Dynamic Nature of Thermodynamics

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Abstract

Empirical observations in thermodynamics have long been distilled into a set of laws and principles. This work looks beyond these formulations to investigate the underlying dynamics from which they emerge. All phenomena in the universe may be understood as manifestations of energy, and the evolution of any system is governed by the transformation of energy among its various forms. Consequently, elucidating the mechanisms of energy conversion is central to understanding the processes that drive system evolution. From this perspective, several pervasive misconceptions in classical thermodynamics become apparent. These insights lead to a unifying view of the fundamental principle governing all processes—the *Restoration Principle*—which explains why systems evolve toward equilibrium and offers a clearer, more coherent framework for thermodynamics.

Introduction

Most principles of classical thermodynamics, from the zeroth through the third law, describe macroscopic behavior without explicitly accounting for the microscopic mechanisms that give rise to them. As a result, thermodynamics has long been prone to persistent misconceptions, including misunderstandings of the nature of temperature and absolute zero, the microscopic origins of thermal expansion and contraction, and the mechanisms underlying Brownian motion and phase transitions. This article summarizes a series of studies that investigate the fundamental processes governing these phenomena, clarifying common misconceptions, developing a conceptually coherent understanding of thermodynamics from first principles, and establishing a more accurate and unified theoretical framework.

Transimpact

When an electron in an atom absorbs energy, it is excited to a higher orbital in a process known as an atomic electron transition. This transition occurs spontaneously and typically within a few nanoseconds. As the valence electron moves to the higher orbital, the outer electron cloud expands, altering the distances to neighboring particles and disturbing the equilibrium spacing maintained by Van der Waals forces. Because the orbital jump happens abruptly, it generates a strong repulsive effect on nearby particles. This phenomenon is analogous to the sudden burst of a popcorn kernel, where the rapid motion exerts substantial force on adjacent objects. This process has been termed transimpact**,** derived from "transition impact."

A transimpact can be illustrated using two monatomic molecules positioned next to each other, as shown in Figure 1. The equilibrium distance between the molecules is determined by a balance of attractive and repulsive van der Waals forces, depicted in Figure 1A. When an atom absorbs energy, an electron transition occurs, causing the molecular size to increase and the distance to the neighboring molecule to decrease, as shown in Figure 1B. This disrupts the van der Waals equilibrium, causing a sudden increase in the repulsive force and pushing the molecules apart, as illustrated in Figure 1C.

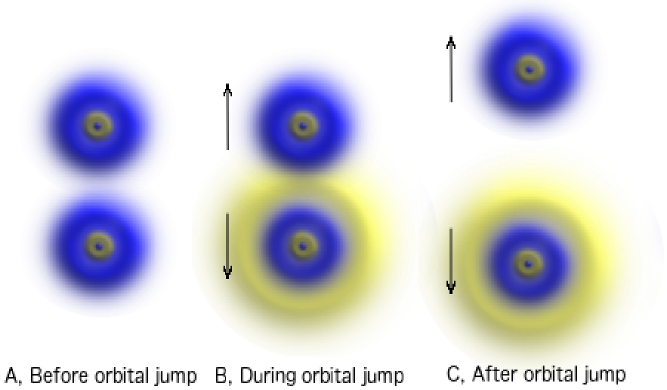

Figure 1: Transimpact due to an atomic electron transition.

Transimpacts are common phenomena that occur whenever an electron undergoes an orbital transition, involving an energy exchange between an atom and its surroundings. Consequently, they play a crucial role in many everyday processes, including Brownian motion and phase transitions. Transimpacts provide the fundamental mechanism by which radiative energy is converted into the kinetic energy of particles, making them one of the key processes driving thermodynamic behavior.

Unveiling the Mechanism Behind Crookes Radiometers

A Crookes radiometer consists of a low-pressure glass bulb containing a set of vanes mounted on a low-friction spindle. Each vane is black on one side and white on the other. When exposed to light, the vanes begin to rotate, with the black sides retreating from the light source; the rotation speed increases with higher light intensity. The mechanism underlying this motion has been debated for over a century. The device's inventor, William Crookes, originally attributed the rotation to photon pressure, as predicted by James Clerk Maxwell. However, this explanation fails on two key grounds: first, if photon pressure were responsible, the vanes would rotate in the opposite direction because the white side reflects more light; second, the rotation should increase as the vacuum improves. In practice, the vanes remain motionless in a hard vacuum, and the maximum rotation occurs at low but nonzero pressures, around one pascal.

Numerous theories attribute the operation of the Crookes radiometer to aerodynamic effects or pressure differentials across the vanes or their edges. However, such air-based explanations fail to account for a critical observation: the rapid onset of rotation, particularly the pronounced initial acceleration. According to Newton's second law, acceleration is proportional to the net applied force; thus, a large initial acceleration implies the presence of a substantial driving force at the very onset of motion upon illumination. For stationary vanes to begin rotating, this force must first overcome the static friction at the spindle. In models based on air pressure or aerodynamic effects, pressure gradients develop only gradually as the gas is heated, resulting in a slow buildup of force and, consequently, an acceleration that increases from zero. This prediction is inconsistent with observations, which consistently show that the acceleration reaches its maximum immediately upon illumination, directly contradicting the gradual-force development expected from air-based mechanisms.

The transimpact theory has been proposed to explain the rotation of Crookes radiometers. When light strikes radiometer vanes, atomic electron transitions occur, triggering instantaneous transimpacts. These transimpacts repel nearby air molecules from the vane surface, transferring momentum to the vane and generating a force. Because the black sides of the vanes absorb light more efficiently, transimpacts occur more frequently there, creating a force imbalance between the two sides. Since transimpacts respond immediately to incident light, the theory predicts the strongest driving force—and thus the greatest acceleration—at the onset of illumination. As the vanes begin to rotate, increasing air resistance gradually reduces their acceleration over time.

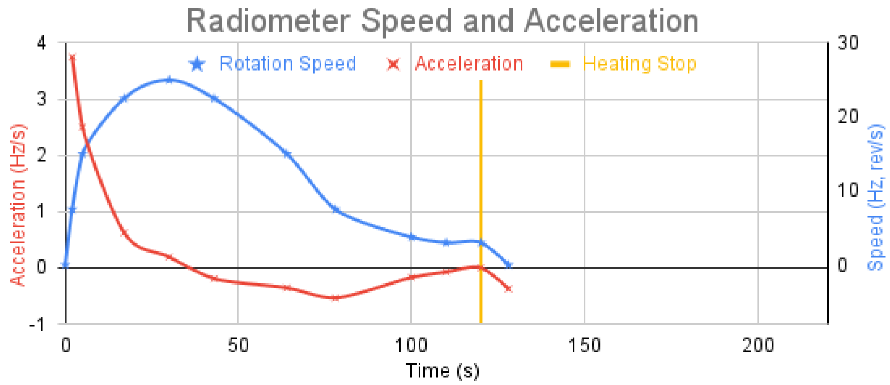

Figure 2: Speed and acceleration profile of the radiometer in a high-power flashlight experiment, showing that the maximum acceleration and driving force occur at the onset of incident light.

Experiments were conducted to test this prediction, and the results closely align with theoretical expectations, as shown in Figure 2. These findings provide strong support for the transimpact theory. The red curve shows that acceleration is maximal immediately upon illumination, indicating that the net driving force is greatest at the onset. As the vanes accelerate, increasing air resistance gradually reduces the net driving force. Eventually, the resistive force balances the driving force, and the system reaches a steady rotational speed.

What Causes Brownian Motion?

