

THERMODYNAMICS: FUNDAMENTALS OF HEAT AND TEMPERATURE

I. INTRODUCTION AND COURSE OVERVIEW

A. Course Schedule Updates

- * New problem set assigned today, due next Wednesday.
- * One final problem set with 2-3 problems will be assigned based on upcoming topics.

B. New Topic: Thermodynamics

- * Focus for the next four lectures: Heat, Temperature, and related concepts.
- * Builds upon intuitive understanding, transitioning to precise physical definitions.

II. DEFINING TEMPERATURE

A. Intuitive Understanding

- * Common notion of "hot" and "cold."
- * This intuition serves as a starting point for more precise definitions.

B. Precision in Physics

- * The concepts of hot and cold require rigorous definition for scientific measurement.
- * This leads to the crucial concept of thermodynamic equilibrium.

III. THERMAL EQUILIBRIUM

A. Definition

- * A state where the macroscopic properties of a system cease to change over time.
- * Similar to mechanical equilibrium, but concerning thermal properties like temperature.

B. Concrete Examples

- * Isolated cup of hot or cold water: If isolated from its surroundings, it will settle to a constant temperature, indicating thermal equilibrium.
- * **Mixing hot and cold water:**

- * Initially, the system is not in equilibrium (temperature is ill-defined across the mixture).
- * After sufficient time, the mixture reaches a uniform, well-defined temperature, signifying a new state of equilibrium.
- * **Gas in a piston:**
- * Initially at equilibrium (piston stationary, pressure uniform).
- * Removing weights causes disequilibrium (piston movement, turbulent gas, non-uniform pressure).
- * Eventually settles to a new equilibrium state with uniform pressure and temperature.

C. Macroscopic vs. Microscopic

- * Macroscopic properties (like temperature) are only well-defined when a system is in equilibrium.
- * Microscopically, atoms and molecules always have defined states (position, velocity).
- * Temperature represents a gross, macroscopic property, not a property of individual molecules.

IV. ZEROth LAW OF THERMODYNAMICS

A. Statement

- * "If two systems are each in thermal equilibrium with a third system, then they are in thermal equilibrium with each other."
- * Symbolically: If A is in equilibrium with B, and B is in equilibrium with C, then A is in equilibrium with C.

B. Significance

- * This seemingly obvious law is fundamental to the concept of temperature.
- * It allows for the global comparison of temperature using a "third system" (e.g., a thermometer).
- * Without this law, a universal temperature scale, where an attribute called temperature can be compared across systems that never directly interact, would not be possible.

V. QUANTITATIVE MEASUREMENT OF TEMPERATURE

A. Need for Quantification

- * Just as "tall" and "short" become precise heights, "hot" and "cold" need quantitative measures.
- * Relies on observable physical properties that change consistently with temperature.

B. Thermal Expansion

- * Principle: Most materials expand when heated and contract when cooled.
- * Example: A metal rod expands when heated.
- * Linear Expansion: The change in length (ΔL) of a material is proportional to its original length (L_0) and the change in temperature (ΔT).
- * $\Delta L = \alpha L_0 \Delta T$
- * α (alpha) is the coefficient of linear expansion, a material-specific property.
- * Volume Expansion: Similarly, the change in volume (ΔV) is proportional to the original volume (V_0) and ΔT .
- * $\Delta V = \beta V_0 \Delta T$
- * β (beta) is the coefficient of volume expansion, also material-specific (approx. 3α for isotropic solids).

C. Liquid-in-Glass Thermometers

- * Mechanism: Utilizes the volume expansion of a liquid (e.g., mercury, alcohol) in a narrow, evacuated tube.
- * Magnification: A small volume expansion of the liquid in the bulb leads to a significant rise in height in the narrow capillary tube ($\Delta V = \text{Area} \times \Delta x$), allowing for precise readings.

