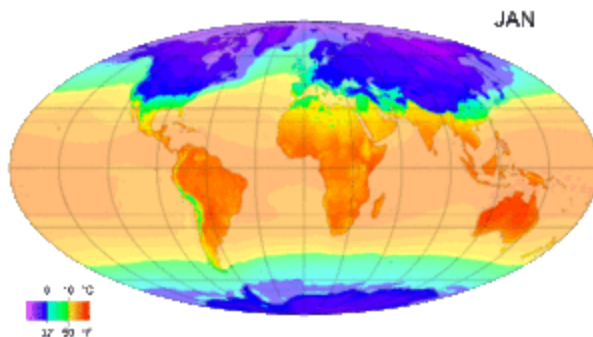# Quantum Mechanics_temperature

A map of global long term monthly average surface air temperatures in Mollweide projection.

A **temperature** is a numerical measure of hot and cold. Its measurement is by detection of heat radiation, particle velocity, kinetic energy, or most commonly, by the bulk behavior of a thermometric material. It may be calibrated in any of various temperature scales, Celsius, Fahrenheit, Kelvin, etc.

Measurements with a small thermometer, or by detection of heat radiation, can show that the temperature of a body of material can vary from time to time and from place to place within it. If changes happen too fast, or with too small a spacing, within a body, it may be impossible to define its temperature. Thus the concept of temperature in general has an empirical content.

Within a body that exchanges no energy or matter with its surroundings, temperature tends to become spatially uniform as time passes. When a path permeable only to heat is open between two bodies, energy always transfers spontaneously as heat from a hotter body to a colder one. The transfer rate depends on the nature of the path. If they are connected by a path permeable only to heat, and no heat flows between them, then the two bodies are equally hot. If changes are slow and spatially smooth enough to allow consistent comparisons of their hotness with other bodies that are respectively in their own states of internal thermodynamic equilibrium, they obey the Zeroth law of thermodynamics and then they have well defined and equal temperatures. Then thermodynamicsprovides a fundamental physical definition of temperature, on an absolute scale, relying on the second law of thermodynamics.

The kinetic theory offers a valuable but limited account of the behavior of the materials of macroscopic systems. It indicates the absolute temperature as proportional to the average kinetic energy of the random microscopic motions of their constituent microscopic particles such as electrons, atoms, and molecules.
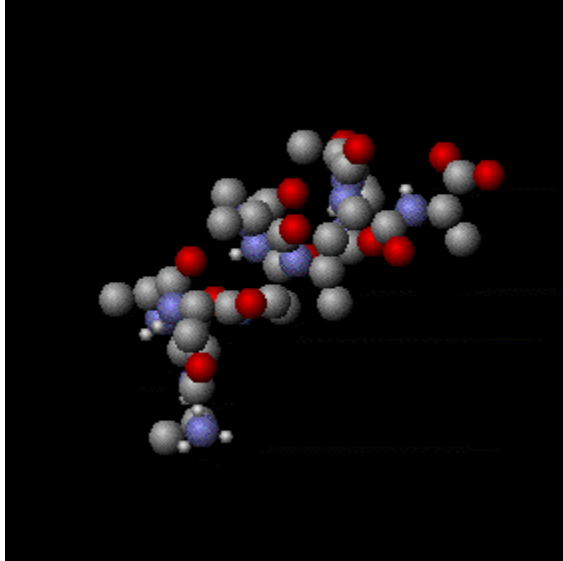

Thermal vibration of a segment of protein alpha helix. The amplitude of the vibrations increases with temperature.

The coldest theoretical temperature is called absolute zero. It can be approached but not reached in any actual physical system. It is denoted by 0 K on the Kelvin scale, −273.15 ℃ on the Celsius scale, and −459.67 ℉ on the Fahrenheit scale. In matter at absolute zero, the motions of microscopic constituents are minimal.

Temperature is important in all fields of natural science, including physics,geology, chemistry, atmospheric sciences and biology.

Use in science

This section **needs additional citations for verification**. Please help improve this article by adding citations to reliable sources. Unsourced material may be challenged and removed. *(January 2013)*

Annual Mean Temperature
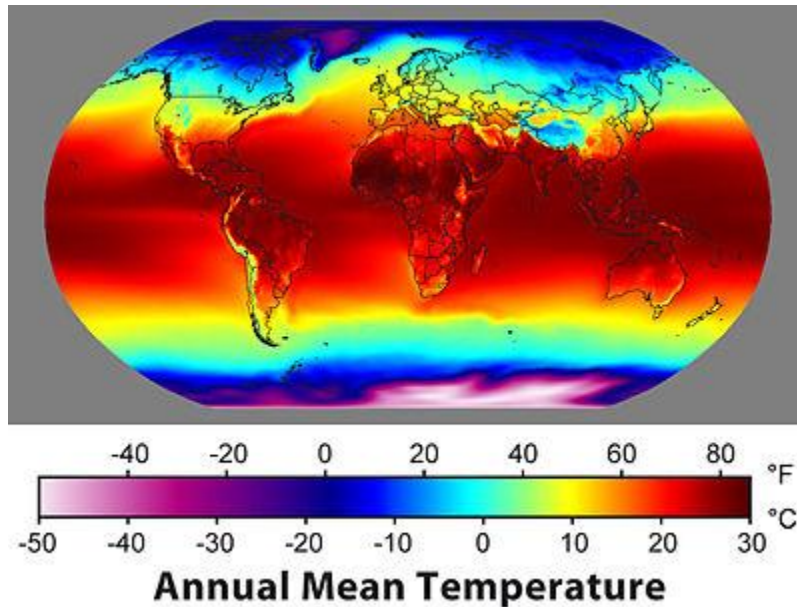
Annual mean temperature around the world

Many things depend on temperature, such as

- physical properties of materials including the phase (solid, liquid, gaseous orplasma), density, solubility, vapor pressure, electrical conductivity
- rate and extent to which chemical reactions occur
- the amount and properties of thermal radiation emitted from the surface of an object
- speed of sound is a function of the square root of the absolute temperature

Temperature scales

See also: Scale of temperature

This section **needs additional citations for verification**. Please help improve this article by adding citations to reliable sources. Unsourced material may be challenged and removed. *(August 2013)*

Temperature scales differ in two ways: the point chosen as zero degrees, and the magnitudes of incremental units or degrees on the scale.

