Entropy Can Decrease

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Abstract

The second law of thermodynamics, which asserts that entropy never decreases, is often regarded as a universal principle. However, this assumption merits re-examination. While entropy generally increases in systems dominated by repulsive forces, it decreases in domains governed by attractive forces such as gravity. This explains why the universe has not succumbed to “heat death” but has instead evolved into a greater organized structure, as stars and galaxies continuously aggregate from diffuse cosmic matter. The conventional view relies on the assumption that energy flows, such as heat transfer, always follow a decreasing gradient of temperature. This assumption does not necessarily hold in systems under attractive forces. Moreover, the second law describes only directional progression, whereas many natural processes are oscillatory or periodical. These observations suggest that entropy increase accounts for only a subset of natural phenomena. To address these limitations, we propose the **Restoration Principle**, which unifies entropy-increasing, entropy-decreasing, and oscillatory processes. Everything in the universe exists in different forms of energy, and all processes in the universe involve transformations among various forms of energy, driven by fundamental attractive or repulsive forces that aggregate or disperse components toward a stable state with a balanced energy distribution. When this balance is disturbed, systems spontaneously act to restore it. The Restoration Principle thus offers a more comprehensive framework for understanding natural evolution than the entropy-based view.

Introduction

The second law of thermodynamics states that the total entropy of an isolated system cannot decrease.[1] A popular application of this principle is that an ice cube in a cup of water at room temperature will melt spontaneously, distributing energy more evenly and increasing the system’s overall entropy. The reverse process, where liquid water spontaneously reassembles into an ice cube without any external input, does not occur, as it would require a decrease in entropy, violating the second law.

The tendency of entropy to increase in spontaneous processes is observed across numerous domains, making it widely regarded as a universal principle of evolution. Originally recognized in classical thermodynamics, the concept of entropy has since been extended, through statistical mechanics, to a broad range of disciplines, from the microscopic realm of statistical physics to the foundations of information theory. Its applications span chemistry, physics, and biology, as well as fields such as cosmology, economics, sociology, meteorology, and climate science.

In its statistical interpretation, entropy is most often associated with disorder, randomness, or uncertainty, and stands as a cornerstone for understanding the direction and behavior of natural processes. Unlike many physical quantities that are symmetric in time, entropy is defined by its one-way increase, a feature often described as the “arrow of time”. This concept has been extended to predict the ultimate fate of the universe: a cold, dark, and inert cosmos in which energy is uniformly dispersed and entropy reaches a maximum—the heat death.

However, accumulating cosmological evidence shows that the universe has not been progressing toward heat death, even after more than 13 billion years of evolution. Instead, it has become increasingly structured: new stars and galaxies continue to form from cosmic clouds rather than dispersing into a uniform sea of particles. Why, then, does the universe not evolve as the second law of thermodynamics might predict? The answer is not that cosmic evolution has yet to reach a mature stage when this tendency manifests, but that the second law is not the universally fundamental principle it is often presumed to be.

The second law of thermodynamics faces several challenges. First, entropy definitions have not captured the underlying nature of all entities in the universe and therefore cannot account for every process it encompasses. Second, the thermodynamic and statistical definitions are not fully equivalent, leading to contradictions in the interpretation of entropy changes, for example, in cases where entropy appears to decrease. Third, the thermodynamic definition relies on temperature as an input variable; because of the long-standing misconception that temperature represents simply the average kinetic energy of particle motion, this understanding is inherently limited and cannot fully capture the complexity of various processes in the universe.

More fundamentally, everything in the universe exists in various forms of energy, and the evolution of any system involves the transformation and redistribution of that energy. Such processes may progress toward or restore from an energy balance, or oscillate around it. The first two scenarios typically follow a directional course in which entropy can either increase or decrease—yet the second law of thermodynamics cannot adequately account for both possibilities. Furthermore, the second law provides no framework for describing oscillatory processes.

To overcome these limitations, a broader theoretical framework is needed—one that accounts not only for entropy-increasing and entropy-decreasing processes but also for oscillatory behaviors. The **Restoration Principle** provides such a foundational framework for system evolution, encompassing both directional change and oscillatory behavior across all types of processes.[2] As such, the Restoration Principle applies universally across all domains and fundamental interactions. Within this broader framework, the second law of thermodynamics can be understood as a specific manifestation of this principle.

