Can Temperature Represent Average Kinetic Energy?

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Abstract

A common misconception is that temperature directly represents the average kinetic energy of particle motion. This view contributes to many misunderstandings in thermodynamics, particularly regarding the limitations of the kinetic theory of gases. Notably, the theory consistently underpredicts the specific heat of polyatomic gases, a discrepancy that has sparked scientific debate for centuries due to a fundamental misinterpretation of what temperature truly represents. The three forms of *dynamic energy*, including potential energy, kinetic energy, and radiative energy, continuously interconvert into each other and play different roles in a thermodynamic system. This study demonstrates that the energy unaccounted for by kinetic theory correlates with the number of molecular bonds, indicating that a significant portion of specific heat is allocated to increase the potential energy stored in these bonds. In practical measurement, thermometers, especially infrared ones, reflect radiation rather than kinetic energy. Planck's law, Stefan-Boltzmann's law, and Wien's displacement law collectively establish a direct relationship between temperature and radiative emission, which originates from potential energy components within a system. Radiative absorption can excite electrons to higher orbitals, effectively storing energy as potential energy. Similarly, molecular bonds function like springs, storing energy when stretched. At higher temperatures, these expanded bonds occupy more space, providing a more accurate explanation for thermal expansion than kinetic vibration alone. Thus, temperature is more accurately understood as a measure of radiative energy, closely related to a system's potential energy. Kinetic energy is indirectly connected to temperature via transformations with other forms of dynamic energy facilitated by *transimpacts* and *impactranses*. This connection breaks down during phase transitions: although kinetic energy changes as molecular bonds are broken or formed, the temperature remains constant, reflecting the system's radiation level tied to potential components rather than its kinetic one. This plateau occurs because potential energy reaches its maximum, constrained by bond extension limits just before rupture.

Introduction

As one of the seven fundamental quantities in physics, temperature is a critical property of systems across all natural sciences, particularly in thermodynamics. But what temperature truly represents at the microscopic scale has not been well understood for a long time. The widely accepted textbook explanation states that temperature corresponds to the average kinetic energy of the particles in a system.[1] Wikipedia provides a similar definition, as shown in Figure 0. However, this interpretation is challenged by how temperature is often measured in practice. For instance, infrared thermometers determine temperature by detecting the radiation emitted by an object without any physical contact. In such cases, there is no direct interaction with the particles themselves, and therefore no direct measurement of their kinetic energy. This disconnect raises fundamental questions about the physical nature of temperature and what it truly quantifies.

Temperature

From Wikipedia, the free encyclopedia

This article is about the physical quantity. Fc
(disambiguation).

Temperature quantitatively expresses the attribute of hotness or coldness. Temperature is measured with a thermometer. It reflects the average kinetic energy of the vibrating and colliding atoms making up a substance.

Figure 0: Screenshot of Wikipedia's definition of *temperature* as it appeared on October 22, 2025.

Furthermore, phase transitions often occur at a constant temperature, even as the system undergoes substantial internal changes. During these processes, molecular bonds are broken and particles are released or reorganized, reflecting significant energetic transformations, particularly in the dynamics of particle motion. If temperature were strictly a measure of the average kinetic energy of such motion, it should vary accordingly during these events. This apparent paradox exposes a deeper inconsistency in the conventional interpretation of temperature and underscores the need to reevaluate its fundamental physical meaning. A more accurate understanding could clarify persistent misconceptions in thermodynamics, including those related to thermal expansion and Brownian motion.

A fundamental misconception about the physical meaning of temperature also lies at the core of the longstanding discrepancy between the specific heat values predicted by the kinetic theory of gases and those measured experimentally—a debate that has persisted for centuries. By analyzing the specific heat of various gases, this study reveals a direct relationship between the deficit of specific heat and the number of molecular bonds, suggesting that temperature is more directly associated with potential and radiative energy. In contrast, its connection to kinetic energy appears to be only indirect. This revised perspective provides a coherent resolution to several enduring problems in thermodynamics and advances a more comprehensive understanding of the true physical nature of temperature.

Limits of the Kinetic Theory

If kinetic energy were the sole determinant of a system's temperature, changes in temperature and kinetic energy could be used to predict the specific heat in gas systems using the kinetic theory. The kinetic theory correlates temperature with the molecular speed of ideal gases, at least within the optimal ranges of assumed temperatures and pressures:[2]

