Understanding Entropy in CH 223

Entropy, **S**, is a measure of the disorder of a system. According to the second law of thermodynamics, the total entropy of the universe is continually increasing, so everything we do as chemists and as human beings must somehow be increasing the quantity of disorder in the universe.

The third law of thermodynamics states that the entropy of each element in some crystalline state at zero Kelvin equals zero. Realize that the chances of finding any element in this idealized (and vague!) crystalline state at 0 K is about zero, so *everything* has some measure of entropy to it; *viz.*, **the entropy of *every* substance will be greater than zero at temperatures greater than zero Kelvin.**

Some generalizations regarding entropy:

1) The entropy of gases is greater than the entropy of liquids, and the entropy of liquids is greater than the entropy of solids.
   Example: $H_20_{(g)}$: $S = 188.8\ J\ K^{-1}\ mol^{-1}$
   Example: $H_20_{(l)}$: $S = 69.9\ J\ K^{-1}\ mol^{-1}$

2) The entropy of complex molecules is greater than the entropy of simple molecules.
   Example: $C_3H_{8(g)}$: $S = 269.9\ J\ K^{-1}\ mol^{-1}$
   Example: $C_2H_{6(g)}$: $S = 229.6\ J\ K^{-1}\ mol^{-1}$
   Example: $CH_{4(g)}$: $S = 186.3\ J\ K^{-1}\ mol^{-1}$

3) In ionic solids, the weaker the ionic charges, the higher the entropy.
   Example: $MgO_{(s)}$: $S = 26.9\ J\ K^{-1}\ mol^{-1}$
   Example: $NaF_{(s)}$: $S = 51.5\ J\ K^{-1}\ mol^{-1}$
   Example: $CH_{4(g)}$: $S = 186.3\ J\ K^{-1}\ mol^{-1}$

4) Entropy usually **increases** when pure liquids or solids dissolve in a solvent.

5) When dissolved gases escape from a solution, entropy usually **increases**.

6) By the second law of thermodynamics,

$$\Delta S_{universe} = \Delta S_{system} + \Delta S_{surroundings}$$

   Occasional localized decreases in entropy are acceptable as long as the overall entropy of the universe continues to increase.