

## **Entropy and the Economic Process**

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### **GLOSSARY**

**Entropy Change:** The amount of heat absorbed by a material divided by the material's absolute temperature.

**Isolated System:** A collection of materials which exchanges neither matter nor energy with its surroundings.

**Macroeconomics:** The study of the economy as an aggregated whole, generally ignoring the behavior of individual industries or consumers and instead concentrating on economy-wide magnitudes such as unemployment, inflation, interest rates, the trade deficit, and growth of gross national product.

**Microeconomics:** The study of the behavior of one or a few firms, industries, or groups of consumers in an economy.

**Natural Resources:** Exhaustible resources such as coal, renewable resources such as fish, and ecosystem capacities useful to humans, such as the ability to cleanse dirty water.

**Neoclassical Economics:** The school of economic thought which postulates that consumers act in a way that maximizes their well-being subject to their income and possibly other constraints, that firms act in a way that maximizes their profits subject to the technological constraints of production, and in general that economic agents act in a way that maximizes an objective function which can in principle be rigorously and completely mathematically described, subject to constraints which can also in principle be so described.

**Second Law of Thermodynamics:** Also known as the Entropy Law, it states that in an isolated system, entropy cannot decrease; it can either rise or remain constant.

**Neoclassical Theory of Value:** The theory that market prices are determined by the intersection of demand curves and supply curves, which in turn arise from the behavior of self-interested atomistic economic agents; these “curves” may be more complicated than simple functions over Euclidian space.

## I. Entropy

In order to explain why a capitalist economy produces certain commodities, and at what price those commodities are exchanged, a description must be given of consumers’ desire and ability to buy the commodity, and of how firms produce that commodity. The description of how firms produce a commodity is in many ways a description of the engineering opportunities open to the firm, although describing the allocation and supervision of labor is also important, as is the firm’s access to diverse forms of financing. In describing the engineering opportunities open to the firm, economists usually have to work with a simplified version of all the potentially available technologies in order for the problem to be tractable. However, it is important for the simplifications to obey fundamental physical laws, or at least not flagrantly violate them, so economists studying production should have a basic knowledge of those laws. The writers who first emphasized the relevance of the laws of thermodynamics to production economics were Frederick Soddy (the 1921 Nobel laureate in Chemistry), Nicholas Georgescu-Roegen, and Herman E. Daly. The Second Law of Thermodynamics, called the Entropy Law, has been the focus of most of the work in this field.

The first law of thermodynamics states that in an isolated system, energy is neither created nor destroyed. This law cannot explain why an ice cube placed on a sidewalk on a hot summer day always melts, because the first law would permit heat to flow from the ice cube as long as an equal amount of heat flowed to the air. Entropy is the concept used to explain why ice cubes on a hot sidewalk melt, and in general why heat always flows from hotter objects to colder ones and why spontaneous physical processes are spontaneous.

The definition of entropy is straightforward. Let  $T$  stand for absolute temperature, that is, temperature in degrees Kelvin, denoted K. (Kelvin temperature is 273.15 degrees higher than Celsius temperature.) Let  $Q$  stand for the flow of heat into a material (or system, which is a collection of materials). The metric unit of heat  $Q$ , as well as that work, is the Joule, denoted ‘J’; the English unit is the calorie or the BTU. Then the change in the entropy “ $S$ ” of the material (or system) is defined to be

$$dS = dQ_{\text{rev}}/T, \tag{1}$$

where the subscript “rev” stands for a reversible process; physicists call a process “reversible” if it involves no dissipative effects such as friction, viscosity, inelasticity, electrical resistance, or magnetic hysteresis [Zemansky 1968 p. 193, p. 215; Mackowiak 1965 p. 59]. As Zemansky adds [ibid. p. 225], if a system undergoes an irreversible process between an initial equilibrium state  $i$  and a final equilibrium state  $f$ , the entropy change of the system is equal to the integral from  $i$  to  $f$  of  $dQ/T$ , taken over any reversible path from  $i$  to  $f$ . No integral is taken over the original irreversible path.