The Introduction of Entropy

In classical thermodynamics, the **entropy** change of a system is defined as an infinitesimal amount of heat transferred from the surroundings to the system during a reversible process divided by the instantaneous temperature of the system during the transfer:[3]

$$(1) \qquad dS = \frac{dQ}{T}$$

Here, *T* is the absolute temperature of the system, *S* denotes the entropy, *dQ* represents the heat transferred, and *dS* indicates the resulting change in entropy due to the heat transfer. This concept was introduced to express Carnot's efficiency in terms of a system property.[4] Carnot efficiency is defined as the ratio of work output to the heat input transferred to a system performing the work. The maximum possible efficiency is determined by Carnot's theorem:

$$\text{(2)} \qquad \frac{W}{Q_h} = \frac{Q_h + Q_c}{Q_h} = 1 - \frac{T_c}{T_h}$$

In this definition, W denotes the work done by the heat Q_h transferred from a hot reservoir at temperature T_h in a Carnot engine, while Q_c is the heat released to a cold reservoir at temperature T_c. The right portion of Equation 2 can be reorganized as:

$$\text{(3)} \qquad \frac{Q_h}{T_h} + \frac{Q_c}{T_c} = \Delta S_h + \Delta S_c = 0$$

Here, ΔS_h and ΔS_c represent the entropy changes of the hot and cold reservoirs, respectively, showing that the total entropy remains unchanged. In contrast, for a heat engine less efficient than a Carnot cycle, the work output would be less than the net input heat (i.e., $W < Q_h + Q_c$) due to energy losses such as friction. In this case, the total entropy increases:

$$\text{(4)} \qquad \Delta S_h + \Delta S_c > 0$$

The results above motivated the introduction of entropy as a meaningful system property for characterizing state changes. However, it is important to recognize that Expression 4 does not represent an entropy increase in the sense described in the second law of thermodynamics. Rather, it reflects an increase due to inefficiencies within the heat engine. In a Carnot cycle, energy is exchanged among four components: a hot reservoir, a cold reservoir, the Carnot engine, and the work-producing subsystem. Expression 4 specifically accounts for the entropy changes in two of these components—the hot and cold reservoirs. In contrast, the entropy increase described by the second law refers to the internal evolution within a single system as it progresses toward a balance.

Statistical Mechanics

Ludwig Boltzmann formulated a statistical definition of entropy by analyzing the microscopic behavior of a system's components. According to his definition, entropy is a logarithmic measure of the number of possible microstates of a system, where each state has a probability p_i of being occupied:[5-10]

$$\text{(5)} \qquad S =- k_B \sum_{i=1}^{n} p_i ln(p_i)$$

In this definition, k_B denotes the Boltzmann constant, and the summation is performed over all possible microstates of the system. To better understand this definition, consider a gas confined within a container. A microstate refers to a specific, detailed configuration of the molecules in the system. For simplicity, assume that all microstates are equally probable. If the system has a total of *n* possible microstates, then the probability of each microstate is $p_i = 1/n$. Under this assumption, the entropy defined in Equation 5 can be rewritten in terms of the probability 1/n or the number of microstates *n*:

$$S =- k_B \sum_{i=1}^{n} \frac{1}{n} ln(\frac{1}{n}) =- k_B n \frac{1}{n} ln(\frac{1}{n}) =- k_B ln(\frac{1}{n}) = k_B ln(n) \tag{6}$$

This expression reflects two closely related definitions of entropy: one as the logarithm of the expected probability (*1/n*) of a system's state, and the other as the logarithm of the number of accessible microstates (*n*). The first definition emphasizes entropy as a measure of the system's tendency to evolve toward its equilibrium, the most probable state, while the second highlights how entropy varies with the number of microscopic configurations.

For example, in a gas, as the temperature decreases, molecules condense into a liquid and settle toward the bottom of the container. This reduces their spatial freedom and limits the number of accessible microstates (*n*), leading to a decrease in entropy. With further cooling, the liquid solidifies, further constraining molecular motion and reducing entropy even more. At absolute zero, the system occupies a single ground state, and therefore, its entropy reaches zero.