(1) $$v = \sqrt{\frac{3k_B T}{m}}$$

Here, m denotes the mass of a gas molecule, T indicates the absolute temperature, k_B represents the Boltzmann constant ($1.38x10^{-23}\ J \cdot K^{-1}$),[3] and v is the predicted average speed of the gas molecules. Consequently, the average kinetic energy due to the motion of a molecule in an ideal gas can be derived from the kinetic theory as follows:

$$(2) \qquad K = \frac{1}{2}mv^2 = \frac{3}{2}k_B T$$

Thus, the kinetic energy K of the system is proportional to its absolute temperature T. This relationship likely underlies the widespread misconception that temperature directly represents the average kinetic energy of the particle motion in a system. Specific heat capacity is the energy required to raise the temperature of a unit mass of a given substance by one degree.[4-5] Accordingly, the specific heat of ideal gases can be predicted using the kinetic theory of gases by dividing both sides of Equation 2 by the molecule mass and temperature, $m \cdot T$:

$$(3) \qquad C_k = \frac{K}{mT} = \frac{3k_B}{2m}$$

Molar specific heat of a substance is the specific heat of a mole of the substance. By multiplying the specific heat by the molar mass, the predicted molar specific heat can be expressed as:

$$(4) \qquad C_t = mN_A C_k = \frac{3}{2}k_B N_A = \frac{3}{2}R$$

Here, N_A denotes Avogadro's number $6.022 \times 10^{23}\ mol^{-1}$) and R is the molar gas constant ($8.314\ J \cdot K^{-1} \cdot mol^{-1}$). According to the kinetic theory, the molar specific heat capacity of an ideal gas should be constant ($12.471\ J \cdot K^{-1} \cdot mol^{-1}$) and identical for all ideal gases. These theoretical values can be directly compared with experimental data. To ensure that the measurements align with the assumptions of the kinetic theory, specific heat values were collected within optimal ranges of pressure (below *1.5 atm*) and temperature (*280–380 K*). Table 1 presents the measured isochoric specific heat capacities for 38 ideal gases in column C_v, along with the corresponding molar specific heats in column C_m. These empirical values are compared with the theoretical specific heats predicted by kinetic theory, listed as C_k (specific heat) and C_t (molar specific heat), derived from Equations (3) and (4).

Gas Name	Molecule	Mass	C_k	C_v	C_t	C_m	Atoms	Bonds
Helium	He	4.003	3.1152	3.1156	12.471	12.4717	1	0
Neon	Ne	20.18	0.6179	0.6180	12.471	12.4712	1	0
Argon	Ar	39.95	0.3121	0.3122	12.471	12.4724	1	0
Krypton	Kr	83.8	0.1488	0.151	12.471	12.6538	1	0
Xenon	Xe	131.3	0.0950	*0.095*	12.471	12.4735	1	0
Hydrogen	H_2	2.016	6.1855	10.183	12.471	20.5289	2	1
Hydrogen Chloride	HCl	36.45	0.3421	0.57	12.471	20.7765	2	1
Hydroxyl	OH	17	0.7335	1.27	12.471	21.5900	2	1
Chloride	Cl_2	70.9	0.1759	0.36	12.471	25.5240	2	1
Bromine	Br_2	159.8	0.0780	0.2	12.471	31.9600	2	1
Oxygen	O_2	32	0.3897	0.658	12.471	21.0560	2	2

Steam	H_2O	18.02	0.6920	1.4108	12.471	25.4226	3	2
Air	$8N_2+2O_2$	28.96	0.4306	0.718	12.471	20.7933	2	2.8
Nitrogen	N_2	28.01	0.4452	0.743	12.471	20.8114	2	3
Carbon Monoxide	CO	28.01	0.4452	0.744	12.471	20.8394	2	3
Ammonia	NH_3	17.03	0.7322	1.66	12.471	28.2698	4	3
Hydrogen Peroxide	H_2O_2	34.01	0.3667	*1.1*	12.471	37.4110	4	3
Methane	CH_4	16.04	0.7774	1.7354	12.471	27.8358	5	4
Carbon Dioxide	CO_2	44.01	0.2833	0.657	12.471	28.9146	3	4
Sulfur Dioxide	SO_2	64.06	0.1947	0.51	12.471	32.6706	3	4
Carbon Disulphide	CS_2	76.13	0.1638	0.55	12.471	41.8715	3	4
Chloroform	$CHCl_3$	126.37	0.0987	0.55	12.471	69.5035	5	4
Acetylene	C_2H_2	26.04	0.4789	1.37	12.471	35.6748	4	5
Methanol	CH_3OH	32.03	0.3893	1.53	12.471	49.0059	6	5
Ethylene	C_2H_4	28.05	0.4446	1.2518	12.471	35.1130	6	6
Ethane	C_2H_6	30.07	0.4147	1.4897	12.471	44.7953	8	7
Ethanol	C_2H_5OH	46.07	0.2707	1.67	12.471	76.9369	9	8
Propene	C_3H_6	42.08	0.2963	1.31	12.471	55.1248	9	9
Propane	C_3H_8	44.1	0.2828	1.4909	12.471	65.7487	11	10
Acetone	$(CH_3)_2CO$	58.08	0.2147	1.32	12.471	76.6656	10	10
Isobutene	C_4H_8	56.11	0.2222	*1.374*	12.471	77.0951	12	12
Butane	C_4H_{10}	58.12	0.2146	1.5734	12.471	91.4460	14	13
Diethyl Ether	$(C_2H_5)_2O$	74.24	0.1680	1.95	12.471	144.7680	15	14
Benzene	C_6H_6	78.11	0.1596	0.99	12.471	77.3289	12	15
Isopentane	C_5H_{12}	72.15	0.1728	*1.548*	12.471	111.6882	17	16
n-Hexane	C_6H_{14}	86.18	0.1447	*1.564*	12.471	134.7855	20	19
n-Heptane	C_7H_{16}	100.2	0.1245	*1.573*	12.471	157.6146	23	22
Octane	C_8H_{18}	114.2	0.1092	1.6385	12.471	187.1167	26	25