Albert Einstein successfully described Brownian motion using a statistical random-walk framework. At the time, his work provided compelling evidence for the existence of molecules, explaining the erratic motion of suspended particles as the result of collisions with randomly moving molecules. Nevertheless, despite the widespread acceptance of this explanation, the fundamental mechanism driving Brownian motion has yet to be definitively proven.

To evaluate the proposed theory, we estimated the velocity a water molecule would need to impart observable Brownian motion to a pollen grain. Our calculations indicate that, to move a Clarkia pulchella pollen grain with a typical diameter of 50–100 µm at the minimum detectable velocity of 300 nm/s, a water molecule would need to travel faster than 3.21 × 10^{7} m/s—over 10% of the speed of light. By comparison, at room temperature, the average velocity of water molecules, based on their kinetic energy distribution, is approximately 590 m/s. Considering that sound propagation depends on molecular interactions, this estimate aligns with the measured speed of sound in water (~1,500 m/s), which represents an approximate upper bound for the average velocity of water molecules. Therefore, the velocity required to explain Brownian motion via direct molecular impacts is physically implausible.

Assuming the speed of water molecules follows a Gaussian distribution with a mean of 1,500 m/s, representing the upper bound of their typical average random motion, the probability of a single molecule exceeding 3.21×10^{7} m/s is less than $10^{-98,000,000}$—effectively zero. Even if surrounding water molecules struck a pollen grain every yoctosecond, the expected time to observe measurable Brownian motion would far exceed the age of the universe.

Even when considering the simultaneous impact of multiple water molecules, producing an observable displacement of a pollen grain would require more than 54,000 molecules to collide in the same direction while none strike from the opposite side. In three-dimensional space, where each molecule can move in infinitely many directions, such precise alignment is extraordinarily improbable. The likelihood that all 54,000 molecules move nearly in the same direction simultaneously is no greater than the exceedingly small probability estimated above. Even under a highly simplified model in which each molecule has only two possible outcomes—collide or not collide—the likelihood of this synchronized action is already vanishingly small, $2^{-54,000}$. For all practical purposes, the chance of such a highly coordinated, asymmetric impact occurring spontaneously is negligible. Consequently, the hypothesis that Brownian motion results from the collective effect of random molecular collisions lacks sufficient support.

If the motion of molecules cannot provide sufficient momentum to drive Brownian motion, what mechanism is truly responsible? We propose that transimpacts can generate far stronger interactions than the random motion of molecules. When a transimpact occurs between a pollen grain and an adjacent water molecule, it produces strong repulsive forces that effectively push the particles apart, offering a more plausible explanation for the observed dynamics of Brownian motion. In fact, the random motion of water molecules themselves originates from transimpacts. Thus, the apparent randomness in the movement of both water molecules and pollen grains can ultimately be traced back to these underlying transimpact interactions.

The Century-Old Dilemma of the Kinetic Theory of Gases

In addition to explaining the operation of Crookes radiometers and providing a mechanism for Brownian motion, transimpacts offer a way for particles to acquire kinetic energy. This association appears to underlie the widely cited connection between temperature and particle kinetic energy, often summarized by the statement that temperature represents the average kinetic energy of molecules. While this notion originates from the kinetic theory of gases and is commonly presented in textbooks and reference sources such as Wikipedia, it is not accurate.

If this claim were accurate, changes in temperature and kinetic energy could be used to predict the specific heat of a system using kinetic theory, as expressed by the following formula:

$$(1) \qquad C_m = \frac{3}{2}R$$

Here, C_m represents the molar-specific heat capacity predicted in the kinetic theory, and R is the gas molar constant. These theoretical predictions can be directly tested through experimental measurements. For monatomic gases—such as helium, neon, argon, krypton, xenon, and radon—kinetic theory predictions closely match experimental data, regardless of atomic mass, with an accuracy of approximately 99%. In contrast, the theory fails to predict the specific heat of polyatomic gases. The predicted values are consistently lower than the measured ones, and the discrepancy grows with increasing molecular complexity. For example, in the case of octane, the predicted specific heat accounts for only about 7% of the measured value, corresponding to a discrepancy of roughly 93%.

To address these discrepancies, the kinetic model was extended to include vibrational and rotational energies in addition to the translational energy of particle motion. Maxwell proposed that energy is equally partitioned among these three modes, and Boltzmann later refined this idea by introducing the concept of degrees of freedom, arguing that energy in a gas is equally distributed among all accessible degrees of freedom. Nevertheless, even with this extension, the model still fails to fully account for the observed discrepancies, a longstanding puzzle in thermodynamics for more than a century. Numerous explanations have been proposed to resolve the apparent breakdown of the kinetic theory, even when extended with the equipartition principle. Boltzmann suggested that gases might not be in complete thermal equilibrium, while Planck and Einstein introduced the concept of the zero-point harmonic oscillator. Kelvin ultimately concluded that the equipartition assumption itself might be flawed—a position later acknowledged by Einstein.

The Neglected Role of Potential Energy

The dilemma of the kinetic theory does not arise from its applicability being limited to ideal gases, as it provides accurate predictions for monatomic gases. Rather, it stems from the assumption that temperature reflects only the kinetic energy of a system, thereby neglecting potential-energy contributions as reservoirs of specific heat. When heat is added to a system, it increases not only the kinetic energy of the particles but also the potential energy between them. For example, when an electron absorbs energy, it may be excited to a higher orbital, thereby storing the absorbed energy as potential energy. Similarly, bonds between atoms and molecules behave like springs, storing potential energy when stretched. As work is performed to increase the separation between bonded atoms or molecules, the energy is retained as potential energy within the bonds. Notably, chemical bonds can typically store a substantially larger fraction of potential energy than electronic orbitals, since atoms or molecules are approximately three orders of magnitude heavier than electrons. This perception becomes particularly evident when comparing molar specific heat values with the number of chemical bonds in gas molecules, as shown in Figure 3.

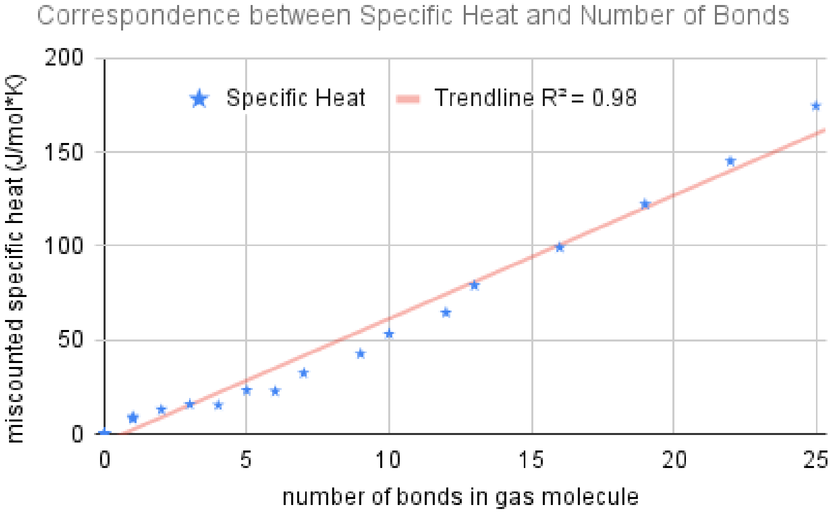

Figure 3: Correlation between molar specific heat and the number of molecular bonds in gas molecules. Each data point represents a distinct gas. The “miscounted” molar specific heat on the vertical axis denotes the measured values offset by the kinetic-theory prediction of 1.5R. The fitted trendline demonstrates a strong linear correlation, with a coefficient of determination R^2 as high as 0.98.