Suppose one has an isolated system containing two bodies, one hot and one cold, placed in thermal contact. When a given quantity of heat,  $Q_0$ , flows from the hotter body to the colder one, the change in entropy of the system is

$$\Delta S = \frac{-Q_0}{T_{\text{hot}}} + \frac{+Q_0}{T_{\text{cold}}}.$$

Since  $T_{\text{hot}} > T_{\text{cold}}$ ,  $\Delta S > 0$ . If heat were to flow in the opposite direction, away from the colder body and toward the warmer body, then the  $Q_0$  terms in the above equation would change sign and  $\Delta S$  would be negative. Because it is the former and not the latter which is always observed, nineteenth-century physicists postulated that in an isolated system, entropy never decreases. This is one form of the second law of thermodynamics.

## II. The Laws of Thermodynamics

The laws of thermodynamics are:

Zeroth Law: Two systems which are in thermal equilibrium with a third are in thermal equilibrium with each other [Dugdale 1996 p. 13].

First Law: During a process in which no heat is exchanged with the environment, the work done is only a function of the initial and final states of the system, not of the path. Furthermore, during *any* process, heat flow,  $Q$ , is equal to  $U_f - U_i + W$  where  $U_f$  and  $U_i$  are the final and initial internal energies of the system and  $W$  is the net work done by the system [op. cit. p. 20; Zemansky pp. 78–79].

Second Law, Kelvin-Planck statement: No process is possible whose *sole* result is the absorption of heat from a reservoir and the conversion of this heat into work [Zemansky 1968 p. 178].

Second Law, Clausius statement: No process is possible whose *sole* result is the transfer of heat from a cooler to a hotter body [op. cit. p. 184].

Second Law, entropy statement: In an isolated system, entropy is nondecreasing [op. cit. p. 234].

Third Law, Unattainability statement: It is impossible to reach absolute zero by any finite number of processes [op. cit. p. 498; Dugdale 1996 p. 177].

Third Law, Nerst-Simon statement: In the limit as temperature goes to zero degrees Kelvin, the entropy change of any reaction is zero [Zemansky 1968 p. 498; Rao 1985 p. 257; Dugdale 1996 pp. 160–161].

Proofs that the three statements of the second law are equivalent can be found in Zemansky’s text, as can proof that the two statements of the third law are equivalent.

### III. Economic Implications of Entropy to Production Theory

**A. Heat Engines.** The first application of entropy was to “heat engines”; these are devices, such as internal combustion engines and power-plant steam engines (in both conventional and nuclear power plants), which use temperature differences to do their work. (A car’s motor, for example, will not work in a room hotter than the spontaneous ignition temperature of gasoline.) The “efficiency” of a heat engine is defined to be its work output divided by its heat input. A surprising and important implication of the Second Law of Thermodynamics is that the maximum efficiency of any heat engine operating between the temperatures  $T_{\text{low}}$  and  $T_{\text{high}}$  is  $1 - (T_{\text{low}}/T_{\text{high}})$ . This is less than 1 because  $T_{\text{low}}$  is strictly greater than zero due to the unattainability statement of the third law. A felicitous explanation of the maximum-efficiency equation appears in a book of Sir Oliver Lodge [Energy, 1929 pp. 39–40], who also explains what it has to do with entropy:

If the quantity of heat energy  $Q$  is lost by a body at temperature  $T$ , and gained by a body at temperature  $T'$ , then the entropy lost is  $Q/T$ , and the [entropy] gained is  $Q/T'$ . The gain is greater than the loss, if  $T'$  is less than  $T$ . But if some of the heat is utilised,—converted into mechanical or other forms of energy by means of some kind of engine,—so that only  $Q'$  is imparted to the colder body (the work done corresponding to the difference  $Q - Q'$ ) then it may be that  $Q/T = Q'/T'$ ; and in that case the entropy of the system remains constant. This is the condition for a perfect engine conducting a reversible operation. Irreversible operations increase entropy or dissipate energy, for the lost availability cannot be recovered. The efficiency of an engine means the ratio between the portion of heat utilised and the heat taken from the source. It is  $(Q - Q')/Q$ ; and we see now that in a perfect engine this would be  $(T - T')/T$ , because the  $Q$ ’s and the  $T$ ’s are then proportional.