Table 1: Discrepancy of isochoric-specific heat between measurements and the predictions in the kinetic theory of gases. Column C_k is the theoretical specific heat projected using Equation (3). Column C_v is measured at a constant volume with an initial pressure of *1 atm* and temperature of *300 K*. The mass unit is 1.66×10^{-27} *kg,* and the unit for specific heat is *kJ/kg•K*. Column C_m is the measured molar-specific heat in *J/mol•K* corresponding to C_v, compared with the theoretical molar-specific heat in C_t. The Atoms and Bonds columns show the number of atoms and bonds in the corresponding gas molecules.

Both the predicted specific heat and molar specific heat are consistently underestimated by the kinetic theory, with the discrepancy increasing alongside molecular complexity. While the ideal gas law has been validated through extensive applications over the past two centuries, and predictions from kinetic theory, such as molecular speeds, correlate well with observed sound propagation in gases, a critical inconsistency remains: if the kinetic theory is fundamentally sound, why do its predictions for specific heat diverge from experimental measurements, especially for complex molecules? This question has remained a longstanding puzzle in the field of thermodynamics for centuries.

To resolve the discrepancies in the kinetic theory, additional components were incorporated into the kinetic energy model: translational energy for particle motion, vibrational energy for particle vibrations, and rotational energy for particle rotation. Initially, the energy predicted by the theory in Equation 3 was assumed to account only for translational motion, with vibrational and rotational contributions considered separately. In response, Maxwell proposed the equipartition of energy among these three modes,[6] effectively multiplying the predicted energy by a factor of three. However, even with this scaling, the model still fails to account for the discrepancies, particularly in the case of molar-specific heat, which increases with molecular complexity, as illustrated in column C_m of Table 1 and Figure 1.

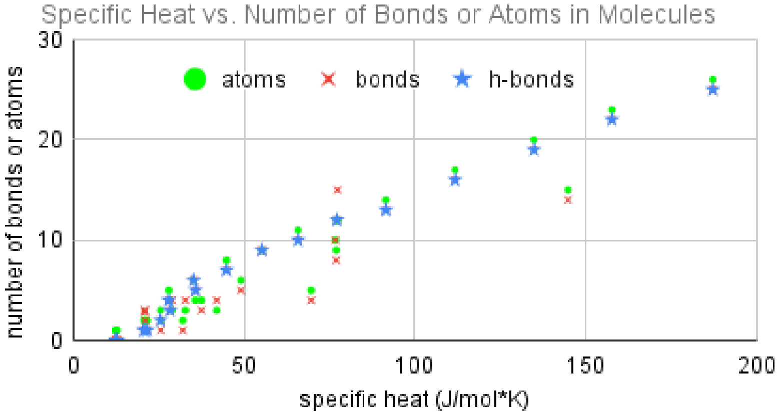


Figure 1: Correlation between measured isochoric molar-specific heat and molecular complexity. The horizontal axis represents the molar-specific heat values from column C_m in Table 1, while the vertical axis indicates molecular complexity, expressed as the number of atoms or bonds from the columns labeled "Atoms" or "Bonds". Green dots indicate the number of atoms in a gas molecule. Blue stars denote the number of hydrogen bonds in aliphatic hydrocarbons (h-bonds), while red crosses represent other types of bonds.

To address this discrepancy, Boltzmann refined the equipartition principle by introducing the concept of degrees of freedom, proposing that energy in a gas is equally distributed among all accessible degrees of freedom.[7–8] Each additional atom in a gas molecule increases from three up to seven degrees of freedom, depending on the complexity of the molecules. Consequently, the predicted specific heat is further scaled up according to the number of atoms within each molecule. To reconcile the model with experimental results, various modifications have been proposed subsequently.[9-10]

To understand the inconsistency, numerous explanations have been proposed to account for the failure of the equipartition principle. Boltzmann suggested that gases might not be in thermal equilibrium.[11] Planck and Einstein introduced the idea of a zero-point harmonic oscillator to explain the discrepancy.[12-13] Kelvin eventually concluded that the equipartition assumption might be incorrect,[14-15] a view later acknowledged by Einstein.[16-17]

Missing Specific Heat Is Stored in Bonds

To escape this conceptual rabbit hole, consider a different question: Is specific heat entirely attributable to kinetic energy? If that were the case, gases, with their significantly greater freedom of molecular motion, should exhibit higher specific heat capacities than liquids. To test this idea, Table 2 compares the specific heat values of various substances in both their liquid and gaseous phases. Contrary to expectations based solely on kinetic energy, all listed substances exhibit higher specific heat in the liquid phase. For example, the specific heat of water is nearly three times that of steam.