The statistical perspective broadens the concept of entropy, linking it to ideas such as disorder, randomness, and uncertainty. For instance, the notion of randomness was further developed in information theory by Claude Shannon, where the information entropy (*H*) is simplified from Definition 5 as follows:[11-12]

$$H =- \sum_{i=1}^{n} p_i log_2(p_i) \tag{7}$$

Note that the Boltzmann constant is omitted in this definition, as it was originally introduced to align the entropy dimension of the statistical definition with its thermodynamic counterpart. Additionally, the natural logarithm (*ln*) is replaced with the base-2 logarithm (log_2), since information in computing is typically represented in binary form.

The statistical definition is what extends the concept of entropy, enabling its application across diverse fields such as physics, chemistry, biology, meteorology, climate science, information theory, cosmology, economics, and sociology. However, it is crucial to recognize that this extension makes sense only from a statistical perspective; applying it beyond this scope can lead to misinterpretation. This distinction highlights the limitations of the second law of thermodynamics. Mixing these two definitions is a common source of confusion when applying the concept of entropy.

Thermodynamic vs. Statistical Entropy

Although Boltzmann is often credited with establishing an equivalence between the thermodynamic and statistical definitions of entropy by analyzing the microscopic behavior of a system's components, the two remain fundamentally distinct, particularly when comparing thermodynamic Definition 1 with statistical Definitions 5 and 6. This divergence becomes especially apparent near absolute zero. At absolute zero, a system occupies a single accessible microstate (the ground state), which corresponds to zero entropy according to Definition 5 or 6. In contrast, the thermodynamic definition predicts a different outcome: as temperature approaches zero, its inverse diverges to infinity, causing the thermodynamic entropy under Definition 1 to also diverge toward infinity. This inconsistency highlights a fundamental disparity between the statistical and thermodynamic formulations of entropy.

Under the statistical definition, a system evolves toward a most likely distribution of its components, whereas under the thermodynamic definition, it moves toward a balanced distribution of its energy. For example, statistical entropy can describe the disorder of a checkerboard arrangement based solely on the randomness of piece placement—a measurement entirely outside the scope of the thermodynamic context. The thermodynamic definition characterizes a system's macrostate based on its energy parameters, while the statistical definition relates entropy to the number of accessible microstates. To establish a true equivalence between these two definitions, one must explicitly connect their input parameters, specifically, the system's energy and its corresponding number of microstates. Unfortunately, such a direct relationship remains elusive. As currently formulated, the statistical definition is too permissive to provide a straightforward and reliable path to the thermodynamic formulation.

Ambiguity presents another significant challenge for the statistical definition of entropy, as the concept of a *microstate* lacks a precise and universally accepted definition. Mathematically, for example, both a 1-meter and a 10-meter line segment contain the same number of points—an uncountably infinite set. This means that a volume of 1,000 cubic meters does not contain more points than a volume of 1 cubic meter. Therefore, it is not strictly valid to claim that a gas molecule has access to more microstates in a larger container based solely on spatial extent. This conceptual ambiguity limits the practical applicability of the statistical definition of entropy, which is why chemists have to rely on the thermodynamic definition, grounded more firmly in measurable quantities, such as temperature and heat.

The perceived equivalence between the two definitions of entropy may arise from the microscopic behavior of gaseous systems, which likely influenced Boltzmann's reasoning. In a gas where particle dynamics are governed primarily by electron–electron repulsion, molecules naturally disperse. As temperature increases, the rise in molecular kinetic energy leads to a broader spatial distribution of particles. Within the statistical framework, greater dispersion corresponds to higher entropy; thus, elevated temperatures are associated with increased entropy in gaseous systems according to the statistical definitions (5) or (6).