The Celsius scale (℃) is used for common temperature measurements in most of the world. It is an empirical scale. It developed by a historical progress, which led to its zero point 0℃ being defined by the freezing point of water, with additional degrees defined so that 100℃ was the boiling point of water, both at sea-level atmospheric pressure. Because of the 100 degree interval, it is called a centigrade scale.[1] Since

the standardization of the kelvin in the International System of Units, it has subsequently been redefined in terms of the equivalent fixing points on the Kelvin scale, and so that a temperature increment of one degree celsius is the same as an increment of one kelvin, though they differ by an additive offset of 273.15.

The United States commonly uses the Fahrenheit scale, on which water freezes at 32 °F and boils at 212 °F at sea-level atmospheric pressure.

Many scientific measurements use the kelvin temperature scale (unit symbol K), named in honor of the Scottish physicist who first defined it. It is a thermodynamic or absolute temperature scale. Its zero point, 0K, is defined to coincide with coldest physically-possible temperature (called absolute zero). Its degrees are defined through thermodynamics. The temperature of absolute zero occurs at 0K = −273.15°C (or −459.67 °F), and the freezing point of water at sea-level atmospheric pressure occurs at 273.15K = 0°C.

The International System of Units (SI) defines a scale and unit for the kelvin or thermodynamic temperature by using the reliably reproducible temperature of the triple point of water as a second reference point (the first reference point being 0 K at absolute zero). The triple point is a singular state with its own unique and invariant temperature and pressure, along with, for a fixed mass of water in a vessel of fixed volume, an autonomically and stably self-determining partition into three mutually contacting phases, vapour, liquid, and solid, dynamically depending only on the total internal energy of the mass of water. For historical reasons, the triple point temperature of water is fixed at 273.16 units of the measurement increment.

Thermodynamic approach to temperature

Temperature is one of the principal quantities in the study of thermodynamics.

Kinds of temperature scale

There is a variety of kinds of temperature scale. It may be convenient to classify them as empirically and theoretically based. Empirical temperature scales are historically older, while theoretically based scales arose in the middle of the nineteenth century.[2][3]

Empirically based scales

Empirically based temperature scales rely directly on measurements of simple physical properties of materials. For example, the length of a column of mercury, confined in a glass-walled capillary tube, is dependent largely on temperature, and is the basis of the very useful mercury-in-glass thermometer. Such scales are valid only within

convenient ranges of temperature. For example, above the boiling point of mercury, a mercury-in-glass thermometer is impracticable. Most materials expand with temperature increase, but some materials, such as water, contract with temperature increase over some specific range, and then they are hardly useful as thermometric materials. A material is of no use as a thermometer near one of its phase-change temperatures, for example its boiling-point.

In spite of these restrictions, most generally used practical thermometers are of the empirically based kind. Especially, it was used for calorimetry, which contributed greatly to the discovery of thermodynamics. Nevertheless, empirical thermometry has serious drawbacks when judged as a basis for theoretical physics. Empirically based thermometers, beyond their base as simple direct measurements of ordinary physical properties of thermometric materials, can be re-calibrated, by use of theoretical physical reasoning, and this can extend their range of adequacy.

Theoretically based scales

Theoretically based temperature scales are based directly on theoretical arguments, especially those of thermodynamics, of kinetic theory, and of quantum mechanics. They rely on theoretical properties of idealized devices and materials. They are more or less comparable with practically feasible physical devices and materials. Theoretically based temperature scales are used to provide calibrating standards for practical empirically based thermometers.

The accepted fundamental thermodynamic temperature scale is the Kelvin scale, based on an ideal cyclic process envisaged for a Carnot heat engine.

An ideal material on which a temperature scale can be based is the ideal gas. The pressure exerted by a fixed volume and mass of an ideal gas is directly proportional to its temperature. Some natural gases show so nearly ideal properties over suitable temperature ranges that they can be used for thermometry; this was important during the development of thermodynamics, and is still of practical importance today.[4][5] The ideal gas thermometer is, however, not theoretically perfect for thermodynamics. This is because the entropy of an ideal gas at its absolute zero of temperature is not a positive semi-definite quantity, which puts the gas in violation of the third law of thermodynamics. The physical reason is that the ideal gas law, exactly read, refers to the limit of infinitely high temperature and zero pressure.[6][7][8]

Measurement of the spectrum of electromagnetic radiation from an ideal three-dimensional black body can provide an accurate temperature measurement because the frequency of maximum spectral radiance of black-body radiation is directly proportional to the temperature of the black body; this is known asWien's displacement law, and has a theoretical explanation in Planck's law and theBose-Einstein law.

Measurement of the spectrum of noise-power produced by an electrical resistor can also provide an accurate temperature measurement. The resistor has two terminals and is in effect a one-dimensional body. The Bose-Einstein law for this case indicates that the noise-power is directly proportional to the temperature of the resistor and to the value of its resistance and to the noise band-width. In a given frequency band, the noise-power has equal contributions from every frequency, and is called Johnson noise. If the value of the resistance is known then the temperature can be found.[9][10]

If molecules, or atoms, or electrons, are emitted from a material and their velocities are measured, the spectrum of their velocities often nearly obeys a theoretical law called the Maxwell–Boltzmann distribution, which gives a well-founded measurement of temperatures for which the law holds.[11] There have not yet been successful experiments of this same kind that directly use theFermi–Dirac distribution for thermometry, but perhaps that will be achieved in future.[12]

Absolute thermodynamic scale

The Kelvin scale is called **absolute** for two reasons. One is Kelvin's, that its formal character is independent of the properties of particular materials. The other reason is that its zero is in a sense absolute, in that it indicates absence of microscopic classical motion of the constituent particles of matter, so that they have a limiting specific heat of zero for zero temperature, according to the third law of thermodynamics. Nevertheless, a Kelvin temperature has a definite numerical value, that has been arbitrarily chosen by tradition. This numerical value also depends on the properties of water, which has a gas-liquid-solid triple point that can be reliably reproduced as a standard experimental phenomenon. The choice of this triple point is also arbitrary and by convention. The Kelvin scale is also called the **thermodynamic scale**.

Definition of the Kelvin scale

The thermodynamic definition of temperature is due to Kelvin.