Name	Molecule	Mass	Liquid (kJ/kg*K)	Gas (kJ/kg*K)	Liquid/Gas Ratio
Ether	$(C_2H_5)_2O$	74.24	2.21	2.01	1.10
Ethanol	C_2H_5OH	46.07	2.3	1.88	1.22
Propane	C_3H_8	44.1	2.4	1.67	1.44
Acetone	$(CH_3)_2CO$	58.08	2.15	1.47	1.46
Carbon Disulfide	CS_2	76.13	0.992	0.67	1.48
Benzene	C_6H_6	78.11	1.8	1.09	1.65
Chloroform	$CHCl_3$	126.37	1.05	0.63	1.67
Bromine	Br_2	159.8	0.47	0.25	1.88
Propene	C_3H_6	42.08	2.85	1.5	1.90
Ammonia	NH_3	17.03	4.6	2.19	2.10
Water	H_2O	18.02	4.19	1.411	2.97

Table 2: Comparison of isobaric specific heat (C_p) between liquids and gases. The specific heat of liquids and gases is from The Engineering ToolBox.

This observation challenges the notion that specific heat is solely attributed to kinetic energy. Is this kinetic assumption fundamentally flawed? Beyond the difference in molecular mobility between gases and liquids, a key distinction is the presence of intermolecular bonds in liquids. Potential energy arises naturally from interactions between particles, such as electrostatic forces between electrons and nuclei. It is well established that chemical bonds store energy, typically measured as enthalpy. Similarly, intermolecular bonds can also act as reservoirs of energy. Like springs, these bonds store energy when stretched. As work is done to increase the distance between bonding atoms or molecules, that energy is retained as potential energy within the bonds. This perspective implies that intermolecular bonds between liquid molecules may be responsible for storing a significant fraction of the energy for the higher specific heat observed in liquids. The intermolecular bond capacity for storing energy may surpass the contribution from particle vibration, particularly since the kinetic energy of molecules in a liquid is relatively low due to high viscosity when compared with a gas.

In practice, specific heat is a valuable tool for calculating the enthalpy associated with chemical bonds. Bond enthalpy and enthalpy of formation are closely related concepts that quantify bond strength by indicating the energy stored under standard conditions (*1 atm* and typically *298.15 K*). However, enthalpy values are often needed at temperatures other than the standard reference. To estimate enthalpy at varying temperatures, chemists use the isobaric specific heat capacity (C_p) and apply the following equation:

$$\Delta H = \int_{t_1}^{t_2} C_p dt \qquad (5)$$

Here, ΔH denotes the enthalpy difference of a substance between two temperatures: t_2 is the target temperature for which the new enthalpy is to be calculated, and t_1 is the standard temperature, typically *298.15 K*, at which enthalpy values are commonly available. This formulation underscores that bond energy is a well-established concept with experimentally validated accuracy.

With this insight, let’s revisit the discrepancies between experimental specific heat values and those predicted by the kinetic theory. Two key observations emerge: (1) For gases composed of monatomic molecules, the predictions of the kinetic theory closely match experimental data, regardless of atomic mass, with an accuracy of approximately *99%*, as shown in the top five rows of Table 1. (2) In contrast, the theory fails to accurately predict the specific heat of gases composed of polyatomic molecules. In these cases, the predicted values are consistently lower than the measured ones, and the discrepancy increases with molecular complexity. For example, the predicted specific heat for octane accounts for only *7%* of the measured value, indicating a *93%* error.

In essence, kinetic theory performs well for gases composed of molecules without internal bonds, indicating the theory is accurate in this case. More importantly, the discrepancy in predicted versus observed specific heat shows a strong correlation with the number of molecular bonds. This trend becomes particularly evident when comparing the molar-specific heat values (column *C*□) with the corresponding bond counts. To illustrate this relationship, Figure 2 presents the correlation between molar-specific heat and the number of molecular bonds. To reduce variability in energy storage due to differences in bond types, the analysis is limited to aliphatic hydrocarbons.

