

Thermodynamics: Heat and Temperature

I. Introduction to Thermodynamics

This topic explores thermodynamics, the study of heat and its relationship to other forms of energy and work. It begins with an intuitive understanding of hot and cold, moving towards precise scientific definitions and laws.

II. Thermal Equilibrium

Definition

A system is in thermal equilibrium when its macroscopic properties (such as temperature, pressure, and volume) remain constant over time. This state is achieved when the system is isolated, preventing any net heat transfer with its surroundings.

Examples

A cup of hot water, if sufficiently isolated, will eventually reach a stable temperature and stay there indefinitely, signifying thermal equilibrium.

When hot and cold water are mixed, the system initially lacks a uniform temperature and is therefore not in equilibrium. Over time, the mixture will attain a single, uniform temperature, establishing a new equilibrium state.

Similarly, a gas within a piston under a fixed load, if undisturbed, reaches equilibrium. If weights are abruptly removed, the piston moves, the gas experiences turbulence, and pressure varies. The system is out of equilibrium until it settles into a new stable position with uniform conditions.

Macroscopic vs. Microscopic Properties

During transitions, macroscopic properties like temperature may not be uniformly defined across the system. However, at the microscopic level, individual atoms and molecules always possess defined states (e.g., specific locations and velocities).

Temperature and thermal equilibrium are macroscopic concepts, only applicable when the system has reached a stable, settled state.

III. Temperature and the Zeroth Law of Thermodynamics

Intuitive Basis

Temperature is fundamentally an intuitive measure of how "hot" or "cold" something

feels.

The Zeroth Law of Thermodynamics

Statement: If two systems (A and B) are each in thermal equilibrium with a third system (C), then systems A and B are also in thermal equilibrium with each other.

Significance: This law is foundational for defining and measuring temperature universally. It allows a thermometer (system C) to accurately compare the temperatures of two objects (A and B) that may never directly interact. It establishes temperature as a transitive property.

IV. Quantitative Measurement of Temperature (Thermometry)

Need for a Scale

To quantify temperature, a standardized numerical scale and units are essential, much like those for length or mass.

Principle of Thermometry

Thermometers rely on physical properties that change predictably with temperature. Examples include:

Thermal Expansion of Solids

The change in length (ΔL) of a solid rod is directly proportional to its original length (L_0) and the change in temperature (ΔT):

$$\Delta L = \alpha L_0 \Delta T$$

Here, α (alpha) is the coefficient of linear expansion, a characteristic property of the material.

Thermal Expansion of Liquids

The change in volume (ΔV) of a liquid is proportional to its initial volume (V_0) and the change in temperature (ΔT):

$$\Delta V = \beta V_0 \Delta T$$

Here, β (beta) is the coefficient of volume expansion, also specific to the material.

Liquid-in-Glass Thermometers

These thermometers use the volume expansion of a liquid (e.g., mercury or alcohol) contained in a thin capillary tube. The narrowness of the tube amplifies small volume changes into easily visible changes in the liquid column's height.

Temperature Scales

Standard temperature scales require well-defined and reproducible reference points.

Celsius Scale (Centigrade)

Reference Points:

Freezing point of water: Defined as 0°C .

Boiling point of water: Defined as 100°C .

Interval: The range between 0°C and 100°C is divided into 100 equal degrees.

Limitations:

The boiling point of water is sensitive to atmospheric pressure (e.g., it boils at a lower temperature at higher altitudes), making it a less precise universal standard.

Different liquids expand non-linearly at different rates. While two liquid thermometers can be calibrated to agree at 0°C and 100°C , they may give slightly different readings at intermediate temperatures. The chosen liquid effectively defines the scale.

V. The Gas Thermometer and Absolute Zero

Superiority of Gas Thermometers

Experiments reveal that for *dilute* gases, the product of pressure (P) and volume (V) exhibits a remarkably linear relationship with temperature.