D. Establishing a Temperature Scale (Arbitrary Markings)

- * Initial idea: Mark arbitrary points on the thermometer (e.g., 0, 5, 19) as long as it's monotonic (temperature increases with height).
- * Problem: Such a scale is not universal or reproducible across different instruments or locations.

VI. TEMPERATURE SCALES

A. Centigrade (Celsius) Scale

- * **Standardization points:**

- * Freezing/Melting point of water (ice-water mixture at standard atmospheric pressure): Defined as 0 degrees Celsius (0°C).
- * Boiling/Condensation point of water (water-steam mixture at standard atmospheric pressure): Defined as 100 degrees Celsius (100°C).
- * Intervals: The range between 0°C and 100°C is divided into 100 equal "degrees."
- * Reproducibility: These points are chosen because they are generally reproducible worldwide under specified conditions (e.g., sea level pressure).

B. Challenges with Liquid Thermometers and Boiling Point

- * Dependence on external factors: The boiling point of water varies with atmospheric pressure (e.g., lower boiling point at higher altitudes like Denver). This makes it an unreliable universal standard for high precision.
- * Non-linearity of expansion: Different liquids expand at slightly different rates over temperature ranges.
- * Example: A mercury thermometer and an alcohol thermometer, calibrated at 0°C and 100°C , may not agree at intermediate temperatures (e.g., one reads 75°C , the other 72°C).
- * This necessitated choosing a "standard" liquid, which was ultimately arbitrary.

C. Gas Thermometer (The Preferred Solution)

- * Principle: The product of pressure (P) and volume (V) of a *dilute* gas shows a linear relationship with temperature.
- * Construction: A gas sample in a container with a movable piston (to measure P and V).
- * Advantage: When calibrated at 0°C and 100°C , *all* dilute gases exhibit this linear relationship and agree on intermediate temperatures.
- * The more dilute the gas, the closer it behaves to an "ideal gas," minimizing intermolecular interactions that could cause non-linearity.
- * Reliability: This consistency across different gases makes the gas thermometer a much more reliable and universal standard for temperature measurement.

VII. ABSOLUTE ZERO AND KELVIN SCALE

A. Discovery of Absolute Zero

- * Extrapolation of Gas Thermometer Data: If the linear P-V vs. T relationship of a dilute

gas is extrapolated below 0°C , the product $P \cdot V$ appears to vanish at a specific temperature.

- * Universal Point: Remarkably, all dilute gases, regardless of their type or amount, extrapolate to the *same* temperature where $P \cdot V$ theoretically becomes zero.
- * Value: This point is approximately -273.15°C (or -273.16°C as mentioned in the lecture).
- * Significance: This indicates a fundamental lower limit to temperature, beyond which pressure cannot be further reduced (as it cannot be negative).

B. The Kelvin Scale (Absolute Temperature Scale)

- * Definition: The zero point of the Kelvin scale (0 K) is set at absolute zero (-273.15°C).
- * Unit Size: One Kelvin (K) is defined to be the same magnitude as one degree Celsius ($^{\circ}\text{C}$).
- * Second Reference Point: To define the straight line, one other universally reproducible point is needed.
- * Triple Point of Water: The unique temperature and pressure at which water, ice, and steam can coexist in stable equilibrium. This is a very specific condition, making it highly reproducible.
- * Value: The triple point of water is defined as 273.16 K (which is 0.01°C).
- * Conversion: Temperature in Kelvin (T_{K}) = Temperature in Celsius (T_{C}) + 273.15.
- * Nomenclature: It is formally "Kelvin" (e.g., 100 K), not "degrees Kelvin" (unlike "degrees Celsius" or "degrees Fahrenheit").
- * Importance: Kelvin is the absolute temperature scale used in scientific contexts, as its zero point is tied to a fundamental physical property, not an arbitrary substance like water.

VIII. INTRODUCTION TO HEAT (Q)

A. Intuitive Understanding of Heat Transfer

- * When a hot object is brought into contact with a cold object, the cold object gets hotter, and the hot object gets cooler.
- * This transfer of "something" from hot to cold was historically called "heat."