For gaseous systems, it can be demonstrated that high temperature is likewise associated with high entropy under the thermodynamic definition. To illustrate this relationship, consider a gas of mass (*m*) with a constant specific heat capacity (*c*) over the relevant temperature range. Under these assumptions, the entropy from the thermodynamic definition (1) can be expressed as:

$$(8) \qquad dS = \frac{dQ}{T} = \frac{mcdT}{T}$$

Integrating this equation over a temperature range from T_0 to T, we found:

$$(9) \qquad S = mc\left(ln(T) - ln(T_0)\right)$$

When T_0 is taken as a fixed starting point, this equation indicates that the entropy of the gas is proportional to the logarithm of T, revealing a direct relationship between temperature and entropy. In this specific system, a parameter relationship can be established between the two definitions, linking higher temperatures to more dispersed molecular distributions and a greater number of accessible microstates. Through this connection, the thermodynamic view of entropy increasing with temperature aligns with the statistical view of entropy increasing with the number of microstates, suggesting a conditional equivalence between the two formulations in such a case.

Nevertheless, it is important to recognize that this apparent alignment arises from a specific physical system and cannot be extended to other contexts, particularly systems under the influence of attractive forces. Although the form of Expression 9 resembles Definitions 5 and 6, they differ fundamentally. Thermodynamic entropy depends explicitly on temperature and is therefore related to the energy level. In contrast, the statistical definition is based solely on the number of accessible microstates, which are only indirectly influenced by energy and temperature in certain systems, e.g., gas systems. As a result, this correspondence between the two definitions may hold in dispersive systems governed primarily by repulsive forces. However, this agreement breaks down in systems dominated by attractive forces. As will be shown later, when the statistical definition is applied to systems dominated by attractive forces, entropy often decreases, whereas thermodynamic entropy does not necessarily behave the same way.

Entropy Decrease Examples

Because the thermodynamic and statistical definitions of entropy are fundamentally distinct, they should not be used interchangeably. Applying one definition within the context of the other often results in confusion and misinterpretation. Specifically, when using the statistical definition, which depends solely on the number of accessible microstates, energy changes should neither be assumed nor imposed, as they are not input variables of the definition. Maintaining this distinction helps prevent misunderstandings and provides a clearer foundation for applying and analyzing the differences between the statistical and thermodynamic definitions.

To illustrate this difference, consider a simple experiment: an isolated system in which water at 0 °C is mixed with ice initially at a temperature below 0 °C. Some of the water may freeze as the ice warms to 0 °C. Under the statistical definition (6), the overall entropy of the system decreases, since the formation of additional ice reduces molecular freedom and decreases the number of accessible microstates. However, according to the thermodynamic definition (1), the overall system's entropy increases. Although the heat exchanged between the water and ice subsystems is equal in magnitude, the temperature of the ice is lower than that of the water. As a result, the entropy gained by the ice is greater than the entropy lost by the water, leading to a net increase in the system's entropy. This example shows both that entropy can decrease and that a single process can yield opposite entropy trends depending on the definition used.

In the ice–water example at room temperature, ice melts spontaneously, and the reverse process does not occur. In contrast, in the above example, water spontaneously assembles into ice. Superficially, the only difference between the two scenarios is temperature; fundamentally, the distinction lies in the dominant intermolecular interactions. Coulomb forces can be either attractive or repulsive: opposite charges attract, while like charges repel. At low temperatures, as molecular vibrations diminish, attractive van der Waals forces become dominant, driving molecular aggregation and crystallization. At high temperatures, greater molecular kinetic energy and vigorous vibrations subside these attractive forces, promoting repulsion between the valence electrons of water molecules to dominate, leading to molecular dispersion, a tendency that is even more pronounced in gases.

This example demonstrates that, because both definitions of entropy were developed primarily from observations of gaseous systems, their application to systems dominated by attractive forces can yield contradictory results. Under such forces, entropy generally decreases according to the statistical definition. For instance, the crystallization of a supersaturated sodium acetate solution represents a statistical decrease in entropy. This can occur even in an isolated container with no heat exchange with the surroundings, where thermodynamic entropy remains constant. A comparable process occurs when sugar crystals gradually form in a sealed bottle of honey—another case of entropy reduction driven by attractive electrostatic interactions.