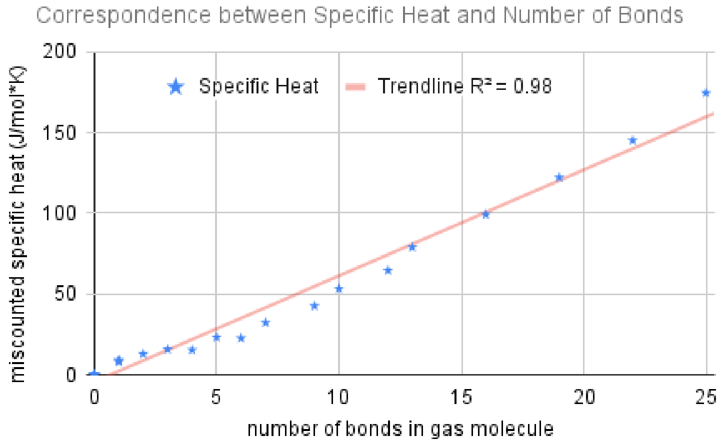


Figure 2: Correspondence between molar-specific heat and the number of bonds. The points in the figure are the selected gases with bonds only associated with aliphatic hydrocarbons, as indicated in Figure 1. The miscounted molar-specific heat is the measurement listed in column C_m of Table 1, offsetting the molar-specific heat predicted in the kinetic theory (C_m-C_t). The slope of the fitted trendline is about *6.56* (*J/mol•K•bond*) with the coefficient of determination (R^2) up to *0.98*.

The following key implications emerge from these observations. The specific heat added to a system is primarily distributed between kinetic and potential energy. The portion predicted by kinetic theory corresponds to the kinetic energy contribution and remains accurate across different gases, regardless of molecular mass or complexity. This suggests that equipartition-based adjustments to kinetic theory may be unnecessary. The remaining energy is primarily allocated to increasing the potential energy stored in molecular bonds, and this portion increases with the number of bonds. Consequently, the temperature cannot be fully represented by kinetic energy alone, as a substantial part of the specific heat is associated with potential energy changes. A small fraction of the energy may also contribute to elevated radiation levels or the excitation of electrons to higher average orbitals; however, this contribution is negligible within the limits of experimental measurement accuracy.

The Connection between Temperature and Potential Energy

The preceding observations seriously challenge the traditional association between temperature and kinetic energy. This raises a critical question: Does temperature more directly represent a system's potential energy? In practical applications, many thermometers, particularly infrared thermometers, do not measure kinetic or potential energy directly. Instead, they detect radiative energy. The behavior of ideal radiative emission is described by Planck's law, which defines the spectral distribution of radiation as a function of wavelength and temperature.[18–20] When integrated over all wavelengths, Planck's law yields the total radiative power P emitted by a perfect blackbody. This power is proportional to the fourth power of absolute temperature, as expressed by the Stefan–Boltzmann law:[21–22]

$$(6) \qquad P = pT^4$$

Here, p is the Stefan-Boltzmann constant. By reorganizing this equation, temperature can be expressed in terms of the emission power of a target system:

$$(7) \qquad T = \sqrt[4]{\frac{P}{p}}$$

This relationship indicates that temperature can be inferred from the radiation emitted by a system, forming the theoretical basis for temperature measurement devices. Infrared thermometers, for instance, determine temperature by detecting the intensity of radiation emitted by a target body, without requiring direct contact.

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. In other words, as the temperature increases, the peak of the radiation curve moves to shorter wavelengths:

$$(8) \qquad \lambda = \frac{b}{T}$$

Here, λ represents the peak wavelength of the radiation curve, and b is Wien's displacement constant. This relationship can also be derived by finding the peak wavelength of Planck's radiation curve, specifically, by taking its derivative for wavelength and solving for the maximum. Rearranging the resulting equation allows the temperature to be expressed in terms of the peak radiation wavelength:

$$(9) \qquad T = \frac{b}{\lambda}$$

This relationship allows temperature to be determined by measuring the peak wavelength of radiation emitted by a system. Wien's displacement law, which describes this inverse relationship between temperature and peak emission wavelength, is evident in everyday observations. For instance, an experienced baker can tell oven temperature by observing the color of the flame. Similarly, the Sun's surface temperature, approximately *5,778 K*, corresponds to a peak emission near *500 nm*, which lies within the visible spectrum. These examples illustrate how temperature is fundamentally linked to the radiative energy emitted by a source.