Significantly, different dilute gases, when calibrated at two points (e.g., 0°C and 100°C), show excellent agreement across the entire temperature range. This inherent linearity makes the gas thermometer a highly universal and reliable standard.

Absolute Zero

When the PV vs. T graph for *any* dilute gas is extrapolated, it consistently indicates that the product PV approaches zero at a common temperature of approximately -273.15°C .

This unique temperature is termed absolute zero. It represents the theoretical lowest possible temperature at which the pressure of an ideal gas would become zero.

The Kelvin Scale (Absolute Temperature Scale)

Definition: The Kelvin scale sets its zero point at absolute zero (0 K).

Reference Point: The triple point of water is defined as 273.16 K. The triple point is the unique temperature and pressure at which water, ice, and steam coexist in stable

equilibrium. This provides a single, highly reproducible reference point for the scale.

Interval: A change of 1 Kelvin (1 K) is precisely equal to a change of 1 degree Celsius (1°C).

Conversion: To convert from Celsius to Kelvin: $T(K) = T(^{\circ}C) + 273.15$.

Notation: Temperatures on the Kelvin scale are referred to simply as "Kelvin" (e.g., 300 K), not "degrees Kelvin."

VI. Heat Transfer (Q)

Definition

Heat (Q) is the transfer of thermal energy between systems or regions within a system due to a temperature difference.

Historical Context (Caloric Theory)

Historically, heat was often conceived as an invisible, self-repelling fluid called "caloric" that flowed from hotter to colder bodies.

Units of Heat

Calorie (cal): Defined as the amount of heat energy required to raise the temperature of 1 gram of water by 1 degree Celsius (specifically, from 14.5°C to 15.5°C).

Kilocalorie (kcal or Calorie with capital C for nutrition): 1 kcal = 1000 cal.

Specific Heat Capacity (c)

Definition: The specific heat capacity (c) of a substance is the amount of heat energy required to raise the temperature of a unit mass of that substance by one degree.

Formula: $\Delta Q = mc\Delta T$

ΔQ : heat transferred (in calories or Joules)

m: mass of the substance (in grams or kilograms)

c: specific heat capacity (in cal/g°C or J/kgK)

ΔT : change in temperature (in °C or K)

Material-Dependent Property: The value of 'c' varies for different materials. For water, c is approximately 1 cal/g°C (or 4186 J/kgK).

Note: Specific heat capacities are not strictly constant; they can vary slightly with temperature and the phase of the substance.

Calorimetry

Principle: In an isolated system, the total heat energy lost by hotter components is equal to the total heat energy gained by colder components. Therefore, the net change in heat for the entire system is zero: $\Sigma(\Delta Q) = 0$.

Problem Solving: When objects at different initial temperatures are brought into thermal contact, they exchange heat until they reach a common final temperature. This principle is used to determine unknown specific heats or final temperatures.

Example: If hot lead pellets are immersed in cold water, the heat lost by the lead is equal to the heat gained by the water. The equation would be:

$$(m_{\text{lead}})(c_{\text{lead}})(T_{\text{final}} - T_{\text{initial_lead}}) + (m_{\text{water}})(c_{\text{water}})(T_{\text{final}} - T_{\text{initial_water}}) = 0.$$

VII. Phase Changes and Latent Heat

Phase Changes

During a phase transition (e.g., solid to liquid, liquid to gas), a substance's temperature remains constant despite continuous absorption or release of heat. This energy is used to break or form intermolecular bonds and rearrange the molecular structure, rather than increasing the kinetic energy (temperature) of the molecules.

Latent Heat (L)

Definition: Latent heat (L) is the amount of heat energy absorbed or released per unit mass of a substance during a phase change, occurring at a constant temperature.