Lodge’s ‘perfect’ heat engine is called a “Carnot engine”; on a temperature-vs.-entropy diagram, the cycle of this engine traces out a rectangle, called a “Carnot cycle.”

This Second-Law limit on the maximum efficiency of heat engines is striking because it establishes a boundary beyond which technological progress

will never cross. This has the salutary effect of throwing scientific doubt on the predictions of technological Panglossians who believe that technical progress is unbounded and can eventually remedy any problem facing industrial civilization. On the other hand, not all engines are heat engines (fuel cells, for example, are not), and the Second-Law maximum-efficiency limitation only applies to heat engines.

All engines, and indeed all uses of energy, are however subject to another limitation due to the entropy law. Lord Kelvin’s “principle of the degradation of energy” states that whenever an irreversible process takes place, the effect on the universe is the same as that which would be produced if a certain quantity of energy were converted from a form in which it was completely available for work into a form in which it is completely unavailable for work. This amount of energy is the temperature of the coldest heat reservoir at hand times the entropy change of the universe brought about by the irreversible process [Zemansky 1968 pp. 236, 237, 239]. As a consequence of this Principle, it is never desirable to recycle energy (because accomplishing such recycling requires using up more energy than would be regenerated).

**B. Chemical Processes.** In order to discuss implications of entropy to chemical rather than mechanical processes—for example, if one wished to analyze fuel cells—it is useful to rewrite (1) as

$$\Delta S = \int_i^f \frac{C_p(T)}{T} dT \quad (2)$$

where ‘ $i$ ’ denotes some convenient initial reference state, ‘ $f$ ’ denotes the state of interest, and  $C_p$  denotes the ‘heat capacity at constant pressure’ of the chemical substance. (The “heat capacity” of a substance (also known as its “specific heat”) is the amount of heat or work needed to raise the temperature of one gram of the substance by one degree Kelvin (or Celsius). A substance’s heat capacity changes slightly depending on its initial temperature.) In a laboratory, (2) can be approximated to any degree of accuracy by replacing the integral with a sum, as in

$$\Delta S \approx \sum_{j=i}^f \frac{C_p(T_j)}{T_j} (1 \text{ K}) \quad (3)$$

where the temperature range from  $T_i$  to  $T_f$  has been broken down into  $j$  parts. Using (3), the entropy of one gram of a substance is experimentally determined in the following way: Start with the substance at a convenient initial temperature and pressure. Measure the heat needed to raise its temperature from  $T_1$  to  $T_2$ , then divide this figure by  $T_2 - T_1$  to obtain the specific heat (at constant pressure). Divide this specific heat by the initial

temperature; this is  $C_p(T_1)/T_1$ . Next measure the heat needed to raise the temperature from  $T_2$  to  $T_3$ , and divide that by  $T_3 - T_2$ , then divide that by the temperature; this is  $C_p(T_2)/T_2$ . Continue until  $T_j$  is reached. The change in entropy of the substance between the initial and final states is approximately equal to the sum of the  $C_p(T_j)/T_j$  terms, times one degree Kelvin to adjust the units.

In this way, the entropy change of different materials can be experimentally determined. The results of such experiments are given—for example, in freshmen chemistry textbooks—in tables of entropies for many substances. For instance, the standard entropy of one mole of copper (63 grams) is  $33 \text{ J/K} = 8 \text{ cal/K}$ , where ‘cal’ denotes calories [Rao 1985 Table C-1], and where “standard” (or “absolute”) entropies are defined with an initial condition of 0 K and one atmosphere pressure and a final condition of 300 K and one atmosphere pressure.

Armed with such knowledge, and the Second Law’s principle that entropy in an isolated system never decreases, the equilibrium proportions of different chemicals reacting with each other can be determined merely by calculating, on paper, which proportions maximize the entropy of the mixture. In other words, once entropy values have been tabulated, the equilibrium direction and extent of any chemical reaction can be determined without further laboratory work. There is a caveat: while the *equilibrium* direction and extent of any chemical reaction can be determined, not all chemical systems go to equilibrium; some get stuck in a “metastable” state in which an “activation energy” barrier blocks their path to the equilibrium. This barrier can be broken by a catalyst, which works by supplying the activation energy, which is then returned in full to the catalyst.