B. Historical "Caloric Fluid" Theory

- * Early scientists imagined heat as an invisible, self-repellent fluid (caloric) that flowed

from hotter to colder bodies.

- * This theory explained many observations but failed to account for heat generated by friction or mechanical work.

C. Defining the Calorie (Unit of Heat)

- * A unit of heat based on water's properties.
- * Definition: One calorie (cal) is the amount of heat required to raise the temperature of one gram of water by one degree Celsius (specifically from 14.5°C to 15.5°C).
- * Kilocalorie (kcal): 1 kcal = 1000 cal; often used for nutritional energy ("Calories").

IX. SPECIFIC HEAT AND LATENT HEAT

A. Specific Heat (c)

- * Definition: The amount of heat required to raise the temperature of one unit mass of a substance by one degree Celsius (or Kelvin).
- * Formula: $\Delta Q = m * c * \Delta T$
- * ΔQ : Heat transferred (calories or joules)
- * m: Mass of the substance (grams or kilograms)
- * c: Specific heat capacity (material-specific constant)
- * ΔT : Change in temperature (final T - initial T)
- * Water: By definition, the specific heat of water (c_{water}) is 1 cal/(g·°C) or 1 kcal/(kg·°C).
- * Other Materials: Specific heat varies greatly between different substances (e.g., copper, wood, gold). These values are determined experimentally.
- * Analogy: Similar to thermal expansion coefficients (α , β), specific heat is a material property that quantifies how much energy is needed to change its temperature.
- * Note: Specific heat is not truly constant but varies slightly with temperature, especially over large ranges. For typical problems, it's treated as constant.

B. Calorimetry Problems (Heat Exchange)

- * Principle: In an isolated system, the total heat exchanged is zero; heat lost by one part of the system is gained by another.
- * $\Sigma(m * c * \Delta T) = 0$
- * Example: Mixing a hot material (lead pellets) with colder water.
- * Heat lost by lead = Heat gained by water.

$$* (m_{\text{Pb}} * c_{\text{Pb}} * (T_{\text{final}} - T_{\text{Pb,initial}})) + (m_{\text{water}} * c_{\text{water}} * (T_{\text{final}} - T_{\text{water,initial}})) = 0$$

* This allows for determination of unknown specific heats or final temperatures.

C. Phase Changes and Latent Heat

* Phenomenon: During a phase change (e.g., melting, freezing, boiling, condensation), a substance absorbs or releases heat **without** a change in temperature.

* Example: Heating ice from -30°C to steam at >100°C.

1. Heating ice: T rises from -30°C to 0°C. ($\Delta Q = m_{\text{ice}} * c_{\text{ice}} * \Delta T$)

2. Melting ice: At 0°C, ice converts to water. Temperature remains constant. ($\Delta Q = m_{\text{melted}} * L_{\text{f}}$)

3. Heating water: T rises from 0°C to 100°C. ($\Delta Q = m_{\text{water}} * c_{\text{water}} * \Delta T$)

4. Vaporizing water: At 100°C, water converts to steam. Temperature remains constant. ($\Delta Q = m_{\text{vaporized}} * L_{\text{v}}$)

5. Heating steam: T rises above 100°C. ($\Delta Q = m_{\text{steam}} * c_{\text{steam}} * \Delta T$)

* Latent Heat (L): The amount of heat required to change the phase of one unit mass of a substance at constant temperature.

* L_{f} (Latent Heat of Fusion): For melting/freezing (e.g., water: 80 cal/g).

* L_{v} (Latent Heat of Vaporization): For boiling/condensation (e.g., water: ~540 cal/g).

* Complex Problems: When mixing substances that undergo phase changes (e.g., ice at -40°C with water at +40°C), one must account for specific heat changes and latent heats.