Formula: $\Delta Q = mL$

ΔQ : heat transferred (in calories or Joules)

m: mass of the substance undergoing the phase change (in grams or kilograms)

L: latent heat (in cal/g or J/kg)

Types of Latent Heat:

Latent Heat of Fusion (L_f): The energy involved in melting (solid to liquid) or freezing (liquid to solid). For water, $L_f \approx 80 \text{ cal/g}$.

Latent Heat of Vaporization (L_v): The energy involved in boiling (liquid to gas) or condensation (gas to liquid). For water, $L_v \approx 540 \text{ cal/g}$.

Note: Specific heats of a substance differ between its various phases (e.g., specific heat of ice is different from that of liquid water).

Solving Problems with Phase Changes

These problems typically involve multiple energy transfer steps:

1. Heating or cooling a substance within a single phase (using $\Delta Q = mc\Delta T$).
2. Undergoing a phase change (using $\Delta Q = mL$).

Strategy: When dealing with mixtures involving phase changes, calculate the heat required for each possible stage. It's often helpful to assume a final state (e.g., all solid, all liquid, or a mixture of both) and then verify if the calculated energy transfer is consistent with that assumption. If not, another final state should be considered.

VIII. Modes of Heat Transfer

Heat energy can be transferred through three primary mechanisms:

Radiation

Mechanism: Transfer of energy via electromagnetic waves.

Requirement: Does not require a medium; it can occur efficiently through a vacuum.

Examples: Heat received from the Sun, warmth felt from a glowing hot object like a space heater's coils.

Convection

Mechanism: Transfer of heat through the actual bulk movement of fluids (liquids or gases).

Process: When a fluid is heated, it becomes less dense and rises. Cooler, denser fluid then sinks to take its place, creating a circulating current (convection current) that distributes heat.

Examples: Boiling water in a pot, atmospheric weather patterns, heat distribution in central heating systems.

Conduction

Mechanism: Transfer of heat through direct contact between particles (atoms or molecules) without any bulk movement of the material. Energy is passed from more energetic particles to less energetic ones through collisions and vibrations.

Example: Heat flowing from a hot stove burner through a metal pan handle to your hand.

Rate of Heat Conduction (Fourier's Law):

The rate of heat flow ($\Delta Q/\Delta t$), also known as heat current (H), through a

material is given by:

$$H = \Delta Q / \Delta t = -\kappa A (\Delta T / \Delta x)$$

κ (kappa): This is the thermal conductivity, a property specific to the material (high for good conductors like metals, low for insulators like wood).

A: The cross-sectional area through which heat flows.

$\Delta T / \Delta x$: The temperature gradient, representing the change in temperature over the distance. The negative sign indicates that heat flows from a region of higher temperature to a region of lower temperature.

IX. Heat as a Form of Energy (Mechanical Equivalent of Heat)

Historical Context

Early scientific observations noted that mechanical work (e.g., from friction or compression) could produce heat, and vice-versa. This provided evidence against the caloric fluid theory.

Joule's Experiment

James Prescott Joule conducted experiments that definitively showed a quantitative relationship between mechanical work and heat produced.

Experiment: In a classic experiment, falling weights rotated a paddle wheel immersed in an insulated container of water. The mechanical potential energy lost by the falling weights was precisely converted into an increase in the water's temperature.

Conclusion: Heat is not a separate substance but rather a form of energy.

Mechanical Equivalent of Heat: Joule's experiments established that 1 calorie is equivalent to approximately 4.186 Joules (often rounded to 4.2 J/cal). This constant provides the conversion factor between mechanical energy (Joules) and thermal energy (calories).

Connection to Conservation of Energy

The recognition of heat as a form of energy led to the formulation of the First Law of Thermodynamics, which is a fundamental statement of the conservation of energy. This law includes heat and internal energy in the overall energy balance.

In processes like inelastic collisions, the "lost" macroscopic kinetic energy is not truly lost but is transformed into the internal energy (heat) of the colliding objects. This internal energy manifests as the increased random kinetic energy of the atoms and

molecules within the material. Thus, energy is conserved across all its forms.