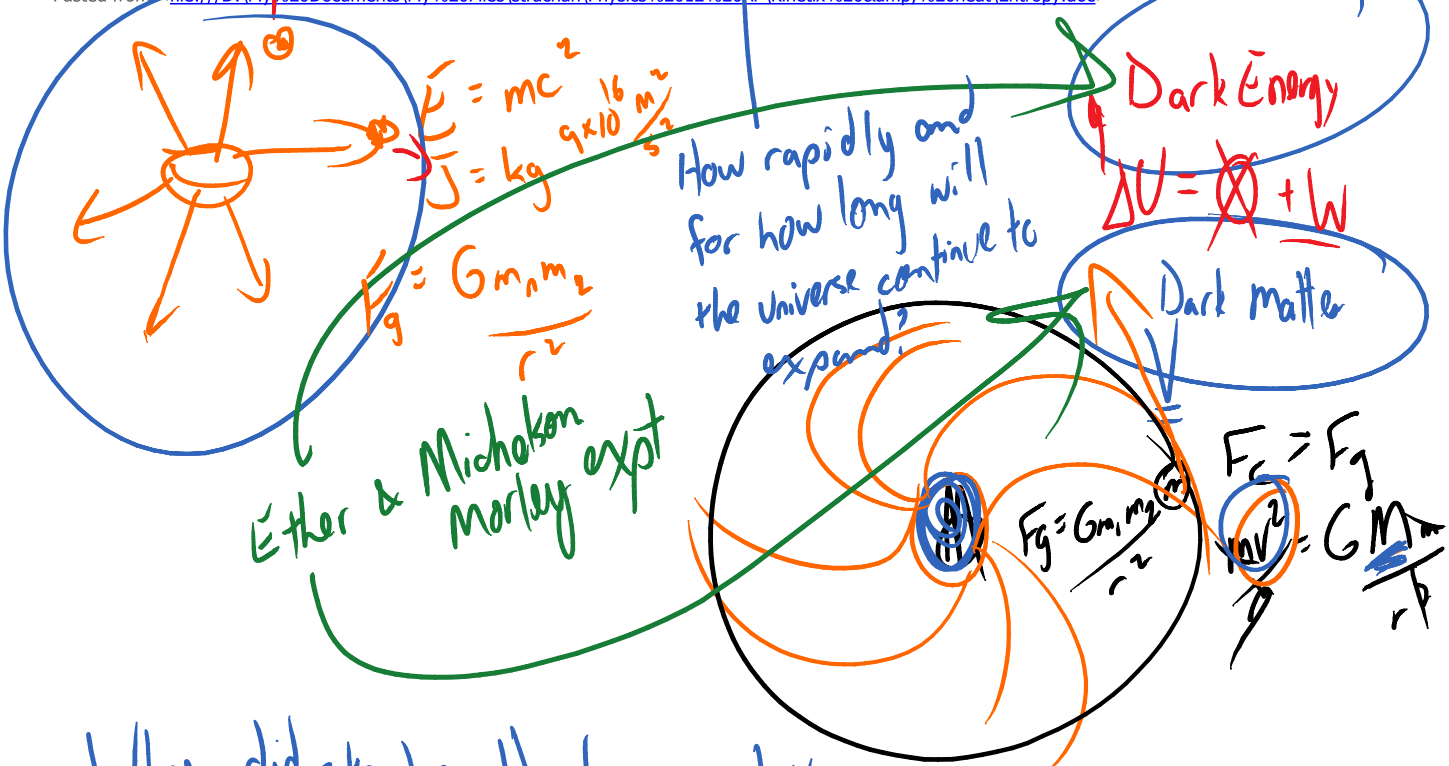

Why did stephen Hawking postulate the existance of Hawking radn in reference to blackholes & thermodynam.

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