* Often involves making an initial assumption (e.g., all ice melts, or all water freezes) and verifying the final temperature (e.g., if calculated T is below 0°C but ice was assumed to melt, the assumption is wrong).

* If the final state is a mixture of phases (e.g., ice and water at 0°C), the problem becomes determining the final amounts of each phase.

X. MODES OF HEAT TRANSFER

A. Radiation

* Mechanism: Transfer of heat via electromagnetic waves (e.g., infrared radiation).

* Requirement: No medium is required; can occur through a vacuum.

* Examples: Heat from the sun, warmth from a glowing space heater.

B. Convection

- * Mechanism: Transfer of heat through the movement of heated fluid (liquid or gas).
- * Process: When a fluid is heated, it expands, becomes less dense, and rises. Cooler, denser fluid then sinks, creating a convection current.
- * Examples: Boiling water in a pot, atmospheric thermal currents, forced-air heating systems.

C. Conduction

- * Mechanism: Transfer of heat through direct contact between particles, without macroscopic movement of the material itself.
- * Process: Vibrating particles in a hotter region transfer energy to adjacent, less energetic particles.
- * Example: Heat moving from a hot stove burner through a metal skillet to its handle.
- * Heat Conduction Formula: The rate of heat flow ($\Delta Q/\Delta t$) through a material.
- * $\Delta Q/\Delta t = -\kappa * A * (\Delta T/\Delta x)$
- * $\Delta Q/\Delta t$: Rate of heat transfer (energy per unit time, e.g., Joules/second or Watts).
- * κ (kappa): Thermal conductivity, a material-specific property (high for metals, low for insulators).
- * A: Cross-sectional area of the material.
- * $\Delta T/\Delta x$: Temperature gradient (change in temperature over a distance, often called temperature difference over length).
- * The negative sign indicates heat flows from higher temperature to lower temperature.
- * Comparison: Allows for direct comparison of how well different materials conduct heat, normalized by their geometry (area, length) and temperature difference.

XI. NATURE OF HEAT AND THE MECHANICAL EQUIVALENT OF HEAT

A. The Problem with "Lost" Mechanical Energy

- * Observations: Mechanical energy (kinetic or potential) seems to "disappear" in inelastic collisions, friction, or drilling processes. However, these processes often result in objects getting hot.
- * Suspicion: This suggested a link between the "lost" mechanical energy and the "gained" heat.

B. Joule's Experiment

- * **Setup:** A falling weight (losing potential energy) turns a paddle wheel submerged in an insulated container of water. The paddles churn the water, causing it to heat up.
- * **Measurements:**
- * **Mechanical energy lost:** Calculated from the mass of the falling weight and the distance it falls (mgh).
- * **Heat gained by water:** Calculated using the specific heat formula ($m_{\text{water}} * c_{\text{water}} * \Delta T_{\text{water}}$).
- * **Discovery:** Joule found a constant proportionality between the "lost" mechanical energy (measured in Joules) and the "gained" heat (measured in calories).
- * **The Mechanical Equivalent of Heat:** Approximately 4.184 Joules per calorie (often rounded to 4.2 J/cal).
- * **Implication:** This landmark experiment demonstrated that heat is not a separate caloric fluid but a form of energy that can be interconverted with mechanical energy.

C. Heat as a Form of Energy (Microscopic Interpretation)

- * **Modern Understanding:** Heat is the transfer of energy due to a temperature difference. Internally, this energy corresponds to the disordered, random kinetic energy of the atoms and molecules within a substance.
- * **Connecting Mechanics and Heat:**
- * In a macroscopic moving object (e.g., a car), the kinetic energy is primarily due to the ordered, bulk motion of its constituent atoms.
- * When objects collide inelastically, this macroscopic kinetic energy is converted into the disordered, random kinetic energy of the atoms and molecules, which manifests as an increase in temperature (heat).
- * The Law of Conservation of Energy remains valid when heat is included as a form of energy in the energy balance.