It is framed in terms of an idealized device called a Carnot engine, imagined to define a continuous cycle of states of its working body. The cycle is imagined to run so slowly that at each point of the cycle the working body is in a state of thermodynamic equilibrium. There are four limbs in such a Carnot cycle. The engine consists of four bodies. The main one is called the working body. Two of them are called heat reservoirs, so large that their respective non-deformation variables are not changed by transfer of energy as heat through a wall permeable only to heat to the working body. The fourth body is able to exchange energy with the working body only through adiabatic work; it may be called the work reservoir. The substances and states of the two heat reservoirs should be chosen so that they are not in thermal equilibrium with one another. This means that they must be at different fixed temperatures, one, labeled here with the number 1, hotter than the other, labeled here with the number 2. This can be tested by connecting the heat reservoirs successively to an auxiliary empirical thermometric body that starts each time at a convenient fixed intermediate temperature. The thermometric body should be composed of a material that has a strictly monotonic relation between its chosen empirical thermometric variable and the amount of adiabatic isochoric work done on it. In order to settle the structure and sense of operation of the Carnot cycle, it is convenient to use such a material also for the working body; because most materials are of this kind, this is hardly a restriction of the generality of this definition. The Carnot cycle is considered to start from an initial condition of the working body that was reached by the completion of a reversible adiabatic compression. From there, the working body is initially connected by a wall permeable only to heat to the heat reservoir number 1, so that during the first limb of the cycle it expands and does work on the work reservoir. The second limb of the cycle sees the working body expand adiabatically and reversibly, with no energy exchanged as heat, but more energy being transferred as work to the work reservoir. The third limb of the cycle sees the working body connected, through a wall permeable only to heat, to the heat reservoir 2, contracting and accepting energy as work from the work reservoir. The cycle is closed by reversible adiabatic compression of the working body, with no energy transferred as heat, but energy being transferred to it as work from the work reservoir.

With this set-up, the four limbs of the reversible Carnot cycle are characterized by amounts of energy transferred, as work from the working body to the work reservoir,

and as heat from the heat reservoirs to the working body. The amounts of energy transferred as heat from the heat reservoirs are measured through the changes in the non-deformation variable of the working body, with reference to the previously known properties of that body, the amounts of work done on the work reservoir, and the first law of thermodynamics. The amounts of energy transferred as heat respectively from reservoir 1 and from reservoir 2 may then be denoted respectively Q_1 and Q_2. Then the absolute or thermodynamic temperatures, T_1 and T_2, of the reservoirs are defined so that to be such that

$$T_1/T_2 = -Q_1/Q_2 \ . \qquad\qquad (1)$$

Kelvin's original work postulating absolute temperature was published in 1848. It was based on the work of Carnot, before the formulation of the first law of thermodynamics. Kelvin wrote in his 1848 paper that his scale was absolute in the sense that it was defined "independently of the properties of any particular kind of matter." His definitive publication, which sets out the definition just stated, was printed in 1853, a paper read in 1851.[13][14][15][16]

This definition rests on the physical assumption that there are readily available walls permeable only to heat. In his detailed definition of a wall permeable only to heat, Carathéodory includes several ideas. The non-deformation state variable of a closed system is represented as a real number. A state of thermal equilibrium between two closed systems connected by a wall permeable only to heat means that a certain mathematical relation holds between the state variables, including the respective non-deformation variables, of those two systems (that particular mathematical relation is regarded by Buchdahl as a preferred statement of the zeroth law of thermodynamics).[17] Also, referring to thermal contact equilibrium, "whenever each of the systems S_1 and S_2 is made to reach equilibrium with a third system S_3 under identical conditions, the systems S_1 and S_2 are in mutual equilibrium."[18] It may viewed as a re-statement of the principle stated byMaxwell in the words: "All heat is of the same kind."[19] This physical idea is also expressed by Bailyn as a possible version of the zeroth law of thermodynamics: "All diathermal walls are equivalent."[20] Thus the present definition of thermodynamic temperature rests on the zeroth law of thermodynamics. Explicitly, this present definition of thermodynamic temperature also rests on the first law of thermodynamics, for the determination of amounts of energy transferred as heat.

Implicitly for this definition, the second law of thermodynamics provides information that establishes the virtuous character of the temperature so defined. It provides that any working substance that complies with the requirement stated in this definition will lead to the same ratio of thermodynamic temperatures, which in this sense is universal, or absolute. The second law of thermodynamics also provides that the thermodynamic temperature defined in this way is positive, because this definition requires that the heat reservoirs not be in thermal equilibrium with one another, and the cycle can be imagined to operate only in one sense if net work is to be supplied to the work reservoir.

Numerical details are settled by making one of the heat reservoirs a cell at the triple point of water, which is defined to have an absolute temperature of 273.16 K.[21] The zeroth law of thermodynamics allows this definition to be used to measure the absolute or thermodynamic temperature of an arbitrary body of interest, by making the other heat reservoir have the same temperature as the body of interest.

Temperature an intensive variable

In thermodynamic terms, temperature is an intensive variable because it is equal to a differential coefficient of one extensive variable with respect to another, for a given body. It thus has the dimensions of a ratio of two extensive variables. In thermodynamics, two bodies are often considered as connected by contact with a common wall, which has some specific permeability properties. Such specific permeability can be referred to a specific intensive variable. An example is a diathermic wall that is permeable only to heat; the intensive variable for this case is temperature. When the two bodies have been in contact for a very long time, and have settled to a permanent steady state, the relevant intensive variables are equal in the two bodies; for a diathermal wall, this statement is sometimes called the zeroth law of thermodynamics.[22][23][24]

In particular, when the body is described by stating its internal energy U, an extensive variable, as a function of its entropy S, also an extensive variable, and other state variables V, N, with $U = U(S, V, N)$, then the temperature is equal to the partial derivative of the internal energy with respect to the entropy:

$$T = \left(\frac{\partial U}{\partial S} \right)_{V,N}. \qquad (2)$$

[23][24][25]

Likewise, when the body is described by stating its entropy S as a function of its internal energy U, and other state variables V, N, with $S = S(U, V, N)$, then the reciprocal of the temperature is equal to the partial derivative of the entropy with respect to the internal energy:

$$\frac{1}{T} = \left(\frac{\partial S}{\partial U}\right)_{V,N}. \qquad (3)$$

[23][25][26]