Consider another example of entropy decrease: when a bottle of raw, unhomogenized milk is left undisturbed, it spontaneously separates into layers—the cream rising to the top and skim milk settling at the bottom. According to the statistical definition, this is a transition from a more disordered, higher-entropy state to a more ordered, lower-entropy configuration. Because energy is not a parameter in the statistical definition, when it is applied to measure entropy, any energy changes should be disregarded. Yet this spontaneous ordering can also occur in an isolated system, without energy exchange with the surroundings. Similar stratification occurs in other mixtures of components with different densities, such as a can of mixed paint separating over time. In such cases, gravitational attraction leads the system toward greater order, highlighting a limitation of the second law of thermodynamics in describing systems dominated by attractive forces.

For further illustration, under the influence of gravity, stars, galaxies, and other celestial bodies are believed to form from vast cosmic gas clouds that initially exhibit high disorder and numerous accessible microstates. As these clouds collapse gravitationally, matter condenses toward their centers of mass, giving rise to more structured systems such as planets, stars, neutron stars, and black holes. This transition represents a decrease in entropy through increased organization. Similarly, the formation of layered internal structures within celestial bodies, such as Earth's differentiated interior, is also gravity-driven and reflects an entropy reduction. While a small volume of air appears as a well-mixed, high-entropy gas, on a planetary scale, the atmosphere is stratified into distinct layers. After all, observational evidence suggests that the universe tends to evolve toward increasingly structured configurations, rather than toward the traditional "heat death" scenario without organized structures, largely due to the pervasive influence of gravity.

These cases challenge the universal applicability of the second law of thermodynamics, which, although effective in describing a wide range of processes, has inherent limitations. The second law reflects the statistical tendency of systems to disperse—a behavior typically associated with repulsive interactions. In contrast, attractive forces in nature can drive systems to aggregate and form more ordered structures, leading to a reduction in entropy as defined by statistical mechanics.

Limitations of The Second Law of Thermodynamics

The examples in the previous section demonstrate that entropy can decrease under the statistical definitions (5) and (6). But can entropy also decrease under the thermodynamic definition (1)? The answer is yes. Consider two systems with temperatures T_1 and T_2, where $T_1 > T_2$. When these systems interact, a quantity of heat (dQ) flows from the higher-temperature system to the lower-temperature one. According to the thermodynamic definition (1), the total change in entropy of the two systems is given by:

(10) $$dS = \frac{dQ}{T_2} - \frac{dQ}{T_1} = \frac{dQ}{T_1 T_2}\left(T_1 - T_2\right) > 0$$

Thus, the total entropy increases. It is important to note that the derivation of Expression 10 relies on the intuitive assumption that energy always flows down a decreasing gradient. This assumption generally holds in systems dominated by repulsive forces, such as gases, where particles tend to disperse, facilitating mixing and heat transfer. Such processes equalize temperature within the system, producing heat flow along the gradient—a behavior long regarded as an intrinsic property of nature. While valid for most small-scale gas systems, this reasoning does not necessarily extend to large-scale systems, particularly those dominated by attractive forces such as gravity. If energy can flow in the opposite direction, entropy will decrease under the thermodynamic definition (1).

To demonstrate that Expression 10 is not universally valid, it is sufficient to identify a single case in which heat flows from a lower-temperature system to a higher-temperature one, thereby making the right-hand side of Expression 10 negative. Consider a black hole as one system and the rest of the universe as the other. Due to the black hole's immense gravitational pull, all incoming electromagnetic waves, including thermal radiation, are drawn from the surroundings into the black hole. The average temperature of the universe is close to absolute zero, whereas the black hole's temperature is significantly higher. In this scenario, energy flows from the low-temperature system to the high-temperature one, leading to a net decrease in total entropy. This example shows that a decreasing energy gradient does not determine heat flow, but rather the direction of gravity. In other words, energy transfer is governed by fundamental forces. This perspective will be further demonstrated in the next section.