These data indicate that specific heat is not solely devoted to increasing kinetic energy; a substantial portion instead contributes to an increase in the potential energy stored in chemical bonds. Moreover, the fraction of energy stored as potential energy grows approximately in proportion to the number of bonds in the gas molecules. This observation indicates that neglecting potential-energy contributions accounts for the discrepancy between the specific heat predicted by kinetic theory and the experimentally measured values. This finding further implies that temperature cannot be interpreted solely as a measure of molecular kinetic energy.

What Energy Does Temperature Measure?

The preceding observations suggest that the specific heat of a system is more closely associated with potential energy, particularly for gases with complex molecular structures, which might imply that temperature is closely related to the system's potential-energy level. However, this interpretation is also not accurate. When a body's temperature is measured using an infrared thermometer, the device detects the level of thermal radiation emitted by the body without any physical contact. In this process, neither the system's potential energy nor its kinetic energy is measured directly.

The radiative behavior of an ideal blackbody is described by Planck's law, which specifies the spectral distribution of radiation as a function of wavelength and temperature. Planck's law implies that the total radiative power P emitted by a perfect blackbody is proportional to the fourth power of its absolute temperature, as expressed by the Stefan–Boltzmann law. By rearranging this relationship, the temperature of a target system can be expressed in terms of its emitted power:

$$(2) \qquad T = \sqrt[4]{\frac{P}{\sigma}}$$

Here, σ is the Stefan-Boltzmann constant. This relationship indicates that temperature can be determined from the radiation emitted by a system, forming the theoretical basis for designing infrared thermometers. Another direct consequence of Planck's radiation law is Wien's displacement law, which states that the peak wavelength of emitted radiation shifts inversely with temperature:

$$(3) \qquad T = \frac{b}{\lambda}$$

Here, λ represents the peak wavelength of the radiation curve, and b is Wien's displacement constant. This relationship allows temperature to be determined by finding the peak wavelength of radiation emitted by a system. For instance, an experienced baker can tell oven temperature by observing the color of the flame. Similarly, the Sun's surface temperature can be determined from its peak emission near 500 nm, corresponding to approximately 5,778 K. These examples illustrate how temperature is primarily linked to the radiative energy emitted by a source.

Nevertheless, some devices determine temperature by measuring a material's electrical resistivity. Since resistivity is closely related to the material's potential-energy states, temperature can, in part, reflect these potential-energy components. This connection arises because thermal emission is linked to the system's potential-energy levels. For example, an electron in a higher orbital, and thus at higher potential energy, is more likely to transition to a lower orbital, releasing energy as radiation. Similarly, a stretched chemical bond tends to contract, emitting its stored potential energy as radiation. Through such mechanisms, the system's potential-energy states influence its emission, which in turn correlates with the measured temperature.

As will be discussed later, the kinetic energy level is also linked to the levels of potential energy and radiative energy, and is thus indirectly reflected by temperature. However, this connection is relatively weak and can sometimes become decoupled, as occurs during phase transitions at constant temperature.

Impactrans

While transimpacts drive the accumulation of kinetic energy within a system, this buildup cannot continue indefinitely. A complementary mechanism must convert kinetic energy into other forms—a process referred to as impactrans. The motion or vibration of particles (including molecules, atoms, or subatomic particles) can lead to collisions with adjacent particles mediated by electrostatic interactions. These collisions may excite electrons to higher orbitals or even eject them entirely, thereby altering their potential energy. Furthermore, the accelerated motion of electrons during such interactions can generate radiation. Through impactrans, kinetic energy is thus transformed into both potential and radiative energy, helping to maintain a dynamic equilibrium among the different forms of energy within the system.

The conversion of kinetic energy into radiative and potential energy through impactrans is evident in many everyday phenomena. A familiar example is the heat generated at the base of an air pump tube. As air is compressed at the tube base, intensified molecular collisions increase the frequency of impactrans events. Through these interactions, electrons can be accelerated and emit radiation, which manifests as the heat observed at the base of the tube. In this way, the kinetic energy of the pump's motion is converted into radiative energy. During this process, some electrons may transition between orbitals, thereby altering their potential energy.

Impactrans can also occur between different systems, explaining the origin of frictional heating. A familiar example is the warmth felt when rubbing one's palms together. Although the palms appear smooth, they are microscopically rough. Kinetic energy from the relative motion of the surfaces is transferred into particle-level collisions at surface asperities. These collisions can accelerate electrons and modify their orbitals or alter bond structures, and the associated changes in potential energy are released in the form of thermal radiation.

In the case of static electricity, rubbing a plastic rod with fur dislodges electrons from atoms, leading to the accumulation of electric charge—an example of impactrans in action. Another illustration is the dent formed when a surface is struck by a hammer. The resulting deformation arises from changes in bond structures: some bonds are compressed while others are stretched, thereby absorbing the hammer's kinetic energy and storing it as potential energy. During the impact, electrons are also accelerated, producing radiation in the form of heat, commonly perceived as impact heating that frequently accompanies collisions. At a fundamental level, this process can be understood as a collective impactrans involving a vast number of interacting particles.

The Triangle of Energy Transformation

The three fundamental forms of energy—kinetic, potential, and radiative—coexist within a system. Kinetic energy is carried by the motion of particles, while radiative energy continuously propagates among them. Potential energy is stored in orbital electrons as well as in chemical and intermolecular bonds. Each form of energy can be converted into the others through the mechanisms indicated along the adjacent edges of the Triangle of Energy Transformation, as illustrated in Figure 4.

For example, through **Absorption**, radiative energy is transformed into potential energy stored in bonds and orbital electrons, whereas through **Emission**, potential energy is released as radiation. Particles can acquire kinetic energy through **Transimpacts**, while kinetic energy can be converted into both potential and radiative energy through **Impactrans** processes. Potential energy can also be transformed into kinetic energy through **Conversion**.

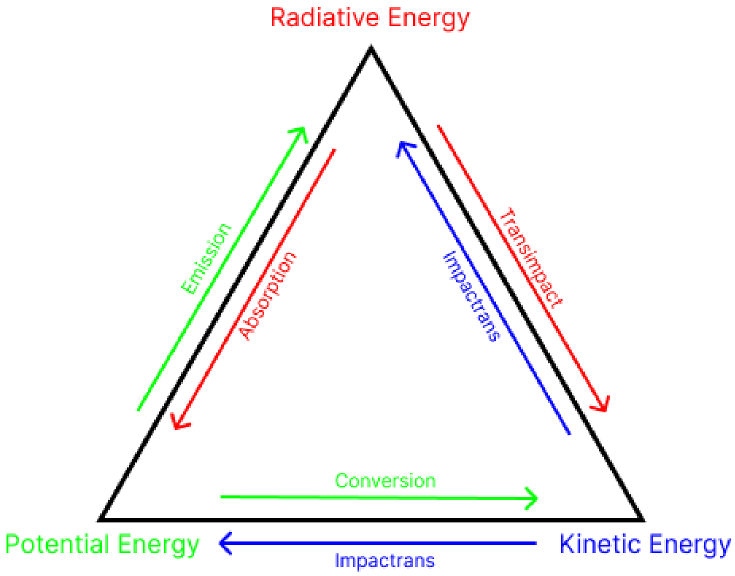

Figure 4: The Triangle of Energy Transformation. Each form of energy is transformed into the other forms along the adjacent edges of the triangle through the mechanisms indicated by the arrows.

Through **absorption**, **emission**, **conversion**, **transimpact**, and **impactrans**, energy can be transformed among three dynamic forms: radiative, potential, and kinetic. Collectively, these mechanisms enable the exchange and redistribution of energy within a system or between interacting systems, driving their evolution toward equilibrium and thereby underpinning the **Zeroth Law of Thermodynamics**.