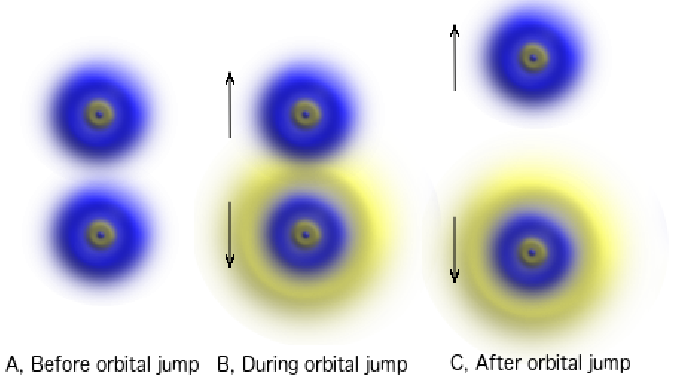
Radiation is closely linked to the potential energy components within a system. When a body accelerates, changes in its force fields can lead to the emission of energy as radiation.[23] For example, an accelerating charged particle, such as an electron transitioning between orbitals, emits energy in the form of radiation, including visible light. Conversely, when radiation is absorbed, an electron may become excited to a higher orbital, storing the absorbed energy as potential energy. Intermolecular and intramolecular bonds can likewise absorb and emit radiation through changes in their potential energy states. For instance, thermal energy is released or absorbed during chemical reactions as intramolecular bonds change. Similarly, during phase transitions, changes in intermolecular bonds result in the absorption or release of latent heat.

In essence, radiative energy originates from the transformation of potential energy stored in bonds or orbital electrons, underscoring the intrinsic connection between these energy forms. This perspective helps explain the observed correlation between bond structure and specific heat in gaseous systems. Consequently, temperature is linked to potential energy through the radiative processes of a system's potential components. As a result, the temperature is more closely associated with a system's radiation level and internal potential energy than with its kinetic energy.

The Origin of Kinetic Energy

Thus far, we have examined the relationships among temperature, radiation, and potential energy. This raises a fundamental question: What role does kinetic energy play in a system, and how does it originate, particularly in the forms of particle motion and vibration? Despite its fundamental importance, this question has received little attention in the scientific literature. However, a study investigating the rotation mechanism of Crookes radiometers provides valuable insight.[24] It proposes that kinetic energy may arise from a process known as *transimpact*, which occurs as a side effect of atomic electron transitions.

Coulomb attraction typically arises between atoms or molecules nearby. However, when they come too close, electron-electron repulsion increases sharply. A balance between these opposing forces is achieved at a characteristic spacing known as the van der Waals distance. This equilibrium is often disrupted during atomic electron transitions. When an electron absorbs energy and becomes excited, it jumps to a higher orbital, typically within a few nanoseconds or less.[25-26] As the electron cloud expands, the volume of the host atom increases, thereby reducing the distance to adjacent atoms. This sudden decrease in spacing disturbs the existing force balance, usually resulting in a surge in repulsive forces that push the atoms apart. This phenomenon, illustrated in Figure 3, is referred to as transimpact.



A, Before orbital jump B, During orbital jump C, After orbital jump

Figure 3: Transimpact due to an atomic electron transition.

Transimpacts are explosive, impulsive events, similar to the sudden burst of popcorn, that forcefully push adjacent particles apart, imparting significant momentum and initiating or amplifying their motion. In Crookes radiometers, the imbalance in transimpacts between the black and white sides of the vanes drives their rotation. The transimpact theory predicts a maximum initial acceleration of this rotation, consistent with experimental observations and helping to resolve a longstanding mystery surrounding the radiometer's operation.[27–31] Atomic electron transitions, which occur routinely at the microscopic level, are the underlying source of transimpacts. As such, transimpacts are pervasive interactions that play a fundamental role in various physical processes, particularly within thermodynamics, such as thermal conduction, phase transitions,[32] and Brownian motion.[33–35] Ultimately, the kinetic energy of microscopic particles arises from radiative energy and reflects changes in potential energy.

While transimpacts enable the accumulation of kinetic energy within a system, this accumulation does not persist indefinitely. A counteracting mechanism must exist to convert kinetic energy into other forms—a process referred to as *impactrans*.[24] In impactrans, the motion or vibration of particles (such as molecules, atoms, or subatomic particles) leads to collisions via electrostatic interactions. These interactions can promote electrons to higher orbitals, or even eject them entirely, thereby altering the system's potential energy. Moreover, the accelerated motion of electrons during such impacts can emit radiation. Through impactrans, kinetic energy is thus transformed into both potential and radiative energy, maintaining the dynamic energy balance within the system.

The conversion of kinetic energy into radiative and potential energy through impactrans is evident in many everyday phenomena. For example, the sensation of warmth when rubbing one's palms together arises from frictional heating—an outcome of impactrans, where kinetic energy from surface motion is transformed into interactions between particles on coarse surfaces. A similar mechanism explains the heat generated at the base of a pump tube, where gas molecules collide more frequently and intensely, increasing the rate of impactrans. In the case of static electricity, rubbing a plastic rod against fur causes electrons to be dislodged from atoms, resulting in the accumulation of electric charge—another manifestation of impactrans.