Entropy decreases not only near black holes, but also around most celestial bodies. Take Earth as an example: by analogy with the gas systems discussed earlier, if we regard Earth as system T_1 and the rest of the universe as system T_2, then traveling backward in time would reveal a hotter Earth gradually dispersing into a cold cosmic cloud, similar to the surrounding universe. This rise in entropy parallels the behavior of gas systems described above, suggesting that Earth and the rest of the universe as a whole evolve toward a *decrease* in total entropy. This example also shows that energy flow does not necessarily follow a decreasing temperature gradient but instead is guided by gravity. On cosmic scales, the evolution of the universe—from the birth of stars to the clustering of galaxies—appears to move toward a lower-entropy state, rather than the uniform particle and energy distribution predicted by the "heat death" scenario, under both statistical and thermodynamic definitions of entropy.

According to Einstein's mass–energy equivalence principle, matter is simply a highly concentrated form of energy. Through chemical and nuclear reactions, the energy stored in matter can be released as heat in the form of radiation. In this sense, the flow of matter is effectively a flow of energy, and conversely, the flow of energy—including heat—necessarily involves the movement of matter. This duality is evident in the earlier example of ice in water, where energy transfer involves not only heat (radiation) but also the potential energy between water molecules. At higher temperatures, such as room temperature, repulsive forces dominate intermolecular interactions: shorter distances correspond to higher potential energy, so molecules tend to separate to minimize it, thereby increasing entropy. Below the freezing point, however, attractive van der Waals forces prevail, and minimizing potential energy drives molecular aggregation into solids, decreasing entropy. In such processes, heat transfer occurs both as thermal and potential energy—both inseparably linked to matter itself.

Finally, entropy increase and decrease represent two possible directions of natural evolution, yet the second law of thermodynamics accounts for only one. Beyond these directional progressions, many natural processes are oscillatory or periodic. A pendulum, for example, swings back and forth around its balance position. On Earth, countless daily and seasonal periodic rhythms arise from its rotation and orbital motion. Indeed, oscillation about its balance is a fundamental characteristic of a system. In chemical equilibrium, reactions do not stop; rather, forward and reverse reactions occur at the same rate. From the perspective of a single component, the system can be viewed as oscillating between reactant and product states. The fundamental driver of all these processes is the tendency of systems to restore energy balance. This universal tendency underlies the functioning of the cosmos and is embodied in the Restoration Principle.

The Restoration Principle

The Restoration Principle is explored in detail in a separate article of the same title.[2] A brief overview is provided here to offer context for understanding the contrasting trends in entropy change. According to Einstein's principle of mass–energy equivalence, everything in the universe exists in some form of energy, with matter itself being a concentrated manifestation. The evolution of any system is driven fundamentally by the transformation of its energy. Such transformations tend either to restore or to oscillate around a balanced distribution of energy. The **Restoration Principle** posits that nature exhibits an inherent tendency for systems to evolve toward such balance—defined by a characteristic ratio among different forms of energy at a given total energy level. This balance emerges through continuous energy transformations as the system seeks greater dynamic stability. It reflects the interplay of competing internal drives, with the dominant force directing redistribution to minimize potential energy while conserving total energy. The core concepts are summarized in the following equations and expressions:

(11) $E = M + D$
(12) $D = U + K + R$
(13) $U : K : R \succ u : k : r$

These relations define a system's state of balanced energy distribution at a given total energy. According to Einstein's mass–energy equivalence principle, the total energy (E) of a system equals its effective mass (m) multiplied by the square of the speed of light (c^2). The effective mass includes not only the system's rest mass but also contributions from all other forms of energy. In Equation 11, this total energy is more informatively expressed as the sum of two components: dynamic energy (D) and matter energy (M). Here, M denotes the energy associated with the system's rest mass at absolute zero, while (D) represents the dynamic portion of the system's energy. Temperature typically represents the measurement of radiation level, which is determined by the total dynamic energy of a system, rather than merely the kinetic energy (K) of particle motion, as many textbooks mistakenly suggest.[13] Absolute zero corresponds to the state in which dynamic energy (D) vanishes, requiring potential, kinetic, and radiative energies all to reach their minimum values simultaneously.[14]

Although the energy bound in matter is not usually reflected in temperature measurements, it can manifest as heat once released. For instance, the potential energy between nucleons is normally confined within matter and treated as part of its mass. During nuclear fusion, this energy is liberated along with a corresponding loss of mass, emitting radiation and giving rise to the extreme temperatures of stars such as the Sun. In positron–electron annihilation, the entire mass of both particles is transformed into radiation. Likewise, any fundamental particle can annihilate with its antiparticle, releasing the full energy of their combined mass. Conversely, in pair production, a high-energy photon interacting near a nucleus can generate an electron–positron pair, illustrating the creation of matter directly from radiant energy.