What Energy Does Temperature Represent?

This question becomes easier to address with the introduction of the **Triangle of Energy Transformation**. The three forms of energy—radiative, potential, and kinetic—continuously transform into one another. They are distinct manifestations of a system's energy and can be collectively described by a new term, Dynamic Energy, defined as:

$$D = R + U + K \tag{4}$$

Here, *D* denotes the system's dynamic energy, consisting of radiative energy *R*, potential energy *U*, and kinetic energy *K*. The continual transformation among these forms drives the system's evolution over time. Treating them collectively as dynamic energy emphasizes their intrinsic interdependence and ever-changing nature. This framework provides both a foundational perspective and an underlying mechanism for understanding thermodynamic behavior.

Therefore, temperature should reflect the dynamic energy of a system. At equilibrium, a system's temperature can be inferred by measuring any one of its components, such as radiative or potential energy, and different types of thermometers are designed on this principle. In contrast, there is no practical device designed to measure kinetic energy directly for temperature determination, largely due to technical limitations. However, this is not the fundamental reason why temperature does not simply represent kinetic energy. Rather, temperature can decouple from kinetic energy during phase transitions that occur at constant temperature—a point that will be explored later in this article regarding phase transitions.

What is The Nature of Absolute Zero?

The concept of absolute zero was first inferred from Charles's law. In classical thermodynamics, absolute zero is defined by the third law of thermodynamics in terms of entropy. With the introduction of **dynamic energy**, the nature of absolute zero can be understood more directly as the state of a system in which its dynamic energy is zero. This implies that all three components of dynamic energy—radiative, potential, and kinetic—are at their lowest possible levels.

According to Einstein's mass–energy equivalence principle, the total energy of a system is intrinsically tied to its mass. This principle reveals that matter itself is a concentrated form of energy, indicating that a system's total energy includes not only its dynamic energy but also the energy contained in matter. Accordingly, the total energy of a system can be decomposed into the sum of dynamic energy D and matter energy M:

$$E = D + M. \tag{5}$$

Temperature represents the **dynamic fraction** (D) of a system's total energy. Under normal conditions on Earth, the energy contained in matter itself—such as the potential energy between nucleons—is not reflected by temperature measurements unless it is released and converted into dynamic energy. Absolute zero corresponds to the complete absence of dynamic energy, meaning that potential, kinetic, and radiative energy are all minimized. For example, for a crystal to reach absolute zero, all radiation must be eliminated, molecular vibrations must cease, all electrons must occupy their lowest possible orbitals, and all bonds must be at their shortest lengths.

Although matter energy is not typically reflected in temperature measurements, it can influence temperature once it is released from matter. For example, nuclear fusion can liberate enormous amounts of potential energy stored between nucleons, accompanied by a corresponding loss of mass. On Jupiter, however, this nuclear potential energy remains locked within matter and is not released spontaneously under normal conditions. In contrast, fusion occurs naturally in the Sun, where this energy is continuously liberated and converted into dynamic energy, producing intense heat. This explains why the Sun, despite having a hydrogen abundance similar to Jupiter's, reaches extremely high temperatures, whereas Jupiter remains comparatively cool.

These examples highlight the importance of distinguishing between **matter energy** and **dynamic energy**. Matter energy refers to the intrinsic energy stored within matter. In contrast, dynamic energy encompasses the forms of energy—kinetic, potential, and radiative—that can interconvert under the prevailing conditions of a system. For instance, the potential energy stored in chemical bonds is typically accessible under terrestrial conditions, while the potential energy in nuclear bonds is normally inaccessible. In the Sun, however, a portion of nuclear potential energy becomes accessible through fusion and is converted into dynamic energy, manifesting directly as the Sun's high temperature.

Why Does Thermal Contraction Occur?

Classical textbooks often attribute thermal expansion to increased molecular vibrations, arguing that as temperature rises, particles vibrate more vigorously and therefore require more space. If vibration alone were responsible for expansion, a substance's volume would increase monotonically with temperature, since vibrational intensity generally increases with additional energy. However, this explanation fails to account for thermal contraction. A well-known example is water, which reaches its maximum density at 4 °C and expands both above and below this temperature. Such behavior directly contradicts the predictions of the vibration-based model, revealing fundamental shortcomings in its assumptions. Moreover, the thermal expansion of many crystals is anisotropic, exhibiting different expansion rates along different crystallographic directions.

Thermal expansion is fundamentally linked to molecular structure, particularly to changes in bonds. As energy is added to a system, intramolecular and intermolecular bonds stretch to accommodate the added energy, increasing the average separation between particles and thereby causing the material to expand. Thermal expansion manifests through bond-structure changes in two principal ways: in crystalline solids, expansion can be direction-dependent due to anisotropy in the crystal lattice; more generally, the volumetric thermal expansion coefficient tends to correlate with molecular complexity, since more complex structures offer greater bond flexibility and a higher capacity for energy absorption.

Importantly, a substance's volume is influenced not only by individual bond lengths but also by the overall geometry of its bonding network. This understanding helps explain anomalous behaviors such as thermal contraction in systems where bonding favors more compact arrangements, notably water near 4 °C. Hydrogen bonds generate strong intermolecular attractions between water molecules. As temperature decreases, these bonds shorten and strengthen, drawing molecules closer together and reducing the volume—an example of normal thermal behavior.

Around 4 °C, however, the hydrogen bonds become sufficiently strong to form a scaffold-like network that constrains water molecules into relatively fixed positions, producing open, hollow structures similar to those found in snowflakes. As the temperature continues to fall, this network becomes increasingly pronounced, leading to further expansion even as the system cools. Although conventional thermal expansion due to bond length changes still operates, its effect is outweighed at lower temperatures by structural rearrangements driven by hydrogen bonding.

What Drives Phase Transitions?

Another common misconception is that phase transitions are driven primarily by increasingly intense molecular vibrations at elevated temperatures. In reality, molecular vibrations arise from **transimpacts**, which are more fundamental and energetic than the vibrations themselves, as they constitute the mechanisms by which vibrational energy is generated and transmitted. In most cases, intermolecular bonds are disrupted by transimpacts before vibrational amplitudes become large enough to induce structural change. Phase transitions are therefore primarily initiated by transimpacts rather than by vibration amplitude alone. More accurately, phase transitions should be understood as the outcome of an interplay among three key factors: ambient pressure, the bond forces that maintain molecular cohesion, and transimpacts that act to destabilize and break those bonds.

At lower temperatures, the transimpact force *T* is weaker than the combined retaining forces of intermolecular bonding *B* and ambient pressure *P*, allowing the system to preserve its structural integrity. As temperature increases, the average intensity of transimpacts grows, while bond strength generally diminishes as intermolecular distances expand to accommodate higher potential energy. Eventually, a critical temperature is reached at which the transimpact force becomes comparable to—and ultimately exceeds—the combined retaining forces (*B* + *P*), triggering a phase transition. This condition can be expressed as an equilibrium among the three forces:

$$T = B + P \tag{6}$$

This equation represents the phase-transition boundary of a substance in a phase diagram. A similar balance applies to gas–plasma transitions, where the dominant binding force is the electrostatic attraction between electrons and their nuclei, and transimpacts are effectively replaced by atomic electron transitions. In this context, the transition corresponds to the ionization of atoms, as electrons gain enough energy to be freed from their nuclei.