**C. General Processes.** The Entropy Law thus governs the efficiency of heat engines, the recycling of energy, and the direction and extent of chemical reactions. More generally, it sets the ground rules within which mechanical and chemical systems operate. Even more generally, it constitutes a fundamental principle of irreversibility, in the following sense: Imagine a system which is isolated. By the Second Law, entropy always increases in such a system (it could remain the same, but only if there were no friction, only “perfect” machines, and so forth). Choose two dates  $t_1$  and  $t_2$  with  $t_1 < t_2$ . Then the entropy of the system at  $t_1$  (call it “ $S_1$ ”) is less than the entropy of the system at  $t_2$  (call it “ $S_2$ ”):  $S_1 < S_2$ . Then it is in principle impossible to reverse all of the changes which have happened in this isolated system between  $t_1$  and  $t_2$ —assuming the system remains isolated—because such a reversal would imply a decrease of entropy. This means that perpetual motion machines are impossible. It means that complete recycling of all a process’s components is impossible; some fresh inputs will always be needed for the process to occur again. *A fortiori*, the economy will always

require natural resource input of some sort from nature. The economy will never be a perpetual motion machine.

Nicholas Georgescu-Roegen wrote that “the economic process has a uni-directional irrevocable evolution” (from low to high entropy), and in fact all processes have such an evolution. The activities of an isolated mine, factory, or household—activities which ultimately turn valuable inputs into waste products—are all physically irrevocable. As Georgescu-Roegen and Herman Daly have emphasized, this replaces the typical freshman textbook conceptualization of the economy as a circular flow with an alternative conceptualization of the economy as entailing a one-way flow of matter and energy from usefulness to uselessness.

The type of economic growth occurring over the last four centuries has greatly accelerated this flow towards the ultimate maximum-entropy equilibrium from which no change could ever be made [Georgescu-Roegen 1971/1993 p. 81].<sup>1</sup> In speeding up this final end, modern economic growth has produced not just more waste but also wonderful “dissipative structures” which, as we learn from the work of Ilya Prigogine, can be the vehicles through which entropy is transformed from low to high. “Dissipative structures” include machines, human beings, and amazing new technologies like R. Dean Astumian’s [2001 p. 60] “molecular motors,” which function precisely because they are not in thermodynamic equilibrium and can use the resulting chaotic turbulence to do work.

Important as the conceptualization of the economy as a one-way process always dependent on fresh inputs from nature is, two caveats are in order. The first is that the word “isolated” is an important qualifier. A farm, for example, is not thermodynamically isolated from the Sun. To construct a proper isolated system for agriculture, the Sun has to be included as part of the system. The second caveat then comes into play: while a system as a whole cannot be returned to its previous state, a component of the system—such as a farm’s soil—can be returned to its previous state, as long as some other component of the system (such as the Sun) suffers an increase in entropy sufficient to offset the decrease in entropy of the component which has returned to its previous state. In principle, this means it may be possible to perfectly recycle the materials of any production process which involves or could be made to involve the Sun (until the Sun dies).

Nicholas Georgescu-Roegen believed such perfect matter recycling to be impossible. Since such an impossibility does not follow from any of the scientifically accepted laws of thermodynamics, he proclaimed this impossibility as a new “Fourth Law of Thermodynamics.” Ecologists often believe that nutrient recycling in ecosystems is 100 per cent complete, and thus

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<sup>1</sup>The fact that the universe as a whole has such an evolution led Kelvin to the idea of the “heat death of the universe.” See <http://scienceworld.wolfram.com/biography/Kelvin.html>.

dispute the Fourth Law’s validity. By contrast, economics Nobel laureate Paul Samuelson [p. xv and xvii, 1999] wrote that Georgescu-Roegen’s “Law of Inevitable Dissipation of Useful Concentrated Matter [...] is good-sense certainty [...]: only if all useful copper and gold and helium could be 100 per cent recycled would a *perpetuum mobile* be possible. If pigs could fly, if my Aunt Sally were a stagecoach, if ... and if ... .” So the truth of the Fourth Law remains in dispute.