Although matter energy (*M*) and dynamic energy (*D*) can transform into one another, their forms and distributions within a system are primarily governed by the system's total energy. Just as water can exist as a solid, liquid, or gas, depending on its dynamic energy, typically represented by temperature, matter in celestial systems evolves into distinct states determined by total energy. When a celestial body exceeds roughly 0.08 solar masses, gravitational collapse becomes strong enough to overcome Coulomb repulsion, bringing nucleons from different atoms sufficiently close for the short-range nuclear force to act. This initiates the process of hydrogen fusion and the birth of a star. Through nuclear fusion, lighter elements combine into heavier ones, progressing up to the formation of iron. Beyond iron, however, nuclei become unstable: Coulomb repulsion among protons outweighs the short-range nuclear attraction, leading heavy elements to undergo fission or radioactive decay into smaller nuclei. Iron thus represents the most stable nucleus and accumulates within stars across successive stellar cycles.

When a star with an initial mass of 8–25 solar masses develops an iron core exceeding the Chandrasekhar limit (~1.4 solar masses),[15-16] electron degeneracy pressure can no longer counteract gravity, and the core collapses. In this process, electrons combine with protons to form neutrons, giving rise to a neutron star in which matter exists predominantly as neutrons. For stars with initial masses above 25 solar masses, the iron core may exceed the Tolman–Oppenheimer–Volkoff (TOV) limit (~2.2 solar masses).[17-18] In such cases, neutron degeneracy pressure, the strong nuclear repulsion force at short distances, is insufficient to resist further collapse, leading to the formation of a black hole. Within a black hole, matter is thought to be compressed to a singularity of effectively infinite density. Its gravitational field is so strong that not even light can escape, rendering direct observation impossible. Instead, black holes are inferred indirectly from the motion of nearby stars and from radiation emitted by accreting matter. These transformations illustrate the universe's tendency to evolve toward characteristic forms of matter and energy distributions under given energy conditions.

Stars with masses insufficient to exceed the Chandrasekhar limit eventually evolve into white or red dwarfs. Objects with masses below about 0.08 solar masses cannot initiate nuclear fusion at all and thus remain as planetary bodies, such as Earth. In such systems, gravitational forces are no longer the dominant factor determining the form of matter. Instead, Coulomb interactions, dominant at microscopic scales, govern the structure and behavior of matter. Because electrostatic forces can be both attractive and repulsive, and because elements can combine in countless ways, an immense diversity of materials arises. Depending on their dynamic energy levels, matter may exist in solid, liquid, or gaseous states, and myriad chemical reactions occur among different compounds. Within this spectrum of energy states, especially under Earth's unique conditions, life can emerge from material complexity, and intelligence can evolve from living systems.

Unlike large-scale systems dominated by gravity, where transformations from matter energy (*M*) to dynamic energy (D) are common, systems governed by Coulomb forces evolve primarily through transformations and exchanges among potential energy (*U*), kinetic energy (*K*), and radiative energy (*R*) within the broader dynamic energy (*D*), as described in Equation 12. These exchanges continually adjust the relative proportions of the three forms of energy, tending toward a stable configuration characterized by the balance ratio (*u:k:r*), as expressed in Equation 13. For an isolated system at a given dynamic energy level, this ratio remains effectively constant. Within a particular phase (solid, liquid, or gas), the ratio typically falls within a narrow range. Any deviation from this balance initiates internal exchanges among the three energy forms to restore equilibrium, a process often manifesting as periodic oscillations around the balance point. By contrast, long-term progressive developments usually reflect transitions from one equilibrium state to another, typically triggered by shifts in the system’s overall dynamic energy. In all cases, the system evolves toward a balanced distribution among its energy forms.