Together, transimpact and impactrans facilitate the transformation of kinetic energy into potential and radiative energy. Heat is transferred through conduction, radiation, and convection. While radiation transfers energy and convection moves both energy and matter through space, the fundamental mechanism behind conduction remains less well understood. The concepts of transimpact and impactrans provide a compelling framework to explain this process, helping to bridge a longstanding gap in our understanding of conductive energy transfer.

Temperature and Dynamic Energy

Thermal energy, often referred to as heat energy, is traditionally defined as the total kinetic energy of a substance's particles, encompassing their translational, rotational, and vibrational motion. Temperature is typically understood as a measure of this thermal energy, and by extension, as an indicator of the average kinetic energy of the particles within a system. However, the preceding observations suggest that these conventional definitions may be inaccurate or fundamentally flawed, overlooking significant contributions from potential energy and radiative processes.

According to this study, potential energy, radiative energy, and kinetic energy are not isolated entities but continuously transform into one another. Radiative and potential energy are exchanged through atomic electron transitions.

Transimpacts convert radiative energy into kinetic energy, while impactranses transform kinetic energy into both potential and radiative energy. Together, these forms constitute the dynamic portion of a system's internal energy. Although all three influence temperature, it is most directly determined by radiative energy, which originates from potential energy. Kinetic energy contributes indirectly, as it is positively correlated with the other two forms. To encapsulate this interrelationship, the concept of *dynamic energy* is defined as the sum of these three interrelated components:

$$(10) \quad T = R + U + K$$

Here, *T* represents a system's dynamic energy, composed of radiative energy (*R*), potential energy (*U*), and kinetic energy (*K*). The continuous transformation among these forms drives the system's evolution. Framing them collectively as *dynamic energy* emphasizes their inherently interactive and ever-changing nature. This conceptual framework provides both a foundational perspective and an underlying mechanism for understanding thermodynamics.

Temperature measures a system's radiation level. As radiation is emitted by its potential energy components, temperature closely corresponds to the system's potential energy. Through processes such as transimpacts and impactranses, kinetic energy continuously transforms and exchanges with other forms of dynamic energy, thereby becoming indirectly linked to temperature. In this way, temperature serves as an indicator of a system's overall dynamic energy level.

The Misinterpreted Nature of Phase Transition

Traditional textbook explanations claim that the latent heat supplied during a phase transition is used to break intermolecular bonds, leading to no net increase in particle kinetic energy and, therefore, no change in temperature. However, this reasoning relies on the implicit assumption that temperature reflects the average kinetic energy of particles—a view that, as this study suggests, is inaccurate.

During a phase transition, such as the melting of ice into water, molecules are freed from their fixed positions within the hydrogen-bond network, leading to an increase in their kinetic energy. If temperature truly reflected the average kinetic energy of a system, one would expect it to rise during the transition. So why does the temperature remain unchanged?

In light of the new insights presented in this study, the apparent paradox becomes more comprehensible. As energy is introduced into a system, electrons are excited to higher orbitals, causing atoms to move farther apart and leading to increased bond lengths and potential energy. This bond stretching weakens intermolecular bonds while simultaneously intensifying transimpacts. Consequently, the temperature rises, reflecting the elevated level of radiation emitted by these increasingly strained bonds. Simultaneously, the system's kinetic energy increases, driven by the elevated transimpacts.

The temperature remains constant during a phase transition primarily because bond extension reaches its limit. In solids and liquids, as energy is added, intermolecular potential energy increases, leading to the stretching of molecular bonds. As bond lengths increase, their strength diminishes. Eventually, the bonds become too weak to withstand the rising intensity of transimpacts and begin to break, marking the onset of the phase transition. At this point, potential energy reaches its upper limit and can no longer increase. As a result, the radiation level stabilizes along with the emitting source. Since temperature reflects the system's radiation level, it remains constant throughout the transition. [30]

During a phase transition, latent heat is used both to break molecular bonds and to increase the kinetic energy of the liberated molecules. As a result, the kinetic energy of particles does indeed rise during phase transitions, contrary to the common assumption that it remains unchanged. However, these kinetic changes do not influence the system's radiative emission and thus are not reflected in temperature. Likewise, the act of breaking bonds does not alter emission levels; only changes in the potential energy components—namely, the bonds themselves—affect radiation. Since the system's overall emission is constrained by the maximum potential energy reached during the phase transition, the temperature, which reflects this radiative level, remains effectively fixed.

With this understanding, we can reexamine the relationship between temperature and a system's kinetic energy. It becomes clear that during phase transitions, when energy is added and kinetic energy increases while temperature remains constant, temperature becomes completely decoupled from kinetic energy. This highlights that temperature does not directly represent the kinetic energy of particle motion. However, outside of phase transitions, during equilibrium states, kinetic energy remains indirectly connected to temperature through ongoing energy exchange with potential and radiative forms.