It is important to note that Equation 6 represents an average balance between the retaining forces (pressure and bonding) and the disruptive force of transimpacts. This equilibrium is typically reached as bonds weaken at elevated temperatures. However, individual transimpacts of exceptionally high intensity can overcome even relatively strong bonds at lower temperatures, ejecting molecules from the surface and producing phenomena such as sublimation or evaporation. A similar principle applies to the photoelectric effect, where high-energy photons can eject electrons from a material at low temperatures without inducing a plasma phase transition. Accordingly, these transitions can be classified as either **bond-limited** or **impact-driven**, as summarized in Table 1.

	Solid to Liquid	Liquid to Gas	Gas to Plasma
Bond-Limited Transition (High Temperature)	Melting	Boiling	Ionization
Impact-Driven Transition (Low Temperature)	Sublimation	Evaporation	Photoelectric Effect

Table 1: Comparison between bond-limited and impact-driven transitions

Why Temperature Remains Constant During Phase Transitions

Classical textbook explanations assert that the latent heat supplied during a phase transition is entirely consumed in breaking intermolecular bonds, resulting in no net increase in molecular kinetic energy and, consequently, no change in temperature. This reasoning relies on the implicit assumption that temperature directly reflects the average kinetic energy of molecules—an assumption that, as this study suggests, may be fundamentally flawed.

During a phase transition, such as the vaporization of water into steam, molecules are released from the hydrogen-bonded molecular network, gaining greater mobility. As these molecules escape the structured network, their kinetic energy must increase. If temperature truly reflected the average kinetic energy of the system, it would be expected to rise during the transition. So why, then, does temperature decouple from the change in molecular kinetic energy?

In light of the insights presented thus far, this apparent paradox becomes more comprehensible. As energy is added to a system before a phase transition, the temperature rises, reflecting an increase in all three components of dynamic energy. In particular, electrons are excited to higher orbitals, causing atoms to move farther apart and thereby stretching and weakening intermolecular bonds. Simultaneously, the system's kinetic energy increases, driven by the intensifying transimpacts.

However, the stretching of bonds has a limit. Eventually, they can no longer withstand the increasing intensity of transimpacts and begin to break, marking the onset of the phase transition. The maximum bond extension places a ceiling on the system's potential energy and the level of radiation emitted by the strained bonds. This effectively anchors the system's temperature as measured via radiative and potential energy. Meanwhile, although the kinetic energy of the molecules continues to increase, no device can directly measure this change during the phase transition.

Variable Temperature Phase Transitions

However, not all phase transitions occur at a constant temperature. For a transition to proceed isothermally, two conditions must be satisfied: first, the transition must occur slowly enough for the anchoring effect to dominate the system; second, the energy required to alter intermolecular bonds must compete for the energy needed to change the system's temperature. For instance, when ice melts into water, the additional energy is used both to break intermolecular bonds and to raise the temperature. These competing demands prevent a simultaneous temperature change during the transition. When this process occurs slowly, it typically proceeds at a constant temperature. However, if the process occurs rapidly—as in supercooled or superheated water—it does not maintain a steady temperature. Another example illustrating the violation of the first condition is the crystallization of supersaturated sodium acetate, which proceeds too quickly for the anchoring effect to stabilize the temperature.

An example that violates the second condition is the superfluid transition of helium. To convert helium from its normal liquid state to the superfluid state, energy is extracted to lower the temperature. During the superfluid transition, the breaking of intermolecular bonds between helium molecules also consumes energy, further reducing the system's temperature. Since these two processes do not compete for energy, the transition does not occur at a fixed temperature. Instead, the temperature gradually decreases throughout the transition, spanning from just above 2.6 K down to 2.17 K.

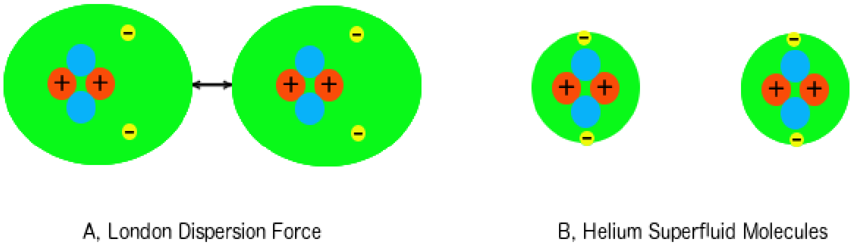

Figure 5: A helium superfluid is a collection of molecules in the absence of their interactions. (A) The interactions form between helium molecules due to London dispersion forces at high temperatures. (B) These interactions vanish as electrons retreat to lower-energy orbitals at low temperatures, weakening the London dispersion forces.

Liquid helium consists of single-atom molecules. The viscosity of the liquid arises from interactions between helium molecules, which are dominated by London dispersion forces. At higher temperatures, electrons occupy higher-energy orbitals and are more susceptible to perturbations, leading to increased polarizability and stronger dispersion forces between molecules, as illustrated in Figure 5A. In contrast, at lower temperatures, helium electrons settle into lower-energy orbitals and become more tightly bound to their nuclei, reducing their polarizability and weakening the dispersion forces. As the temperature continues to decrease, these forces diminish to the point that intermolecular interactions become negligible. As shown in Figure 5B, the loss of attractive forces effectively eliminates bonding between molecules. In the absence of these interactions, individual helium molecules can move independently, resulting in a fluid with vanishing viscosity—a defining characteristic of superfluidity. In this sense, helium superfluid is not a conventional interacting fluid but rather a collection of freely moving, effectively non-interacting atoms.

As energy is removed from liquid helium, its temperature decreases, and intermolecular interactions weaken as the London dispersion forces are reduced. Because the disruption of these weak intermolecular bonds also consumes energy, this process further contributes to the temperature decrease. Consequently, the system does not remain at a fixed temperature during the transition but instead evolves gradually over a finite temperature range. This range can be approximately identified by the irregular changes in the curvature of the helium specific heat curve, as shown in Figure 6.

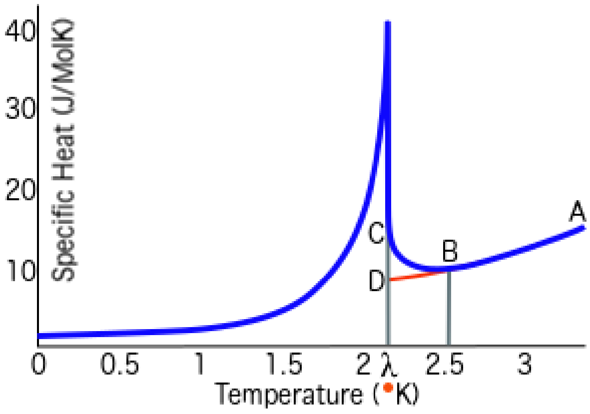

Figure 6: Identifying the temperature range of the helium-4 superfluid transition. The blue curve depicts the specific heat of helium-4 as a function of temperature. The irregular change in curvature between points B and C suggests that the transition begins before the λ-point at approximately 2.17 K.

In a homogeneous liquid, the specific heat typically decreases smoothly from point A as energy is removed. Based on this behavior, one would expect the curve to continue its downward trend from A through B to D, as indicated by the red extrapolated curve. However, a distinct deviation occurs at point B: the pronounced upward curvature from B to C signifies a dramatic increase in specific heat, indicating a substantial rise in the number of superfluid atoms—those that have broken free from the interactions between liquid helium molecules—and signaling a major phase transition within the liquid.

The markedly high specific heat to the left of the λ-point indicates that molecules in the superfluid phase have a higher specific heat than those in the normal liquid phase. This elevated specific heat likely arises from the absence of intermolecular bonds, as freely moving molecules can carry significantly more kinetic energy, contributing to the overall heat capacity. Consequently, the increasing specific heat observed to the right of the λ-point suggests the early appearance of superfluid molecules, implying that the transition begins before reaching the λ-point. Therefore, the onset temperature of the superfluid transition should be considered at least from point B, if not earlier. The temperature range of the helium-4 superfluid transition can thus be approximated as spanning from point B to the λ-point. However, the upper limit of this range remains uncertain due to the subtle curvature changes that mark the initial onset of the transition.