The above definition, equation (1), of the absolute temperature is due to Kelvin. It refers to systems closed to transfer of matter, and has special emphasis on directly experimental procedures. A presentation of thermodynamics by Gibbs starts at a more abstract level and deals with systems open to the transfer of matter; in this development of thermodynamics, the equations (2) and (3) above are actually alternative definitions of temperature.[27]

Temperature local when local thermodynamic equilibrium prevails

Real world bodies are often not in thermodynamic equilibrium and not homogeneous. For study by methods of classical irreversible thermodynamics, a body is usually spatially and temporally divided conceptually into 'cells' of small size. If classical thermodynamic equilibrium conditions for matter are fulfilled to good approximation in such a 'cell', then it is homogeneous and a temperature exists for it. If this is so for every 'cell' of the body, then local thermodynamic equilibrium is said to prevail throughout the body.[28][29][30][31][32]

It makes good sense, for example, to say of the extensive variable U, or of the extensive variable S, that it has a density per unit volume, or a quantity per unit mass of the system, but it makes no sense to speak of density of temperature per unit volume or quantity of temperature per unit mass of the system. On the other hand, it makes no sense to speak of the internal energy at a point, while when local thermodynamic equilibrium prevails, it makes good sense to speak of the temperature at a point. Consequently, temperature can vary from point to point in a medium that is not in global thermodynamic equilibrium, but in which there is local thermodynamic equilibrium.

Thus, when local thermodynamic equilibrium prevails in a body, temperature can be regarded as a spatially varying local property in that body, and this is because temperature is an intensive variable.

Kinetic theory approach to temperature

A more thorough account of this is below at Theoretical foundation.

kinetic theory provides a microscopic explanation of temperature, based on macroscopic systems' being composed of many microscopic particles, such asmolecules and ions of various species, the particles of a species being all alike. It explains macroscopic phenomena through the classical mechanics of the microscopic particles. The equipartition theorem of kinetic theory asserts that each classical degree of freedom of a freely moving particle has an average kinetic energy of $k_B T/2$ where k_B denotes Boltzmann's constant. The translational motion of the particle has three degrees of freedom, so that, except at very low temperatures where quantum effects predominate, the average translational kinetic energy of a freely moving particle in a system with temperature T will be$3k_B T/2$.

It is possible to measure the average kinetic energy of constituent microscopic particles if they are allowed to escape from the bulk of the system. The spectrum of velocities has to be measured, and the average calculated from that. It is not necessarily the case that the particles that escape and are measured have the same velocity distribution as the particles that remain in the bulk of the system, but sometimes a good sample is possible.

molecules, such as oxygen (O_2), have more degrees of freedom than single spherical atoms: they undergo rotational and vibrational motions as well as translations. Heating results in an increase in temperature due to an increase in the average translational kinetic energy of the molecules. Heating will also cause, through equipartitioning, the energy associated with vibrational and rotational modes to increase. Thus a diatomic gas will require more energy input to increase its temperature by a certain amount, i.e. it will have a greater heat capacity than a monatomic gas.

The process of cooling involves removing internal energy from a system. When no more energy can be removed, the system is at absolute zero, though this cannot be achieved experimentally. Absolute zero is the null point of the thermodynamic temperature scale, also called absolute temperature. If it were possible to cool a system to absolute zero, all classical motion of its particles would cease and they would be at complete rest in this classical sense. Microscopically in the description of quantum mechanics, however, matter still has zero-point energyeven at absolute zero, because of the uncertainty principle.

Basic theory

Temperature is a measure of a quality of a state of a material [33] The quality may be regarded as a more abstract entity than any particular temperature scale that measures it, and is called *hotness* by some writers. The quality of hotness refers to the state of material only in a particular locality, and in general, apart from bodies held in a steady state of thermodynamic equilibrium, hotness varies from place to place. It is not necessarily the case that a material in a particular place is in a state that is steady and nearly homogeneous enough to allow it to have a well-defined hotness or temperature. Hotness may be represented abstractly as a one-dimensional manifold. Every valid temperature scale has its own one-to-one map into the hotness manifold.[34][35]

When two systems in thermal contact are at the same temperature no heat transfers between them. When a temperature difference does exist heat flows spontaneously from the warmer system to the colder system until they are in thermal equilibrium. Heat transfer occurs by conduction or by thermal radiation.[36][37][38][39][40][41][42][43]

Experimental physicists, for example Galileo and Newton,[44] found that there are indefinitely many empirical temperature scales. Nevertheless, the zeroth law of thermodynamics says that they all measure the same quality.

Temperature for bodies in thermodynamic equilibrium

For experimental physics, hotness means that, when comparing any two given bodies in their respective separate thermodynamic equilibria, any two suitably given empirical thermometers with numerical scale readings will agree as to which is the hotter of the two given bodies, or that they have the same temperature.[45] This does not require the two thermometers to have a linear relation between their numerical scale readings, but it does require that the relation between their numerical readings shall be strictly monotonic.[46][47] A definite sense of greater hotness can be had, independently of calorimetry, of thermodynamics, and of properties of particular materials, from Wien's displacement law of thermal radiation: the temperature of a bath of thermal radiation is proportional, by a universal constant, to the frequency of the maximum of its frequency spectrum; this frequency is always positive, but can have values that tend to zero. Thermal radiation is initially defined for a cavity in thermodynamic equilibrium. These physical facts justify a mathematical statement that hotness exists on an ordered one-dimensional manifold. This is a fundamental character of temperature and thermometers for bodies in their own thermodynamic equilibrium.[2][34][35][48][49]

Except for a system undergoing a first-order phase change such as the melting of ice, as a closed system receives heat, without change in its volume and without change in external force fields acting on it, its temperature rises. For a system undergoing such a phase change so slowly that departure from thermodynamic equilibrium can be neglected, its temperature remains constant as the system is supplied with latent heat. Conversely, a loss of heat from a closed system, without phase change, without change of volume, and without change in external force fields acting on it, decreases its temperature.[50]

Temperature for bodies in a steady state but not in thermodynamic equilibrium

While for bodies in their own thermodynamic equilibrium states, the notion of temperature requires that all empirical thermometers must agree as to which of two bodies is the hotter or that they are at the same temperature, this requirement is not safe for bodies that are in steady states though not in thermodynamic equilibrium. It can then well be that different empirical thermometers disagree about which is the hotter, and if this is so, then at least one of the bodies does not have a well defined absolute thermodynamic temperature. Nevertheless, any one given body and any one suitable empirical thermometer can still support notions of empirical, non-absolute, hotness and temperature, for a suitable range of processes. This is a matter for study in non-equilibrium thermodynamics.