It is important to emphasize that temperature reflects a system’s radiative energy, not its kinetic energy. The level and relative proportion of radiative energy are determined by the system’s total dynamic energy, which in turn dictates the dominant forces controlling the state of matter—solid, liquid, or gas. At low dynamic energy levels, Coulomb attraction between molecules dominates, forming bonds and increasing the system’s potential energy. To sustain high potential energy under low temperatures, water crystallizes into the solid state. Similarly, the direction of a chemical reaction is governed by the system’s dynamic energy: the input of radiative energy to a heat-sensitive reaction can shift its course by altering the balance among energy components. Likewise, changing the quantities of reactants or products modifies potential energy contributions and thereby the overall energy balance. In short, a system’s total dynamic energy determines both the balance ratio of its energy components and its state of matter.

Changes in entropy—whether increase or decrease—reflect a system’s tendency to restore a balanced energy distribution at different energy levels. At low energy levels, attractive forces dominate, promoting the formation of additional bonds between particles and the storage of greater potential energy. In such conditions, balanced distributions are typically characterized by a high proportion of potential energy. Under these forces, particles tend to aggregate, reducing the number of accessible microstates and thus lowering entropy. This behavior explains the spontaneous formation of crystals and solid matter at low temperatures, as well as why celestial systems evolve from diffuse cosmic clouds into stars and galaxies under gravity, resulting in a decrease in entropy rather than a progression toward “heat death”.

At high dynamic energy levels, repulsive interactions between particles, primarily arising from Coulomb repulsion, become dominant, breaking bonds and dispersing the system. This accounts for why substances melt or vaporize at high temperatures, processes that are accompanied by an increase in entropy. As bonds are broken, potential energy contributions diminish, lowering the potential share within the dynamic energy while relatively increasing other energy forms. From the perspective of the Restoration Principle, such entropy increases represent the system’s response to reestablish a balanced energy distribution under high energy conditions.

This principle likewise helps explain why many objects degrade over time—a process often described as an increase in entropy. Such materials are typically formed under conditions of different energy, and their eventual breakdown reflects the system’s natural tendency to reestablish a balanced energy configuration at a new energy level. For instance, polylactic acid (PLA), a bioplastic derived from corn starch, is synthesized at high temperatures of about 170–200 °C. Once at room temperature, the energy level shifts; to restore balance at this lower level, PLA undergoes hydrolysis of its ester bonds, resulting in rapid degradation. Yet not all natural processes follow the path of breakdown. Crystallization, for example, reflects the opposite trend: the system becomes more ordered and stable, representing another manifestation of the Restoration Principle.

Conclusions

The central idea behind classical entropy is that energy tends to disperse. This holds in systems dominated by repulsive interactions, but the trend can reverse when attractive interactions prevail. Moreover, the thermodynamic and statistical definitions of entropy are not fully equivalent, raising doubts about the universal validity of the second law of thermodynamics—a law that rests on the precise meaning of entropy itself. Challenges to the second law go beyond theoretical inconsistencies; they strike at its core claim that entropy never decreases. Addressing this issue calls for a broader conceptual framework. The **Restoration Principle** offers such a framework by describing the inherent tendency of systems to evolve toward a balanced distribution of energy through the transformation and redistribution of potential, kinetic, and radiative forms. Unlike the second law, which recognizes only entropy-increasing behavior, the Restoration Principle accommodates both entropy-increasing and entropy-decreasing processes, as well as oscillatory dynamics. It thus provides a more general foundation for describing system evolution under the influence of all fundamental forces. Since everything in the universe is ultimately composed of energy, the evolution of any system can be understood as a process of restoring—or oscillating around—its energy balance. The Restoration Principle, therefore, has broad applicability across scientific disciplines, offering deeper insight into the dynamics of natural phenomena.

Revision History

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- 12/17/2025: Adding Links to Summaries of Related Articles

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- https://cs.stanford.edu/people/zjl/abstract.html, PDF
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- https://doi.org/10.5281/zenodo.17967154, PDF

Further Literature

- Misconceptions in Thermodynamics (PDF: DOI) (中文: DOI)
- The Mechanism Driving Crookes Radiometers (PDF: DOI) (中文: DOI)
- The Cause of Brownian Motion (PDF: DOI) (中文: DOI)
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