As an aside it should be mentioned that not only does  $t_1 < t_2$  in a closed system imply  $S_1 \leq S_2$ , but  $S_1 < S_2$  in a closed system also implies  $t_1 < t_2$ . The second statement has been used to explain the direction of time, “time’s arrow.” For example, it has been proposed that since we live in a universe whose expansion causes the universe’s entropy  $S$  to change, the reason we perceive time to increase in the direction it does is that that causes us to conclude that the universe’s  $S$  is increasing, in accordance with the Second Law of Thermodynamics, not decreasing, which would violate the Second Law.

#### IV. Entropy and Neoclassical Economic Analysis

Neoclassical microeconomic theorists have no difficulty incorporating scientific constraints into their work. Indeed, as far back as Tjalling Koopmans’ and Gerard Debreu’s work of the 1950’s [Debreu 1959 p. 40, (d); Koopmans 1951 p. 48], general equilibrium theorists have explicitly incorporated an irreversibility assumption into their work, as well as a more basic assumption that there cannot be outputs without inputs. Unfortunately, general equilibrium analysis is conducted at a level of abstraction that makes it almost useless for policy applications. (For example, “applied general equilibrium” analysis of trade policy *assumes* full employment, which begs the question of free trade’s effects on employment.)

Neoclassical analysis of long-run economic growth has been conducted, not using microeconomic general equilibrium, but instead using a macroeconomic framework in which aggregate economic output is assumed to be a function  $f(L, K)$  of the economy’s stock of labor  $L$  and physical capital  $K$ , or sometimes a function  $f(L, K, R)$  including natural resources  $R$ . Inspired by the work of Robert U. Ayres, one notices that elementary mass balance (conservation of matter) is lacking in  $f(L, K)$ —for where does  $K$  come from? Surely not from  $L$ , and most physical capital nowadays is not alive and able to reproduce itself!—and it is also lacking in  $f(L, K, R)$  in the common formulation where  $K$  can be maintained at a constant or even increasing level while  $R$  goes to zero.

Neoclassical resource economists have constructed a very useful body of microeconomic knowledge about how profit-maximizing firms exploit the environment, as well as the implications of different types of government



regulation of that activity. However, these results usually do not treat the overall issue of natural resources in long-run growth.

Fundamentally, no economic technique will be able to definitively foretell the abundance and accessibility of particular natural resources far in the future, nor the need for those resources in the future. Educated guesses are possible, but novelty, contingency, and lack of predetermination mean that it is not possible to predict the future. There is no basis to confidently assume that the future will resemble the recent past few centuries, in which technological progress made it possible and desirable to exploit vast, hitherto useless stocks of many different materials found in nature.

## V. Interpretations and Misinterpretations of Entropy

Equations (1), (2), and (3) are completely straightforward in their literal meaning. However, they give no intuitive insight into what entropy is. Other famous formulas in physics, such as  $F = ma$  or even  $E = mc^2$ , deal with concepts we have at least some everyday familiarity with or at least understanding of, such as force, mass, acceleration, energy, and the speed of light. By contrast,  $Q/T$  or  $C_p/T$  are remote and hard to grasp. This has led scientists to come up with “interpretations” or “explanations” of entropy. Some of these interpretations are extremely valuable to scientific specialists in the field. Others are unfortunate misinterpretations. Neither the useful ones nor the useless ones have much importance to economists, but their continuing great influence on the historical development of the field makes it important to understand them.