What Does a System Equilibrate To?

It is commonly observed that a system will evolve toward a state of equilibrium. What, then, does a system equilibrate to? The energy within a system undergoes continuous transformation among different forms. Over time, these transformations converge toward stable distributions among the various forms of energy, which can be expressed as:

$$U : K : R \succ u : k : r \tag{7}$$

$$u + k + r = 1 \tag{8}$$

These relations indicate that the three forms of dynamic energy—potential energy (U), kinetic energy (K), and radiative energy (R)—tend to converge toward a stable ratio $u : k : r$, where $u + k + r = 1$. For any system, regardless of its initial energy distribution, each energy form continuously converts into the others along the edges of the Triangle of Energy Transformation, with conversion rates proportional to its fractional contribution. An energy form with a larger fraction transforms into the others more rapidly, thereby reducing its relative share. Over time, the transformation rates balance, and the relative proportions of the three energy forms converge to a stable distribution, known as the equilibrium ratio.

The equilibrium distribution ratio among different forms of energy may differ across systems and may vary with the system's state. Importantly, this equilibrium is dynamically maintained rather than static. A disturbance to the equilibrium triggers a spontaneous process that either restores the original equilibrium—if the total dynamic energy remains unchanged—or establishes a new equilibrium—if the total dynamic energy has changed. This behavior is referred to as the **Restoration Principle**.

Everything in the universe can be understood as a manifestation of energy, and every process is fundamentally an expression of energy transformation. The Restoration Principle is therefore foundational to all natural phenomena. A periodic change may be interpreted as a restoration process oscillating about an equilibrium state, such as a reversible chemical reaction. By contrast, a directional progression may represent a segment of a longer-term process that restores a system toward equilibrium. For instance, the second law of thermodynamics encompasses a subset of these processes, describing cases in which entropy evolves in a direction that promotes the equilibration of energy distribution. This topic will be examined in detail in the following sections.

A Simplified Model for Understanding Entropy

The second law of thermodynamics is often misunderstood; it describes the typical evolution of a system in terms of entropy. To facilitate further discussion, we present an intuitive framework for understanding entropy. In classical thermodynamics, the change in entropy of a system (dS) is defined as the infinitesimal amount of heat transferred (dQ) from the surroundings to the system during a reversible process, divided by the system's instantaneous temperature (T) at the time of transfer:

$$dS = \frac{dQ}{T} \tag{9}$$

To grasp the nature of entropy, consider a gaseous system of mass *m* with a specific heat capacity *c* over a small temperature range. This allows us to treat *c* as constant within the small range. Under these simplifications, the entropy defined above can be expressed in terms of the temperature change resulting from the heat input:

$$(10) \qquad dS = \frac{dQ}{T} = \frac{mcdT}{T}$$

Let *k* stand for the constant *mc*. By integrating this equation over a temperature range from 1 K to *T*, we found:

$$(11) \qquad S = kln(T)$$

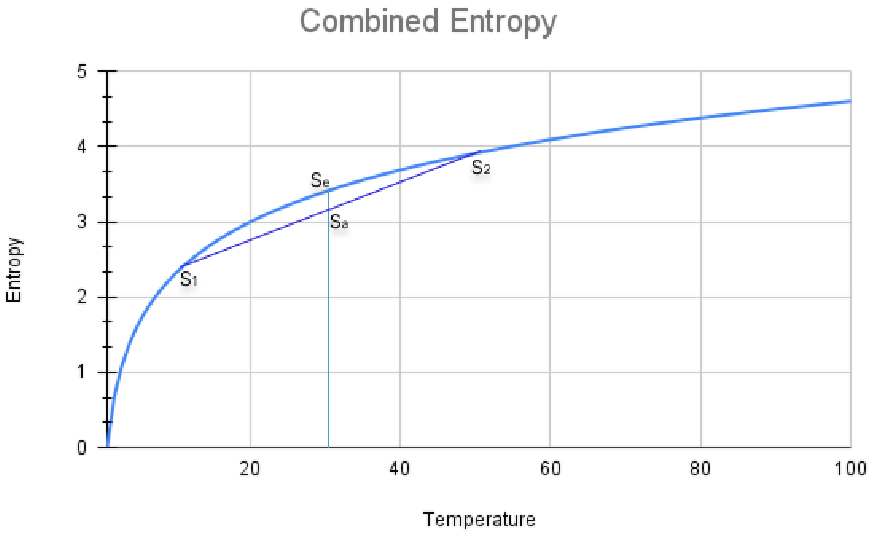

Figure 7: The relationship between entropy and temperature in a simplified model.

In this model, it is easy to see that as heat is added to a system, its temperature rises and its entropy increases, as illustrated in Figure 7. More importantly, entropy increases logarithmically with temperature.

The Nature of the Second Law of Thermodynamics

The simplified model above offers valuable insight into the concept of entropy without sacrificing generality. Consider, for instance, two systems of equal mass but with different temperatures and initial entropies, S_1 and S_2. The average entropy, S_a, represents the initial entropy of the combined system. As the two systems interact and exchange heat, the total entropy becomes S_e at equilibrium. As illustrated in the figure, the entropy of the combined system at equilibrium is always greater than the average of the two initial entropies, demonstrating the directional evolution of a system as described by the second law of thermodynamics.

In essence, the change in entropy within the combined system is driven by the uneven distribution of dynamic energy between the two initial systems. When the systems interact, a restoration process is triggered. At equilibrium, a new ratio among different forms of dynamic energy is established throughout the combined system.

This unidirectional increase in entropy is not confined to the specific example discussed above but reflects a general principle embodied in the definition of entropy. Consider two systems at temperatures T_1 and T_2, with $T_2 > T_1$. When the systems interact, an infinitesimal amount of heat dQ flows from the higher-temperature system to the lower-temperature one. According to Definition (9), the total change in entropy of the two-system ensemble is given by:

$$(12) \qquad dS = \frac{dQ}{T_1} - \frac{dQ}{T_2} = \frac{dQ}{T_1 T_2}\left(T_2 - T_1\right) > 0$$

Since $T_2 > T_1$, the total entropy of the combined system always increases, indicating that the system naturally evolves toward higher entropy. The system's entropy reaches its maximum at equilibrium, where a new stable distribution is established among the different forms of dynamic energy. In other words, a state of maximum entropy corresponds to the system's equilibrium, with a stable distribution of energy across its various forms. Therefore, the second law of thermodynamics describes only a subset of the processes governed by the **Restoration Principle**.

Entropy Can Decrease

It is important to note that the derivation of Expression (12) rests on the empirical assumption that heat always flows down a temperature gradient. This assumption, however, is not universally valid and can break down in systems dominated by attractive forces, particularly on large spatial scales. While a small volume of gas may remain well mixed and nearly uniform, a planetary-scale atmosphere is typically stratified into distinct layers with different temperatures. Although temperature gradients can drive localized convection, the large-scale structure of Earth's atmosphere is governed primarily by gravity. Consequently, the entropy of the real atmosphere does not attain its maximum value, highlighting a limitation in the direct application of the second law of thermodynamics to such systems.