Temperature for bodies not in a steady state

When a body is not in a steady state, then the notion of temperature becomes even less safe than for a body in a steady state not in thermodynamic equilibrium. This is also a matter for study in non-equilibrium thermodynamics.

Thermodynamic equilibrium axiomatics

For axiomatic treatment of thermodynamic equilibrium, since the 1930s, it has become customary to refer to a Zeroth law of thermodynamics. The customarily stated minimalist version of such a law postulates only that all bodies, which when thermally connected would be in thermal equilibrium, should be said to have the same temperature by definition, but by itself does not establish temperature as a quantity expressed as a real number on a scale. A more physically informative version of such a law views empirical temperature as a chart on a hotness manifold.[34][49][51] While the zeroth law permits the definitions of many different empirical scales of temperature, the second law of thermodynamics selects the definition of a single preferred, absolute temperature, unique up to an arbitrary scale factor, whence called

thethermodynamic temperature.[2][34][52][53][54][55] If internal energy is considered as a function of the volume and entropy of a homogeneous system in thermodynamic equilibrium, thermodynamic absolute temperature appears as the partial derivative of internal energy with respect the entropy at constant volume. Its natural, intrinsic origin or null point is absolute zero at which the entropy of any system is at a minimum. Although this is the lowest absolute temperature described by the model, the third law of thermodynamics postulates that absolute zero cannot be attained by any physical system.

Heat capacity

See also: heat capacity and calorimetry

When a sample is heated, meaning it receives thermal energy from an external source, some of the introduced heat is converted into kinetic energy, the rest to other forms of internal energy, specific to the material. The amount converted into kinetic energy causes the temperature of the material to rise. The introduced heat (ΔQ) divided by the observed temperature change is the heat capacity (C) of the material.

$$C = \frac{\Delta Q}{\Delta T}$$

If heat capacity is measured for a well defined amount of substance, the specific heat is the measure of the heat required to increase the temperature of such a unit quantity by one unit of temperature. For example, to raise the temperature of water by one kelvin (equal to one degree Celsius) requires 4186 joules per kilogram(J/kg)..

Temperature measurement

A typical Celsius thermometer measures a winter day temperature of –17℃.

See also: Timeline of temperature and pressure measurement technology, International Temperature Scale of 1990 and Comparison of temperature scales

Temperature measurement using modern scientific thermometers and temperature scales goes back at least as far as the early 18th century, when Gabriel Fahrenheit adapted a thermometer (switching to mercury) and a scale both developed by Ole Christensen Rømer. Fahrenheit's scale is still in use in the United States for non-scientific applications.

Temperature is measured with thermometers that may be calibrated to a variety of temperature scales. In most of the world (except for Belize, Myanmar, Liberiaand the United States), the Celsius scale is used for most temperature measuring purposes. Most scientists measure temperature using the Celsius scale and thermodynamic temperature using the Kelvin scale, which is the Celsius scale offset so that its null point is 0K = −273.15℃, or absolute zero. Many engineering fields in the U.S., notably high-tech and US federal specifications (civil and military), also use the Kelvin and Celsius scales. Other engineering fields in the U.S. also rely upon the Rankine scale (a shifted Fahrenheit scale) when working in thermodynamic-related disciplines such as combustion.

Units

The basic unit of temperature in the <u>International System of Units</u> (SI) is the<u>Kelvin</u>. It has the symbol K.

For everyday applications, it is often convenient to use the Celsius scale, in which0℃ corresponds very closely to the <u>freezing point</u> of water and 100℃ is its<u>boiling point</u> at sea level. Because liquid droplets commonly exist in clouds at sub-zero temperatures, 0℃ is better defined as the melting point of ice. In this scale a temperature difference of 1 degree Celsius is the same as a 1kelvin increment, but the scale is offset by the temperature at which ice melts (273.15 K).

By international agreement[56] the Kelvin and Celsius scales are defined by two fixing points: <u>absolute zero</u> and the <u>triple point</u> of <u>Vienna Standard Mean Ocean Water</u>, which is water specially prepared with a specified blend of hydrogen and oxygen isotopes. Absolute zero is defined as precisely 0K and −273.15℃. It is the temperature at which all classical translational motion of the particles comprising matter ceases and they are at complete rest in the classical model. Quantum-mechanically, however, zero-point motion remains and has an associated energy, the <u>zero-point energy</u>. Matter is in its <u>ground state</u>,[57] and contains no <u>thermal energy</u>. The triple point of water is defined as 273.16K and 0.01℃. This definition serves the following purposes: it fixes the magnitude of the kelvin as being precisely 1 part in 273.16 parts of the difference between absolute zero and the triple point of water; it establishes that one kelvin has precisely the same magnitude as one degree on the Celsius scale; and it establishes the difference between the null points of these scales as being 273.15K (0K = −273.15℃ and273.16K = 0.01℃).

In the United States, the Fahrenheit scale is widely used. On this scale the freezing point of water corresponds to 32 ℉ and the boiling point to 212 ℉. The Rankine scale, still used in fields of chemical engineering in the U.S., is an absolute scale based on the Fahrenheit increment.

Conversion

The following table shows the <u>temperature conversion formulas</u> for conversions to and from the Celsius scale.

	from <u>Celsius</u>	to Celsius
Fahrenheit	[℉] = [℃] × ⁹⁄₅ + 32	[℃] = ([℉] − 32) × ⁵⁄₉
Kelvin	[K] = [℃] + 273.15	[℃] = [K] − 273.15

Rankine $[°R] = ([°C] + 273.15) \times \frac{9}{5}$ $[°C] = ([°R] − 491.67) \times \frac{5}{9}$

Delisle $[°De] = (100 − [°C]) \times \frac{3}{2}$ $[°C] = 100 − [°De] \times \frac{2}{3}$

Newton $[°N] = [°C] \times \frac{33}{100}$ $[°C] = [°N] \times \frac{100}{33}$

Réaumur $[°Ré] = [°C] \times \frac{4}{5}$ $[°C] = [°Ré] \times \frac{5}{4}$

Rømer $[°Rø] = [°C] \times \frac{21}{40} + 7.5$ $[°C] = ([°Rø] − 7.5) \times \frac{40}{21}$

Plasma physics

The field of plasma physics deals with phenomena of electromagnetic nature that involve very high temperatures. It is customary to express temperature in electronvolts (eV) or kiloelectronvolts (keV), where 1 eV = 11605K. In the study of QCD matter one routinely encounters temperatures of the order of a few hundred MeV, equivalent to about 10^{12}K.