**A. Disorder, Randomness, and Probability.** One of the most surprising consequences of equation (1) comes from analyzing the mixing of two ideal gases in a container whose walls allow no heat to flow. (An “ideal” gas has molecules which do not interact with their neighbors.) At first sight, the formula textbooks give for the resulting change of entropy—which is

$$\Delta S = -R \left[ n_1 \ln \frac{n_1}{n_1 + n_2} + n_2 \ln \frac{n_2}{n_1 + n_2} \right] \quad (4)$$

where  $n_1$  and  $n_2$  are the moles of the gases and  $R$  is the universal gas constant (8.31 J/(K · mole)—seems wrong, because since no heat can flow into or out of the container, it seems from (1) that the system’s entropy cannot change. This doubt about (4)’s validity is incorrect. Entropy change is not  $\int_i^f dQ/T$  (which is zero in this case) but rather  $\int_i^f dQ/T$  over a reversible path. The mixing of two ideal gases is not reversible. The only way to calculate the entropy change accompanying the mixing of two ideal gases is to replace the original irreversible path linking the states before and after the mixing with a reversible path linking those same states, then performing the integration. Because entropy change is a function of the initial and final states only,

entropy change along the new reversible path (which can be calculated using (1)) must be the same as the entropy change along the original irreversible path. Along the reversible path, heat does flow, and the entropy change turns out to be (4), which is greater than zero.<sup>2</sup>

The remarkable result that mixing two gases increases entropy has led observers to think that there is more to the idea of entropy than expressed by (1)—that entropy measures increasing mixing, or spatial disorder, or even some general notion of “order,” *per se*. This is incorrect. Increasing entropy does *not* mean increasing disorder in any intuitive, lay-person sense. For example, the spontaneous—i.e., entropy-increasing—evolution of an oil/water mixture is towards “oil on the top, water on the bottom,” which is hardly increasing disorder according to any intuitive meaning of the word “disorder.”

Confusing the issue is that while increasing entropy does not mean increasing disorder in any intuitive, lay-person sense, it does mean increasing disorder in an abstract, highly technical sense defined by the Boltzmann equation

$$S = k \ln W + S_0, \quad (5)$$

where  $k$  is Boltzmann’s constant ( $1.38 \times 10^{-23}$  J/K) [Rao p.258], where the constant  $S_0$  is, by tradition but not by necessity, set equal to zero, and where  $W$  is defined by (6), where  $N$  is the number of particles in the system and  $n_i$  is the number of particles in energy level  $i$ :

$$W = \max_{n_0, n_1, \dots, n_k} \frac{N!}{n_0! n_1! \dots n_k!} \quad \text{such that} \quad (6)$$

$$\sum_i n_i = N \quad \text{and} \quad (7)$$

$$\sum_i n_i \cdot (\text{energy level } i) = \text{total system energy}. \quad (8)$$

(Note that  $N!/(n_0!n_1! \dots n_k!)$  is the number of ways that  $N$  distinguishable objects can be put into groups of  $n_0, n_1, \dots, n_k$  objects.)<sup>3</sup>

While  $W$  defined in (6) is not a random variable, and comes from combinatorics not statistics, Boltzmann christened it the “thermodynamic probability,” which has led observers to incorrectly think that entropy has something to do with randomness. On the other hand, there is another interpretation of Boltzmann’s ideas which does involve randomness. In this interpretation, the “max” operator in (6) is dropped, and  $W$  becomes a random variable which takes on the value  $N!/(n_0!n_1! \dots n_k!)$  with probability

$$\frac{N!/(n_0!n_1! \dots n_k!)}{\sum_{n_0, n_1, \dots, n_k} N!/(n_0!n_1! \dots n_k!)}$$

<sup>2</sup>See Beard and Lozada [1999 pp. 86–88.]

<sup>3</sup>An excellent derivation of (5) is given by Dugdale [1996 pp. 70–99].)

(see Dugdale p. 73, 76). The problem with this interpretation is that it allows the Second Law to be violated with nonzero probability, and such a violation has never been observed.

Boltzmann's ideas have grown into the important science of statistical mechanics, by which the entropies of various substances are *explained*, from a molecular level; the older classical thermodynamics described how to measure entropy, but could not explain why it should take on a particular value for a particular substance. For extremely simple physical systems, statistical mechanics does imply that lower entropy states are those which are intuitively more ordered and less random. Thus one mole of a substance in the form of a simple solid has lower entropy than if that mole is in the form of a simple liquid, and that in turn has lower entropy than if that mole is in the form of a simple gas. Also, random mixtures of simple substances have higher entropy than if those substances are ordered into unmixed, pure components, as seen in (4). However, most substances are not "simple" in the sense necessary for this intuition to be valid, as the oil and water example showed. Treatments which ignore this caveat—such as Erwin Schrödinger's famous 1944 essay "What is Life"—need to be read with the greatest caution.