Debunking the Thermal Expansion Myth

Classical textbooks often misleadingly attribute thermal expansion to increased molecular vibrations, suggesting that as the temperature rises, particles vibrate more vigorously and require additional space. If vibration alone were responsible for expansion, then a substance's volume should increase monotonically with temperature, since vibrational intensity generally grows with heat. However, this explanation fails to account for thermal contraction, such as water, which contracts as the temperature rises above 0 °C and reaches maximum density at 4 °C before expanding again. Such behavior directly contradicts the predictions of the vibration-based model and exposes fundamental flaws in its underlying assumptions.

Thermal expansion is fundamentally related to changes in the molecular bond structure. When a substance absorbs heat, electrons are excited to higher orbitals, causing atoms to move farther apart and increasing the system's potential energy. This results in an overall increase in volume. To accommodate the added energy, intermolecular bonds stretch further, contributing to the expansion of matter. The thermal expansion reflects changes in bond structure in two key ways: In crystalline solids, expansion can vary with direction, depending on the anisotropy of the crystal lattice. The volumetric thermal expansion coefficient tends to correlate with molecular complexity, as more intricate structures generally allow for greater bond flexibility and energy absorption.

Importantly, a substance's volume is determined not only by bond lengths but also by bond structures. This distinction helps explain unusual behaviors like thermal contraction in systems where bonding leads to more compact arrangements, such as the hydrogen bonding in water near 4 °C. Hydrogen bonds drive the intermolecular attraction between water molecules. As temperature decreases, these bonds become shorter and stronger, reducing volume and limiting molecular mobility. Around 4 °C, the hydrogen bonds become strong enough to form a scaffold-like network, causing water molecules to lock into specific positions. This creates hollow, less dense structures reminiscent of those in snowflakes. As the temperature drops further, this crystalline arrangement becomes more pronounced, resulting in volume expansion despite declining dynamic energy. Although conventional thermal expansion due to bond lengthening still occurs, its effect is outweighed by the structural changes driven by hydrogen bonding at lower temperatures.

At higher temperatures, the hydrogen bonds lengthen and weaken. Eventually, transimpacts disrupt these scaffolds, causing molecules to lose their fixed relative positions. Without this structural framework, water molecules pack more densely than in ice. As a result, above 4 °C, normal thermal expansion driven by changes in bond length becomes dominant, governing water's typical expansion behavior.

It is important to recognize that attractive forces between particles manifest not only as bonds between atoms or molecules but also as the attraction between electrons and their host nuclei. As a result, even substances composed of monatomic molecules exhibit thermal expansion, typically uniform in all directions due to their symmetric structure. In contrast, asymmetric thermal expansion or contraction is more common in crystals with complex or anisotropic bonding arrangements. These patterns of expansion cannot be explained by conventional vibration theory.

Conclusions

Understanding the roles of the three dynamic forms of energy, potential, kinetic, and radiative, is essential to resolving longstanding misconceptions in thermodynamics. These forms make up the dynamic component of a system's total energy and continually interconvert through atomic electron transitions, *transimpacts,* and *impactranses*, which underlie many thermodynamic phenomena, such as thermal conduction, thermal expansion, phase transition, Brownian motion, and the rotation of Crookes radiometers. Temperature reflects the level of a system's *dynamic energy* by directly measuring its radiative energy, which originates from potential energy sources. Kinetic energy, while not directly represented by temperature, is indirectly linked through energy exchanges facilitated by transimpacts and impactranses. Thermal expansion results from changes in bond structures in response to variations in a system's energy level, typically through bond stretching to accommodate additional energy. Phase transitions, on the other hand, occur at constant temperatures because of a fundamental limit on bond extension. Once this limit is reached, potential energy and its corresponding radiation level can no longer increase, causing the temperature to plateau. This explains why the kinetic energy of free particles may continue to change during a phase transition, even as the temperature remains fixed.

Revision History

- 09/15/2023: Initial Post on Stanford Site
- 10/22/2023: Removing Measurements of Specific Heat at Constant Pressure
- 11/01/2025: Published on Zenodo
- 12/17/2025: Adding Links to Summaries of Related Articles

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

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)
- 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)
- Is There a Sea of Free Electrons in Metals? (PDF: DOI) (中文: DOI)
- Electron Tunnel (PDF: DOI) (中文: DOI)
- Unified Theory of Low and High-Temperature Superconductivity (PDF: DOI) (中文: DOI)
- LK-99 Limitations and Significances (PDF: DOI) (中文: DOI)
- Superconductor Origin of Earth's Magnetic Field (PDF: DOI) (中文: DOI)
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- The Evolution from the Law of Gravitation to General Relativity (PDF: DOI) (中文: DOI)
- The Simplest Derivation of $E = mc^2$ (PDF: DOI) (中文: DOI)
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- AI Contamination (PDF) (中文)
- DeepSeek pk ChatGPT (PDF) (中文)

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