You can experience directly the mass, volume, or temperature of a substance, but you cannot experience its entropy. Consequently you may have the feeling that entropy is somehow less real than other properties of matter. We hope to show in this section that it is quite easy to predict whether the entropy under one set of circumstances will be larger than under another set of circumstances, and also to explain why. With a little practice in making such predictions in simple cases you will acquire an intuitive feel for entropy and it will lose its air of mystery.

The entropy of a substance depends on two things: first, the *state* of a substance—its temperature, pressure, and amount; and second, how the substance is *structured at the molecular level*. We will discuss how state properties affect entropy first.

**Temperature** As we saw in the last section, there should be only one way of arranging the energy in a perfect crystal at 0 K. If $W = 1$, then $S = k \ln W = 0$; so that the *entropy should be zero at the absolute zero of temperature*. This rule, known as the **third law of thermodynamics**, is obeyed by all solids unless some randomness of arrangement is accidentally “frozen” into the crystal. As energy is fed into the crystal with increasing temperature, we find that an increasing number of alternative ways of dividing the energy between the atoms become possible. $W$ increases, and so does $S$. Without exception the entropy of any pure substance *always increases with temperature*.

**Volume and Pressure** $^{-1}$ *increases with increasing volume* $\text{decreases with increasing pressure* Amount of Substance* proportional to the amount of substance* extensive properties*}

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