**B. Absolute Entropy.** Notice that (1) defines only changes in entropy; absolute amounts of entropy are not defined. It is sometimes asserted that, by contrast, in Statistical Mechanics absolute entropy is defined—equivalently, that "the constant  $S_0$  in (5) must be zero." Modern treatments show that this is not true [Dugdale 1996 p. 99; Max Planck's erroneous treatment of this issue is refuted in Beard and Lozada 1999 p. 118 fn. 12]. This controversy has no bearing on laboratory entropy measurements, which always measure entropy differences anyway. (*A fortiori*, it has no bearing on processes whose thermodynamic analyses use or could use those laboratory measurements, so the point is one of mathematical and theoretical correctness rather than scientific or applied correctness.)

**C. Thermal vs. Configurational Entropy.** Occasionally, scientists find it useful to distinguish the contribution to entropy change represented by (4) from other contributions to entropy change. When this is done, the component of the entropy change given by (4) is called "configurational entropy" or "the entropy of mixing," and the other components of the entropy change are together called "thermal entropy." The choice of the term "thermal entropy" is unfortunate, because it can mislead one into thinking that configurational entropy has nothing to do with thermal changes, and hence that configurational entropy has nothing to do with (1). Configurational entropy is actually just as much a component of (1) as thermal entropy is.

**D. Entropy and Free Energy.** To calculate the entropy change when an event occurs, the resulting entropy change of the system's surroundings

must be added to the entropy change undergone by the system itself. Under conditions of constant pressure and temperature, the former is “the energy released by the system into its environment” (traditionally represented by  $-\Delta H$ ) times  $1/T$ , and the latter is  $\Delta S$ . Adding these two entropy changes together yields a total entropy change of  $\Delta H/T - \Delta S = (1/T)(\Delta H - T\Delta S) = (1/T)\Delta G$  where  $\Delta G$  is known as the “Gibbs Free Energy.” Tables of Gibbs Free Energies of common substances are readily available, and such a table is all one needs to determine how entropy changes when an event occurs under conditions of constant pressure and temperature. This makes Gibbs Free Energy extremely valuable for chemists, but it should not obscure the fact that entropy is the fundamental physical quantity, and that Gibbs Free Energy is a derived quantity of no independent physical importance. Similarly, Helmholtz Free Energy  $F = U - TS$  (where  $U$  is internal energy) is useful in situations where volume and temperature  $T$  are constant, but it is not a fundamental quantity either.<sup>4</sup>

**E. Information.** An equation which looks like (5) with (6) appears in the field of information theory, with of course entirely different meanings for the symbols. This does not imply any connection between entropy and information, any more than the similarity between the formula for the sum of scores on two student examinations ( $e_1 + e_2$ ) and the formula for the sum of the energy of two hotel elevators ( $e_1 + e_2$ ) implies any connection between student examination scores and hotel elevators. The “entropy” of information theory, while no doubt quite useful to practitioners of that field, has nothing to do with the Second Law of Thermodynamics.<sup>5</sup>

**F. Minimum Energy Requirements.** Near the end of Section III (A) it was shown that the amount of energy  $E$  degraded during an irreversible process is the temperature of the coldest heat reservoir at hand (call it  $T_0$ ) times the entropy change of the universe brought about by the irreversible process:

$$E = T_0 \Delta S. \quad (9)$$

Suppose we desire to obtain pure copper from oxide copper ores of the type often found in Africa. Because CuO is thermodynamically stable at room temperatures,<sup>6</sup> breaking up CuO into copper and oxygen would cause an entropy decrease, so it cannot occur spontaneously. Some authors have

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<sup>4</sup>See also Beard and Lozada [1999 Section 5.5 and footnote 10] and Lozada [1999, Appendix].

<sup>5</sup>Much has been written on this topic; for example, see Beard and Lozada [1999 pp. 109–111.]