Similarly, the evolution of the universe illustrates the local aggregation of energy rather than its uniform dispersion, contrary to the traditional "heat death" scenario suggested by the second law of thermodynamics. Under the influence of gravity, stars, galaxies, and other celestial bodies are thought to form from vast cosmic gas clouds with initially low energy density. As these clouds collapse gravitationally, matter concentrates toward their centers of mass, giving rise to highly concentrated energy structures and more organized systems such as planets, stars, neutron stars, and black holes. These processes can be seen as the inverse of the dispersive tendency described by the second law.

Decreases in entropy are not uncommon in systems governed by attractive forces, particularly in large-scale systems dominated by gravity. Under sufficiently strong attractive interactions, such as Coulomb forces, this kind of "anti-gradient" energy flow can also occur at small scales. Moreover, the evolution of systems in the universe does not always follow a single, unidirectional path as suggested by the second law; it can also exhibit periodic or reversible behavior, which the second law does not fully account for.

The Difference Between Classical and Statistical Entropy

In addition to the classical definition of entropy (9), Ludwig Boltzmann introduced a statistical formulation based on the probability of microstates. A microstate represents a specific configuration of all components of a system. Under this statistical definition, entropy is expressed as:

$$(13) \qquad S = -k_B \sum_{i=1}^{N} p_i ln(p_i)$$

In this definition, k_B denotes the Boltzmann constant, p_i represents the probability of each microstate, and the summation is taken over all possible microstates (N) of the system. For clarity, if we assume that all microstates are equally probable, the statistical definition can be simplified as:

$$(14) \qquad S = k_B ln(N)$$

Despite this simplification, the formulation retains sufficient generality to enable a clear comparison with the classical definition of entropy. Because it represents a specific case of the general statistical definition, consistency between the general formulation and the classical definition should also be reflected in this simplified scenario. Conversely, if inconsistency arises even in this special case, it would call into question the validity of the general equivalence.

Although Boltzmann is often credited with establishing an equivalence between the statistical and classical definitions of entropy through analyses of microscopic behavior, the two definitions are fundamentally distinct. The classical definition characterizes the macroscopic distribution of a system's energy, whereas the statistical definition describes the probabilistic distribution of its microscopic constituents. Under the statistical definition, a system evolves toward the most probable configuration of its components, that is, statistical equilibrium; under the classical definition, it evolves toward a stable distribution of energy, namely, energetic equilibrium. While these equilibria may coincide in systems dominated by repulsive interactions, where components naturally disperse, the two definitions embody fundamentally different perspectives and can diverge in systems governed primarily by attractive forces.

Furthermore, if the two definitions were truly equivalent, their predictions should agree under all conditions. However, a divergence becomes apparent as a system's temperature approaches 0 K. In this limit, the system occupies a single accessible microstate—the ground state—resulting in zero entropy according to both statistical Definitions (13) and (14). In contrast, the classical definition predicts markedly different behavior: as the temperature approaches 0 K, its reciprocal diverges, causing the entropy in Definition (9) to diverge as well. Specifically, within the simplified model described by Equation (11), the entropy tends toward negative infinity as the temperature approaches 0 K. This discrepancy highlights a fundamental distinction between the classical and statistical formulations of entropy.

Summary

Since all entities in the universe exist as energy in one form or another, including matter itself, the evolution of any system is fundamentally governed by the transformation of its energy. This evolution is driven by the system's inherent tendency toward equilibrium, facilitated by the continuous exchange among different forms of dynamic energy. Equilibrium corresponds to a state in which energy is stably distributed among these forms—a concept formalized by the **Restoration Principle**, which encapsulates the fundamental driving force behind all processes. Viewing thermodynamic phenomena from this perspective enables a deeper examination, clarifies common misconceptions, and fosters an accurate understanding of system dynamics.

Revision History

- 12/28/2025: Initial Post on Zenodo
- 01/02/2026: Revision for Conceptual Clarity
- 01/08/2026: Revised Impactrans Section

Links to Summaries of Related Articles

- https://cs.stanford.edu/people/zjl/abstract.html, PDF
- https://sites.google.com/view/zjl/abstracts, PDF
- https://xenon.stanford.edu/~zjl/abstract.html, PDF
- https://doi.org/10.5281/zenodo.17967154, PDF

Related 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)
- Can Temperature Represent Average Kinetic Energy? (PDF: DOI) (中文: DOI)
- The Nature of Absolute Zero Temperature (PDF: DOI) (中文: DOI)
- The Triangle of Energy Transformation (PDF: DOI) (中文: DOI)
- Is Thermal Expansion Due to Particle Vibration? (PDF: DOI) (中文: DOI)
- Superfluids Are Not Fluids (PDF: DOI) (中文: DOI)
- Why a Phase Transition Temperature Remains Constant (PDF: DOI) (中文: DOI)
- What Causes Friction to Produce Heat? (PDF: DOI) (中文: DOI)
- The Easiest Way to Grasp Entropy (PDF: DOI) (中文: DOI)
- Entropy Can Decrease (PDF: DOI) (中文: DOI)
- The Restoration Principle (PDF: DOI) (中文: DOI)

Bibliography

- Armin, H. (1971). “The Genesis of Quantum Theory (1899–1913)”. Translated by Claude W. Nash, ed., Cambridge, MA.
- Bardeen, J.; et al. (1957). “Theory of Superconductivity”. *Physical Review*. **108**. p. 1175. doi:10.1103/physrev.108.1175.
- Bednorz, J.G. & Müller, K.A. (1986). "Possible high T_c Superconductivity in the Ba−La−Cu−O System". *Z. Phys. B*. **64** (1): 189–193. doi:10.1007/BF01303701.
- Boltzmann, L. (1895). “On Certain Questions of the Theory of Gases”. *Nature*, **51** (1322): 413-415.