Theoretical foundation

See also: thermodynamic temperature

Historically, there are several scientific approaches to the explanation of temperature: the classical thermodynamic description based on macroscopic empirical variables that can be measured in a laboratory; the kinetic theory of gases which relates the macroscopic description to the probability distribution of the energy of motion of gas particles; and a microscopic explanation based on statistical physics and quantum mechanics. In addition, rigorous and purely mathematical treatments have provided an axiomatic approach to classical thermodynamics and temperature.[58] Statistical physics provides a deeper understanding by describing the atomic behavior of matter, and derives macroscopic properties from statistical averages of microscopic states, including both classical and quantum states. In the fundamental physical description, using natural units, temperature may be measured directly in units of energy. However, in the practical systems of measurement for science, technology, and commerce, such as the modern metric system of units, the macroscopic and the microscopic descriptions are interrelated by the Boltzmann constant, a proportionality factor that scales temperature to the microscopic mean kinetic energy.

The microscopic description in statistical mechanics is based on a model that analyzes a system into its fundamental particles of matter or into a set of classical or quantum-mechanical oscillators and considers the system as a statistical ensemble of microstates. As a collection of classical material particles, temperature is a measure of the mean energy of motion, called kinetic energy, of the particles, whether

in solids, liquids, gases, or plasmas. The kinetic energy, a concept of <u>classical mechanics</u>, is half the <u>mass</u> of a particle times its <u>speed</u>squared. In this mechanical interpretation of thermal motion, the kinetic energies of material particles may reside in the velocity of the particles of their translational or vibrational motion or in the inertia of their rotational modes. In monatomic <u>perfect gases</u> and, approximately, in most gases, temperature is a measure of the mean particle kinetic energy. It also determines the probability distribution function of the energy. In condensed matter, and particularly in solids, this purely mechanical description is often less useful and the oscillator model provides a better description to account for quantum mechanical phenomena. Temperature determines the statistical occupation of the microstates of the ensemble. The microscopic definition of temperature is only meaningful in the <u>thermodynamic limit</u>, meaning for large ensembles of states or particles, to fulfill the requirements of the statistical model.

In the context of thermodynamics, the kinetic energy is also referred to as<u>thermal energy</u>. The thermal energy may be partitioned into independent components attributed to the <u>degrees of freedom</u> of the particles or to the modes of oscillators in a <u>thermodynamic system</u>. In general, the number of these degrees of freedom that are available for the <u>equipartitioning</u> of energy depend on the temperature, i.e. the energy region of the interactions under consideration. For solids, the thermal energy is associated primarily with the <u>vibrations</u> of its atoms or molecules about their equilibrium position. In an <u>ideal monatomic gas</u>, the kinetic energy is found exclusively in the purely translational motions of the particles. In other systems, <u>vibrational</u> and <u>rotational</u> motions also contribute degrees of freedom.

Kinetic theory of gases
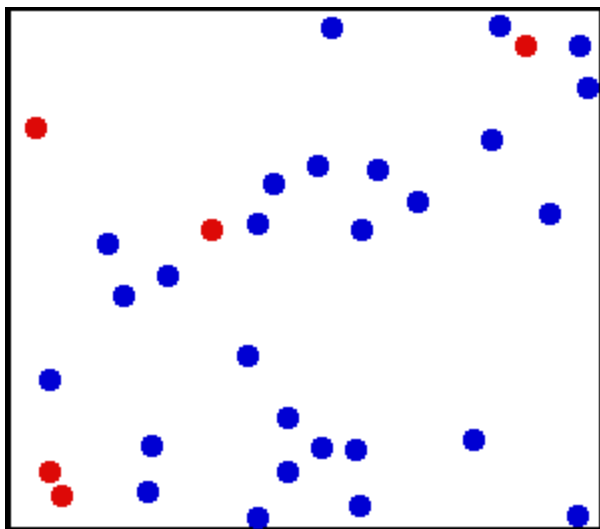
A theoretical understanding of temperature in an ideal gas can be obtained from the underlined kinetic theory.

Maxwell and Boltzmann developed a kinetic theory that yields a fundamental understanding of temperature in gases.[59] This theory also explains the ideal gaslaw and the observed heat capacity of monatomic (or 'noble') gases.[60][61][62]
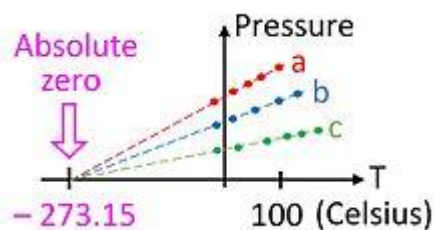
Plots of pressure vs temperature for three different gas samples extrapolate to absolute zero.

The ideal gas law is based on observed empirical relationships between pressure (p), volume (V), and temperature (T), and was recognized long before the kinetic theory of gases was developed (see Boyle's and Charles's laws). The ideal gas law states:[63]

$$pV = nRT$$

where n is the number of moles of gas and R = 8.314472(15) J mol^{-1}K^{-1} is the gas constant.

This relationship gives us our first hint that there is an absolute zero on the temperature scale, because it only holds if the temperature is measured on anabsolute scale such as Kelvins. The ideal gas law allows one to measure temperature on this absolute scale using the gas thermometer. The temperature in kelvins can be

defined as the pressure in pascals of one mole of gas in a container of one cubic meter, divided by the gas constant.

Although it is not a particularly convenient device, the gas thermometer provides an essential theoretical basis by which all thermometers can be calibrated. As a practical matter it is not possible to use a gas thermometer to measure absolute zero temperature since the gases tend to condense into a liquid long before the temperature reaches zero. It is possible, however, to extrapolate to absolute zero by using the ideal gas law, as shown in the figure.