<sup>6</sup>If  $dS$  is the entropy change,  $dH$  is the enthalpy change,  $dG$  is the Gibbs Free Energy change, and  $T$  is absolute temperature, it can be shown that  $dS = (dH - dG)/T$ , so  $dG = dH - TdS$ . A process is spontaneous if  $dH < 0$  and  $dS > 0$ , making  $dG < 0$ . In order for CuO to be stable, we would want  $\text{CuO} \rightarrow \text{Cu} + 0.5\text{O}_2$  to have a positive  $dG$ , which it does from King et al. [1973] Table 30.

calculated the amount of the entropy decrease, then used (9) to assert that  $T_0$  times that amount is a minimum amount of energy necessary to liberate the copper from the CuO. (Their general proposition is that there is a definite theoretical limit to the minimum energy input to a process and that there is no substitute for this minimum energy.)<sup>7</sup> There may be some sense in which this is true, but no amount of an “energy commodity” (such as coal or electricity) is needed to obtain copper from CuO. What is done is simply to introduce sulfuric acid into the environment of the copper oxide [Biswas and Davenport 1980 pp. 254–5, 258]; a typical reaction is  $\text{CuO} + \text{H}_2\text{SO}_4 \longrightarrow \text{Cu}^{++} + \text{SO}_4^{-} + \text{H}_2\text{O}$  [Biswas and Davenport 1994 p. 15], which increases entropy and is thus spontaneous. This example shows that there is no minimum “energy commodity” requirement for physical, chemical, and hence industrial processes.<sup>8</sup>

**G. Entropy and Energy.** There is no such thing as ‘low-entropy energy.’ Entropy is a characteristic of matter, and a substance’s entropy changes when the substance reversibly absorbs or emits energy (or when an irreversible process involves the substance and a reversible process linking the same initial and final states would entail the substance’s absorption or emission of energy); see (1).

**H. Entropy Theory of Value.** A glance at a table of standard entropies for different substances will show the error of believing that “the lower an object’s entropy, the higher its economic value.” Similarly, each time one heats a kettle of water in a kitchen, the water becomes more useful at the same time as its entropy increases. Adding impurities to iron to create steel simultaneously increases the entropy of the metal and increases its economic value. The entropy of oceanic pollutants is decreased by their concentration in the flesh of oysters, but this decrease of entropy makes the oysters unsafe to eat, thus less valuable to humans. So there is no general connection between low entropy and high economic value, nor between high entropy and low economic value. It is true that the process of production and consumption turns valuable inputs into useless waste, with the useless waste having, as a total assembly, higher entropy than the valuable inputs. However, this generalization only applies to assemblies as a whole, and not to their component parts.

Furthermore, what is “valuable” will always be partially determined by psychological, historical, sociological, and in any case non-physical traits unrelated to entropy; a personal computer made in 1987 was quite valuable then, but would be considered junk now, even if its physical traits are unchanged. Even a simple container of oxygen, whose entropy decreased when

<sup>7</sup>See Chapman and Roberts [1983, pp. 10, 24–25].

<sup>8</sup>Conversely,  $\text{Cu}_2\text{S}$ , from which copper is often made in North America, is thermodynamically metastable. Obtaining copper from  $\text{Cu}_2\text{S}$  results in a net output of energy; a typical reaction is  $\text{Cu}_2\text{S} + \text{O}_2 \longrightarrow 2\text{Cu} + \text{SO}_2 + \text{energy}$ . See Beard and Lozada [1999 (5.8)].

it was extracted from air, can be “useful” and “valuable” in one situation—helping hospital patients breathe—and be an unwelcome hazard in another situation, such as near a stove.

Suppose we agree to draw process boundaries large enough so that the process’s components form a closed system. Then, as mentioned in Section IIC, *all* processes increase (or leave unchanged) entropy, regardless of whether the process creates economic value or destroys it.

## VI. Conclusion

The Entropy Law implies that economic production will never be entirely self-contained or self-replicating, let alone self-expanding. Production will always require fresh inputs from nature, and nature’s size, both of terrestrial stocks and of the rate of solar flow, is finite.

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