- Brush, S.G.; Everitt, C.W.F. (1969). “Maxwell, Osborne Reynolds, and the Radiometer”. *Historical Studies in the Physical Sciences*, vol. 1, 1969, pp. 105–125.
- Carey, V. (1999). “Statistical Thermodynamics and Microscale Thermophysics”. Cambridge University Press.
- Clausius, R. (1857). “Über die Art der Bewegung, welche wir Wärme nennen”. *Annalen der Physik und Chemie* (in German). **176** (3): 353–79. Bibcode:1857AnP...176..353C. doi:10.1002/andp.18571760302. Facsimile at the Bibliothèque nationale de France (pp. 353–79).
- Crookes, W. (1874). “On Attraction and Repulsion Resulting from Radiation”. *Philosophical Transactions of the Royal Society of London*. **164**: 501–527. doi:10.1098/rstl.1874.0015.
- De Podesta, M.; et al. (2013). “A Low-uncertainty Measurement of the Boltzmann Constant”. *Metrologia*, **50** (4): S213-S216, BIPM & IOP Publishing Ltd.
- Draper, J.W. (1847). “On the Production of Light by Heat”. *London, Edinburgh, and Dublin Philosophical Magazine and Journal of Science*, series 3, **30**: 345–360.
- Drozdov, A.; et al. (2015). "Conventional Superconductivity at 203 Kelvin at High Pressures in the Sulfur Hydride System". *Nature*. **525** (2–3): 73–76. arXiv:1506.08190.
- Drude, P. (1900). "Zur Elektronentheorie der Metalle". *Annalen der Physik*. **306** (3): 566–613. doi:10.1002/andp.19003060312.
- Dugdale, J.S. (1996). “Entropy and Its Physical Meaning”. *Taylor & Francis*. p. 13. ISBN 978-0-7484-0569-5.
- Einstein, A. (1905), “Zur Elektrodynamik bewegter Körper”, Annalen der Physik (in German), **322** (10): 891–921, Bibcode:1905AnP...322..891E, doi:10.1002/andp.19053221004 (English translation).
- Einstein, A. and Stern, O. (1913). “Some Arguments for the Assumption of Molecular Agitation at Absolute Zero”. *Ann. Phys.*, **40** (551) 551-560.
- Einstein, A. (1993). “The Collected Papers of Albert Einstein. 3”. *English translation by Beck, A. Princeton University Press*.
- Eremets, M.I.; et al. (2022). "High-Temperature Superconductivity in Hydrides: Experimental Evidence and Details". *Journal of Superconductivity and Novel Magnetism*. **35** (4): 965–977. arXiv:2201.05137. doi:10.1007/s10948-022-06148-1.
- Gay-Lussac, J.L. (1802), “Recherches sur la Dilatation des gaz et des vapeurs”. *Annales de Chimie*, **XLIII**: 137. English translation (extract).
- Gibbs, P. (1996). “How Does a Light-Mill Work?” *Usenet Physics FAQ*. Retrieved 8 August 2014.
- Goldstein, H. (1980). “Classical Mechanics”. (2nd edition). Addison-Wesley.
- Gourgy, J.B. (1979), “Charles the Obsecure”. *Isis* **70** (254): 576-579.
- Han, L.H.; et al. (2010). “Light-Powered Micromotor Driven by Geometry-Assisted, Asymmetric Photon-heating and Subsequent Gas Convection”. *Applied Physics Letters*. **96** (21): 213509(1–3). doi:10.1063/1.3431741.
- Itano, W.M.; et al. (2015). “Early Observations of Macroscopic Quantum Jumps in Single Atoms”. *International Journal of Mass Spectrometry*. **377**: 403. doi:10.1016/j.ijms.2014.07.005.
- Jaynes, R. (2019). “Radiometer in the Microwave”. *YouTube channel* @jaynesnetwork, https://www.youtube.com/watch?v=OGX0-C1FXYA.
- Kapitza, P. (1938). “Viscosity of liquid helium below the λ-point”. *Nature*. **141** (3558), 74. Bibcode:1938Natur.141...74K. doi:10.1038/141074a0. S2CID 3997900.
- Knizhnik, K. (2016). “Derivation of the Stefan–Boltzmann Law”. *Johns Hopkins University – Department of Physics & Astronomy*.
- Kraftmakher, Y. (2014). *Experiments and demonstrations in physics* (2nd ed.). Singapore: World Scientific. p. 179. ISBN 9789814434904.

- Krönig, A. (1856). “Grundzüge einer Theorie der Gase”. *Annalen der Physik und Chemie* (in German). **99** (10): 315–22. Bibcode:1856AnP...175..315K. doi:10.1002/andp.18561751008. Facsimile at the Bibliothèque nationale de France (pp. 315–22).
- Leite, F. L.; et al. (2012). “Theoretical models for surface forces and adhesion and their measurement using atomic force microscopy”. *International Journal of Molecular Sciences*. **13** (12), 12773–856. doi:10.3390/ijms131012773. PMC 3497299. PMID 23202925.
- Machado, A.; et al. (2012). “Defect Structure Versus Superconductivity in MeB2 Compounds (Me = Refractory Metals) and One-Dimensional Superconductors”, doi:10.5772/48625.
- Maxwell, J.C. (1875). “The Question of the Atom”. *J. Chem. Soc.* (London), Facsimile published in Mary Jo Nye, **1875** (28): 493-508.
- Maxwell, J.C. (1879). “On Stresses in Rarefied Gases Arising from Inequalities of Temperature”. *Philosophical Transactions of the Royal Society of London*. **170**: 231–256. doi:10.1098/rstl.1879.0067.
- Mayhew, K.W. (2017). “A New Perspective for Kinetic Theory and Heat Capacity”. *Prog. in Phys.,* **13** (3): 165-173.
- Narimanov, E.E.; Smolyaninov, I.I. (2012). “Beyond Stefan–Boltzmann Law: Thermal Hyper-Conductivity”. *Conference on Lasers and Electro-Optics* 2012. OSA Technical Digest. Optical Society of America. pp. QM2E.1. doi:10.1364/QELS.2012.QM2E.1.
- Osborne, R. (1879). “On Certain Dimensional Properties of Matter in the Gaseous State”. *Royal Society Phil. Trans.*, Part 2.
- Pais, A. (1982). “ Subtle is the Lord”. Oxford University Press. Oxford UK.
- Planck M. (1901). “On the Law of the Energy Distribution in the Normal Spectrum”. *Ann. Phys.*, **4** (553): 1-11.
- Planck, M. (1914). “The Theory of Heat Radiation*”. Masius, M. (transl.)* (2nd ed.). P. Blakiston's Son & Co. OL 7154661M.
- Planck, M. (1915). “Eight Lectures on Theoretical Physics*”. Wills, A. P. (transl.). Dover Publications.*
- Rayleigh, J.W.S. (1900). “The Law of Partition of Kinetic Energy”. *Phil. Mag.*, **49**: 98–118.
- Reif, F. (1965). “Fundamentals of Statistical and Thermal Physics”. *McGraw-Hill*, New York.
- Schilling, A.; et al. (1993). "Superconductivity above 130 K in the Hg–Ba–Ca–Cu–O System". *Nature*. **363** (6424): 56–58. doi:10.1038/363056a0.
- Souza, E.; et al. (2006). “Improvement of Metallic Joint Electrical Conductivity Using a Novel Conductive Paste Produced from Recycled Residues”. *Rev. Esc. Minas* Vol. 59, No. 2.
- Takahashi, H.; et al. (2008). "Superconductivity at 43 K in an Iron-based Layered Compound LaO$_{1-x}$F$_x$FeAs". *Nature*. **453** (7193): 376–378. doi:10.1038/nature06972. PMID 18432191.
- Thomson, W. (1904). “Historical and Philosophical Perspectives”. *Baltimore Lectures*. Baltimore: Johns Hopkins University Press. Sec. 27.
- Tsekov, R. (1995). "Brownian Motion of Molecules: the Classical Theory". *Ann. Univ. Sofia*. **88**: 57. arXiv:1005.1490. Bibcode:1995AUSFC..88...57T. The behavior of a Brownian particle is quite irregular and can be described only in the frames of a statistical approach.
- Tuckerman, M.E. (2010). “Statistical Mechanics: Theory and Molecular Simulation”. (1st ed.). P. 87. ISBS 978-0-19-852526-4.
- Vaidya, R.; et al. (2003). “Effect of Pressure on Electrical Resistance of WSe2 Single Crystal”. Pramana -- *Journal of Physics*, Vol. 61, No. 1, pp. 183-186.
- Vertical Learning Labs. (2011). “*Molecular Speed Distribution*”. Introduction to Chemistry, Unit 1.5.
- Vijay, R.; et al. (2011). “Observation of Quantum Jumps in a Superconducting Artificial Atom”. *Physical Review Letters*. **106** (11): 110502. arXiv:1009.2969. doi:10.1103/PhysRevLett.106.110502. PMID 21469850.

- Worrall, J. (1982), “The Pressure of Light: The Strange Case of the Vacillating 'Crucial Experiment'”. *Studies in History and Philosophy of Science*, **13** (2): 133–171, doi:10.1016/0039-3681(82)90023-1.
- Wu, M.K.; et al. (1987). "Superconductivity at 93 K in a New Mixed-Phase Y-Ba-Cu-O Compound System at Ambient Pressure". *Physical Review Letters*. **58** (9): 908–910. doi:10.1103/PhysRevLett.58.908. PMID 10035069.
- Yarris, L. (2010). “Nano-Sized Light Mill Drives Micro-sized Disks”. *Physorg. Retrieved* 6.