The kinetic theory assumes that pressure is caused by the force associated with individual atoms striking the walls, and that all energy is translational kinetic energy. Using a sophisticated symmetry argument,[64] Boltzmann deduced what is now called the Maxwell–Boltzmann probability distribution function for the velocity of particles in an ideal gas. From that probability distribution function, the average kinetic energy, E_k (per particle), of a monatomic ideal gas is:[61][65]

$$E_\mathrm{k} = \frac{1}{2}mv^2_\mathrm{rms} = \frac{3}{2}kT \,,$$

where the Boltzmann constant, k, is the Avogadro number divided by the ideal gas constant, and v_{rms} is the root–mean–square speed. Thus the ideal gas law states that internal energy is directly proportional to temperature.[66] This direct proportionality between temperature and internal energy is a special case of theequipartition theorem, and holds only in the classical limit of an ideal gas. It does not hold for most substances, although it is true that temperature is a monotonic(non-decreasing) function of internal energy.

Zeroth law of thermodynamics

Main article: Zeroth law of thermodynamics

When two otherwise isolated bodies are connected together by a rigid physical path impermeable to matter, there is spontaneous transfer of energy as heat from the hotter to the colder of them. Eventually they reach a state of mutual thermal equilibrium, in which heat transfer has ceased, and the bodies' respective state variables have settled to become unchanging.

One statement of the Zeroth law of thermodynamics is that if two systems are each in thermal equilibrium with a third system, then they are also in thermal equilibrium with each other.

This statement helps to define temperature but it does not, by itself, complete the definition. An empirical temperature is a numerical scale for the hotness of a thermodynamic system. Such hotness may be defined as existing on a one-dimensional manifold, stretching between hot and cold. Sometimes the zeroth law is stated to include the existence of a unique universal hotness manifold, and of numerical scales on it, so as to provide a complete definition of empirical temperature.[51] To be suitable for empirical thermometry, a material must have a monontonic relation between hotness and some easily measured state variable, such as pressure or volume, when all other relevant coordinates are fixed. An exceptionally suitable system is the ideal gas, which can provide a temperature scale that matches the absolute Kelvin scale. The Kelvin scale is defined on the basis of the second law of thermodynamics.

Second law of thermodynamics

Main article: second law of thermodynamics

In the previous section certain properties of temperature were expressed by the zeroth law of thermodynamics. It is also possible to define temperature in terms of the second law of thermodynamics which deals with entropy. Entropy is often thought of as a measure of the disorder in a system. The second law states that any process will result in either no change or a net increase in the entropy of the universe. This can be understood in terms of probability.

For example, in a series of coin tosses, a perfectly ordered system would be one in which either every toss comes up heads or every toss comes up tails. This means that for a perfectly ordered set of coin tosses, there is only one set of toss outcomes possible: the set in which 100% of tosses come up the same. On the other hand, there are multiple combinations that can result in disordered or mixed systems, where some fraction are heads and the rest tails. A disordered system can be 90% heads and 10% tails, or it could be 98% heads and 2% tails, et cetera. As the number of coin tosses increases, the number of possible combinations corresponding to imperfectly ordered systems increases. For a very large number of coin tosses, the combinations to ~50% heads and ~50% tails dominates and obtaining an outcome significantly different from 50/50 becomes extremely unlikely. Thus the system naturally progresses to a state of maximum disorder or entropy.

It has been previously stated that temperature governs the transfer of heat between two systems and it was just shown that the universe tends to progress so as to maximize entropy, which is expected of any natural system. Thus, it is expected that there is some relationship between temperature and entropy. To find this relationship, the relationship between heat, work and temperature is first considered. A heat engine is a device for converting thermal energy into mechanical energy, resulting in the performance of work, and analysis of the Carnot heat engine provides the necessary relationships. The work from a heat engine corresponds to the difference between the heat put into the system at the high temperature, q_H and the heat ejected at the low temperature, q_C. The efficiency is the work divided by the heat put into the system or:

$$\text{efficiency} = \frac{w_{cy}}{q_H} = \frac{q_H - q_C}{q_H} = 1 - \frac{q_C}{q_H} \qquad (4)$$

where w_{cy} is the work done per cycle. The efficiency depends only on q_C/q_H. Because q_C and q_H correspond to heat transfer at the temperatures T_C and T_H, respectively, q_C/q_H should be some function of these temperatures:

$$\frac{q_C}{q_H} = f(T_H, T_C) \qquad (5)$$

Carnot's theorem states that all reversible engines operating between the same heat reservoirs are equally efficient. Thus, a heat engine operating between T_1 and T_3 must have the same efficiency as one consisting of two cycles, one between T_1 and T_2, and the second between T_2 and T_3. This can only be the case if:

$$q_{13} = \frac{q_1 q_2}{q_2 q_3}$$

which implies:

$$q_{13} = f(T_1, T_3) = f(T_1, T_2) f(T_2, T_3)$$

Since the first function is independent of T_2, this temperature must cancel on the right side, meaning $f(T_1, T_3)$ is of the form $g(T_1)/g(T_3)$ (i.e. $f(T_1, T_3)$ = $f(T_1, T_2)f(T_2, T_3)$ = $g(T_1)/g(T_2) \cdot g(T_2)/g(T_3) = g(T_1)/g(T_3)$), where g is a function of a single temperature. A temperature scale can now be chosen with the property that:

$$\frac{q_C}{q_H} = \frac{T_C}{T_H} \qquad (6)$$

Substituting Equation 4 back into Equation 2 gives a relationship for the efficiency in terms of temperature:

$$\text{efficiency} = 1 - \frac{q_C}{q_H} = 1 - \frac{T_C}{T_H} \qquad (7)$$

Notice that for $T_C = 0$ K the efficiency is 100% and that efficiency becomes greater than 100% below 0 K. Since an efficiency greater than 100% violates the first law of thermodynamics, this implies that 0 K is the minimum possible temperature. In fact the lowest temperature ever obtained in a macroscopic system was 20 nK, which was achieved in 1995 at NIST. Subtracting the right hand side of Equation 5 from the middle portion and rearranging gives:

$$\frac{q_H}{T_H} - \frac{q_C}{T_C} = 0$$

where the negative sign indicates heat ejected from the system. This relationship suggests the existence of a state function, S, defined by:

$$dS = \frac{dq_{\mathrm{rev}}}{T} \tag{8}$$

where the subscript indicates a reversible process. The change of this state function around any cycle is zero, as is necessary for any state function. This function corresponds to the entropy of the system, which was described previously. Rearranging Equation 6 gives a new definition for temperature in terms of entropy and heat:

$$T = \frac{dq_{\mathrm{rev}}}{dS} \tag{9}$$

For a system, where entropy S(E) is a function of its energy E, the temperature T is given by:

$$T^{-1} = \frac{d}{dE} S(E) \tag{10}$$
,

i.e. the reciprocal of the temperature is the rate of increase of entropy with respect to energy.

