

THERMODYNAMICS: HEAT AND TEMPERATURE

1. Introduction to Thermodynamics

- * Thermodynamics is a fundamental branch of physics concerned with heat and its relation to other forms of energy and work.
- * It shifts focus from the mechanics of individual particles to the macroscopic properties of systems, particularly those involving energy transfer as heat.
- * Key concepts include temperature, heat, energy, and entropy.

2. Thermal Equilibrium

- * **Definition:** A system is in thermal equilibrium when its macroscopic properties, such as temperature, pressure, and volume, cease to change over time, provided the system is isolated from its surroundings.
- * **Macroscopic vs. Microscopic:**
- * At the macroscopic level, a system in equilibrium appears static, with constant measurable properties.
- * At the microscopic level, constituent atoms and molecules are always in motion, possessing well-defined positions and velocities. Thermal equilibrium refers to the average, observable behavior, not the instantaneous state of every particle.
- * **Examples:**
- * An isolated cup of hot water or cold water, given sufficient time, will reach a state of thermal equilibrium, maintaining a constant temperature.
- * When hot and cold water are mixed, the system is initially not in equilibrium, as different parts have different temperatures. After a period of transition and mixing, the mixture will reach a new, uniform temperature, signifying equilibrium.
- * A gas in a piston with constant weights is in equilibrium. If weights are suddenly removed, the piston moves, and the gas becomes turbulent (not in equilibrium). After settling, it reaches a new equilibrium state at a different volume and pressure.

3. Defining Temperature

- * **Intuitive Notion:** Temperature is intuitively understood as the degree of hotness or coldness of an object.
- * **Need for Precision:** Physics requires a quantitative and universally reproducible

definition of temperature.

- * **The Zeroth Law of Thermodynamics:**

- * **Statement:** If two systems (A and B) are each in thermal equilibrium with a third system (C), then they are in thermal equilibrium with each other (A and B are at the same temperature).

- * **Significance:** This law establishes temperature as a transitive property, allowing for the consistent measurement and comparison of temperatures between systems that are not in direct contact. It underpins the use of thermometers, where the thermometer (C) establishes a common temperature for two other systems (A and B).

4. Quantitative Temperature Measurement and Scales

- * **Early Attempts: Thermal Expansion:**

- * Many materials expand predictably when heated. This property can be used to define a temperature scale.

- * **Linear Expansion:** The change in length (ΔL) of a rod is proportional to its original length (L_0) and the change in temperature (ΔT): $\Delta L = \alpha L_0 \Delta T$, where α is the coefficient of linear expansion (material-dependent).

- * **Volume Expansion:** The change in volume (ΔV) of a substance is proportional to its original volume (V_0) and the change in temperature (ΔT): $\Delta V = \beta V_0 \Delta T$, where β is the coefficient of volume expansion (material-dependent).

- * **Liquid-in-Glass Thermometers:** These utilize the volume expansion of a liquid (e.g., mercury, alcohol) in a thin capillary tube to magnify the change in volume into a visible change in height.

- * **Standardization of Temperature Scales:**

- * To create a reproducible scale, two fixed and easily achievable reference points are required, with the interval between them divided into equal steps.

- * **Celsius (Centigrade) Scale:**

- * Freezing/Melting Point of Water (at standard atmospheric pressure): Defined as 0 °C.

- * Boiling Point of Water (at standard atmospheric pressure): Defined as 100 °C.

- * The interval between these points is divided into 100 equal degrees.

- * **Limitations of Liquid Thermometers:** Different liquids expand at slightly different non-linear rates, meaning two thermometers calibrated at 0 °C and 100 °C might not agree on intermediate temperatures (e.g., 50 °C). This necessitates choosing a standard substance.

- * **Limitations of Water's Fixed Points:** The boiling point of water varies with atmospheric pressure (e.g., lower at high altitudes), making it an unreliable universal standard without specifying pressure.

5. The Gas Thermometer and Absolute Zero

- * **Principle:** Dilute gases exhibit a nearly linear relationship between the product of pressure (P) and volume (V) and temperature. Importantly, for very dilute gases, this relationship is universal, meaning all dilute gases behave similarly.
- * **Universality:** When calibrated at 0 °C and 100 °C, dilute gas thermometers agree on intermediate temperatures, overcoming the limitation of liquid thermometers.
- * **Absolute Zero:** Extrapolating the PV vs. Temperature relationship for a dilute gas reveals that PV approaches zero at a specific temperature, approximately -273.15 °C. This point, where the pressure (and thus molecular motion) theoretically ceases, is defined as the absolute zero of temperature.
- * **Significance of Absolute Zero:**
 - * It is the theoretical lowest possible temperature, a universal and fundamental constant of nature, independent of any specific substance like water.
 - * It implies a minimum energy state for matter, where molecular motion is minimized (though quantum mechanics indicates residual zero-point energy).
- * **Kelvin (Absolute) Temperature Scale:**
 - * **Definition:** The Kelvin scale is an absolute thermodynamic temperature scale where 0 K corresponds to absolute zero.
 - * **Reference Point:** The triple point of water (the unique temperature and pressure at which water, ice, and water vapor coexist in equilibrium) is defined as 273.16 K. This allows for a single, highly reproducible reference point for calibration.
 - * **Relationship to Celsius:** $T(K) = T(^{\circ}C) + 273.15$. A 1 K change is equivalent to a 1 °C change.
 - * **Nomenclature:** Temperatures on the Kelvin scale are referred to as "Kelvin" (e.g., 273 K), not "degrees Kelvin."

6. Heat (Q)

- * **Definition:** Heat is the transfer of thermal energy between systems due to a temperature difference. It is energy in transit.
- * **Historical Concept (Caloric Theory):** Early theories proposed heat as an invisible

fluid called "caloric" that flowed from hotter to colder bodies.