Definition from statistical mechanics

statistical mechanics defines temperature based on a system's fundamental degrees of freedom. Eq.(10) is the defining relation of temperature. Eq. (9) can be derived from the principles underlying the fundamental thermodynamic relation.

Generalized temperature from single particle statistics

It is possible to extend the definition of temperature even to systems of few particles, like in a quantum dot. The generalized temperature is obtained by considering time ensembles instead of configuration space ensembles given in statistical mechanics in the case of thermal and particle exchange between a small system of fermions (N even less than 10) with a single/double occupancy system. The finite quantum grand

canonical   ensemble,[67] obtained   under   the   hypothesis   of ergodicity and orthodicity,[68] allows to express the generalized temperature from the ratio of the average time of occupation τ_1 and τ_2 of the single/double occupancy system:[69]

$$T = k^{-1} \ln 2 \frac{\tau_2}{\tau_1} \left(E - E_F \left(1 + \frac{3}{2N} \right) \right),$$

where E_F is the Fermi energy which tends to the ordinary temperature when N goes to infinity.

Negative temperature

Main article: Negative temperature

On the empirical temperature scales, which are not referenced to absolute zero, a negative temperature is one below the zero-point of the scale used. For example,dry ice has a sublimation temperature of $-78.5°C$ which is equivalent to $-109.3°F$. On the absolute Kelvin scale, however, this temperature is 194.6 K. On the absolute scale of thermodynamic temperature no material can exhibit a temperature smaller than or equal to 0 K, both of which are forbidden by the third law of thermodynamics.

In the quantum mechanical description of electron and nuclear spin systems that have a limited number of possible states, and therefore a discrete upper limit of energy they can attain, it is possible to obtain a Negative temperature, which is numerically indeed less than absolute zero. However, this is not the macroscopic temperature of the material, but instead the temperature of only very specific degrees of freedom, that are isolated from others and do not exchange energy by virtue of the equipartition theorem.

A negative temperature is experimentally achieved with suitable radio frequency techniques that cause a population inversion of spin states from the ground state. As the energy in the system increases upon population of the upper states, the entropy increases as well, as the system becomes less ordered, but attains a maximum value when the spins are evenly distributed among ground and excited states, after which it begins to decrease, once again achieving a state of higher order as the upper states begin to fill exclusively. At the point of maximum entropy, the temperature function shows the behavior of a singularity, because the slope of the entropy function decreases to zero at first and then turns negative. Since temperature is the inverse of the derivative of the entropy, the temperature formally goes to infinity at this point, and switches to negative infinity as the slope turns negative. At energies higher than this point, the spin degree of freedom therefore exhibits formally a negative

thermodynamic temperature. As the energy increases further by continued population of the excited state, the negative temperature approaches zero asymptotically.[70] As the energy of the system increases in the population inversion, a system with a negative temperature is not colder than absolute zero, but rather it has a higher energy than at positive temperature, and may be said to be in fact hotter at negative temperatures. When brought into contact with a system at a positive temperature, energy will be transferred from the negative temperature regime to the positive temperature region.

Examples of temperature

Main article: Orders of magnitude (temperature)

	Temperature		Peak emittancewavelength[71]
	Kelvin	Degrees Celsius	of black–body radiation
absolute zero (precisely by definition)	0 K	−273.15 ℃	cannot be defined
Coldest temperature achieved[72]	100 pK	−273.149999999900 ℃	29,000 km
Coldest Bose-Einstein condensate[73]	450 pK	−273.14999999955 ℃	6,400 km
One millikelvin (precisely by definition)	0.001 K	−273.149 ℃	2.89777 m (radio, FM band)[74]
Water's triple point (precisely by definition)	273.16 K	0.01 ℃	10,608.3 nm (long wavelength I.R.)
Water's boiling point[A]	373.1339 K	99.9839 ℃	7,766.03 nm (mid wavelength I.R.)
Incandescent lamp[B]	2500 K	≈2,200 ℃	1,160 nm (near infrared)[C]
Sun's visible surface[D][75]	5,778 K	5,505 ℃	501.5 nm (green–blue light)
Lightning bolt's channel[E]	28 kK	28,000 ℃	100 nm (far ultraviolet light)
Sun's core[E]	16 MK	16 million ℃	0.18 nm (X–rays)

Thermonuclear weapon (peak temperature)[E][76]	350 MK	350 million ℃	8.3×10^{-3} nm (gamma rays)
Sandia National Labs' Z machine[E][77]	2 GK	2 billion ℃	1.4×10^{-3} nm (gamma rays)[F]
Core of a high-mass star on its last day[E][78]	3 GK	3 billion ℃	1×10^{-3} nm (gamma rays)
Merging binary neutron star system[E][79]	350 GK	350 billion ℃	8×10^{-6} nm (gamma rays)
Relativistic Heavy Ion Collider[E][80]	1 TK	1 trillion ℃	3×10^{-6} nm (gamma rays)
CERN's proton vs nucleus collisions[E][81]	10 TK	10 trillion ℃	3×10^{-7} nm (gamma rays)
Universe 5.391×10^{-44} s after the Big Bang[E]	1.417×10^{32} K	1.417×10^{32} ℃	1.616×10^{-27} nm (Planck Length)[82]

- [A] For Vienna Standard Mean Ocean Water at one standard atmosphere (101.325 kPa) when calibrated strictly per the two-point definition of thermodynamic temperature.
- [B] The 2500 K value is approximate. The 273.15 K difference between K and ℃ is rounded to 300 K to avoid false precision in the Celsius value.
- [C] For a true black-body (which tungsten filaments are not). Tungsten filaments' emissivity is greater at shorter wavelengths, which makes them appear whiter.
- [D] Effective photosphere temperature. The 273.15 K difference between K and ℃ is rounded to 273 K to avoid false precision in the Celsius value.
- [E] The 273.15 K difference between K and ℃ is without the precision of these values.
- [F] For a true black-body (which the plasma was not). The Z machine's dominant emission originated from 40 MK electrons (soft x-ray emissions) within the plasma.