- * **The Calorie:**

- * **Definition:** Historically, one calorie (cal) was defined as the amount of heat required to raise the temperature of 1 gram of water by 1 degree Celsius (from 14.5 °C to 15.5 °C).

- * **Kilocalorie (kcal or Calorie):** 1 kilocalorie = 1000 calories, often used in nutrition.

- * **Specific Heat (c):**

- * **Definition:** The specific heat of a substance is the amount of heat energy required to raise the temperature of a unit mass of that substance by one degree Celsius (or Kelvin). It is a material-dependent property.

- * **Formula:** $Q = mc\Delta T$, where Q is the heat transferred, m is the mass, c is the specific heat, and ΔT is the change in temperature.

- * **Water's Specific Heat:** Approximately 1 cal/(g·°C) or 4184 J/(kg·K).

- * **Limitations:** Specific heat is not truly constant; it can vary with temperature and pressure, especially over wide ranges.

- * **Calorimetry: Conservation of Heat:**

- * **Principle:** In an isolated system where heat is exchanged between substances, the total heat lost by hotter bodies equals the total heat gained by colder bodies. Mathematically, the sum of all heat changes ($\sum Q$) in an isolated system is zero.

- * **Application:** Used to determine the specific heat of unknown materials or to calculate final equilibrium temperatures in mixtures.

7. Phase Changes

- * **Definition:** A phase change is a transition of a substance from one state of matter (e.g., solid, liquid, gas) to another. These changes occur at specific temperatures and pressures.

- * **Latent Heat (L):**

- * **Definition:** The energy absorbed or released during a phase change without a change in temperature. This energy is used to rearrange the molecular structure rather than increase kinetic energy.

- * **Formula:** $Q = mL$, where m is the mass of the substance undergoing a phase change and L is the latent heat for that specific transition.

- * **Latent Heat of Fusion (L_f):** Heat required to change a substance from solid to liquid (or vice versa). For water, $L_f \approx 80$ cal/g.

- * **Latent Heat of Vaporization (L_v):** Heat required to change a substance from liquid to gas (or vice versa). For water, $L_v \approx 540 \text{ cal/g}$.
- * **Temperature Plateaus:** During a phase change, even as heat is continuously added or removed, the temperature of the substance remains constant until the entire sample has completed the transition.
- * **Complex Calorimetry Problems:** Problems involving phase changes require considering the heat absorbed/released during temperature changes ($mc\Delta T$) and during phase changes (mL). The final state of the mixture (e.g., all liquid, all solid, or a mix of phases) depends on the total energy balance.

8. Mechanisms of Heat Transfer

* **1. Radiation:**

- * **Mechanism:** Transfer of thermal energy via electromagnetic waves (e.g., infrared radiation).
- * **Medium Requirement:** Does not require a medium; can occur through a vacuum.
- * **Example:** Heat from the sun reaching Earth, heat felt from a glowing space heater.

* **2. Convection:**

- * **Mechanism:** Transfer of heat through the bulk movement of fluids (liquids or gases). Hotter, less dense fluid rises, and colder, denser fluid sinks, creating convection currents.
- * **Medium Requirement:** Requires a fluid medium.
- * **Example:** Boiling water in a pot, atmospheric weather patterns (warm air rising).

* **3. Conduction:**

- * **Mechanism:** Transfer of heat through direct contact between particles, without macroscopic movement of the material itself. Energy is transferred via collisions between atoms or molecules.
- * **Medium Requirement:** Requires direct physical contact between materials.
- * **Example:** Heat traveling along a metal spoon placed in hot soup, heat transfer through a pan on a stove.
- * **Rate of Heat Conduction (Fourier's Law):**
- * The rate of heat flow ($Q/\Delta t$) through a material is proportional to the cross-sectional area (A), the temperature difference (ΔT) across the material, and inversely proportional to the thickness/length (Δx) through which the heat travels.

- * $Q/\Delta t = \kappa A(\Delta T/\Delta x)$, where κ (kappa) is the thermal conductivity of the material.
- * **Thermal Conductivity (κ):** A material-dependent property that quantifies how well a substance conducts heat. Metals typically have high thermal conductivity, while insulators (like wood or air) have low thermal conductivity.
- * **Reservoir:** A large body whose temperature remains constant despite heat transfer into or out of it, often idealized for practical problems (e.g., the ocean, a large room).

9. Heat as a Form of Energy: The First Law of Thermodynamics (Introduction)

- * **Observation:** In many physical processes, mechanical energy (kinetic, potential) appears to be "lost," yet the system's temperature increases (e.g., inelastic collisions, drilling, friction).
- * **Joule's Experiment:**
- * **Setup:** A falling weight drives a paddle wheel that churns water in an insulated container.
- * **Observation:** The mechanical potential energy of the falling weight is converted into internal energy (heat) of the water, causing its temperature to rise.
- * **Conclusion:** There is a direct proportionality between the amount of mechanical energy "lost" (measured in Joules) and the amount of heat "gained" (measured in calories).
- * **Mechanical Equivalent of Heat:** Approximately 1 calorie = 4.184 Joules (often rounded to 4.2 J/cal).
- * **Implication:** Heat is not a separate substance but another form of energy. The principle of conservation of energy can be extended to include heat.
- * **Microscopic View:** Heat represents the disordered, random kinetic energy of the atoms and molecules within a substance. Macroscopic kinetic energy (e.g., a moving car) is the ordered motion of the entire system, while heat is the random internal motion. When macroscopic motion is lost (e.g., in a collision), it is transformed into this random molecular motion, causing a temperature increase.