

# THERMODYNAMICS: TEMPERATURE AND HEAT

## 1. Introduction to Thermodynamics

- \* Thermodynamics is the study of heat, temperature, and their relation to energy and work.
- \* This section begins a fresh look at these fundamental concepts, some of which may be familiar from prior studies.

## 2. Defining Temperature

- \* Intuitive Notion of Temperature
- \* Temperature is initially understood through the senses as "hot" or "cold."
- \* For precise physical understanding, this intuition needs refinement.
- \* Thermodynamic Equilibrium
- \* A fundamental concept in thermodynamics.
- \* Definition: A system is in thermal equilibrium when its macroscopic properties (e.g., temperature, pressure, volume) have stopped changing over time.
- \* **Example:**
- \* A cup of hot water, isolated from its surroundings, will eventually reach a stable temperature; it is then in thermal equilibrium. The same applies to a cup of cold water.
- \* If hot and cold water are mixed, the system is initially not in equilibrium (different parts have different temperatures). After a sufficient waiting period, the mixture will reach a uniform, well-defined temperature, signifying a return to equilibrium.
- \* **Distinction:**
- \* Macroscopic properties (like overall temperature) are only well-defined for systems in equilibrium.
- \* Microscopic properties (like individual atomic velocities) are always defined, but thermodynamics focuses on bulk behavior.
- \* Zeroth Law of Thermodynamics
- \* Statement: If two systems (A and B) are each in thermal equilibrium with a third system (C), then A and B are in thermal equilibrium with each other (i.e., they are at

the same temperature).

- \* Significance: This law allows for the consistent definition and measurement of temperature. It enables the use of thermometers to compare temperatures of systems that are not in direct contact. Without it, there would be no global concept of temperature.

- \* Temperature Scales

- \* Quantitative Measurement

- \* To quantify "hot" or "cold," we need a standard and a scale.

- \* Early methods relied on thermal expansion of materials (e.g., expansion of a metal rod or liquid volume).

- \* Liquid-in-Glass Thermometers

- \* Principle: Liquids expand when heated. By placing a liquid in a bulb connected to a narrow capillary tube, even small volume changes cause a significant and visible change in the liquid column's height. This magnifies the expansion.

- \* **Calibration:**

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

- \* **Two reproducible reference points are chosen:**

- \* Freezing/Melting point of water (ice-water mixture at standard atmospheric pressure) defined as 0 degrees Celsius ( $0^{\circ}\text{C}$ ).

- \* Boiling point of water (water-steam mixture at standard atmospheric pressure) defined as 100 degrees Celsius ( $100^{\circ}\text{C}$ ).

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

- \* Fahrenheit Scale: Another common scale with different reference points ( $32^{\circ}\text{F}$  and  $212^{\circ}\text{F}$  for water's freezing and boiling points, respectively).

- \* Limitations of Early Thermometers

- \* Boiling point of water is sensitive to pressure (e.g., lower boiling point at high altitudes).

- \* Different liquids (e.g., mercury vs. alcohol) expand at non-linear rates relative to each other, meaning they would only agree at the calibration points ( $0^{\circ}\text{C}$  and  $100^{\circ}\text{C}$ ), but not necessarily at intermediate temperatures. This requires choosing a specific liquid as a standard, leading to potential inconsistencies.

- \* Gas Thermometer

- \* Principle: For dilute gases, the product of pressure (P) and volume (V) is observed to be linearly proportional to temperature ( $PV \propto T$ ).

- \* Advantage: All \*dilute\* gases exhibit this linear relationship and agree on temperature readings across the entire range once calibrated at two points (e.g., 0°C and 100°C). This makes the gas thermometer a much more reliable and universally consistent instrument than liquid thermometers.

- \* Absolute Zero and the Kelvin Scale

- \* Extrapolation of Gas Thermometer Data: If the PV vs. T graph for any dilute gas is extrapolated to lower temperatures, the product PV approaches zero at a specific temperature: -273.15°C (or -273.16°C in older definitions).

- \* Absolute Zero: This temperature is defined as the absolute zero of temperature, signifying the lowest possible temperature at which the pressure of an ideal gas would theoretically vanish. No further cooling is possible beyond this point.

- \* **Kelvin Scale:**

- \* Definition: The Kelvin scale (K) is an absolute temperature scale where 0 Kelvin (0 K) is set at absolute zero.

- \* Reference Point: The triple point of water (the unique temperature and pressure at which ice, liquid water, and water vapor coexist in thermal equilibrium) is defined as 273.16 K. This provides a single, universally reproducible standard.

- \* Conversion: A temperature difference of 1 Kelvin is equal to 1 degree Celsius. To convert from Celsius to Kelvin, add 273.15 to the Celsius temperature ( $T_K = T_C + 273.15$ ).

- \* Convention: Temperatures on the Kelvin scale are referred to simply as "Kelvin" (e.g., 273 K), not "degrees Kelvin."

- \* Significance: The Kelvin scale is based on fundamental physical principles (behavior of dilute gases) rather than arbitrary material properties or water's phase changes, making it a universal and scientifically robust temperature scale. All future discussions of T (temperature) will refer to Kelvin unless otherwise specified.

### 3. Understanding Heat

- \* Historical Context and Units

- \* Definition: Heat (Q) is the transfer of thermal energy from a hotter body to a colder

body.

- \* Historical View: Initially conceived as a "caloric fluid" that flowed from hot to cold.

- \* **Units:**

- \* Calorie (cal): Historically 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, to be precise).

- \* Kilocalorie (kcal or Calorie): 1 kcal = 1000 cal. Often used for food energy.

- \* Specific Heat Capacity (c)

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

- \* Formula:  $\Delta Q = m * c * \Delta T$

- \*  $\Delta Q$  = amount of heat transferred (calories or Joules)

- \* m = mass of the substance (grams or kilograms)

- \* c = specific heat capacity of the substance (cal/g°C or J/kgK)

- \*  $\Delta T$  = change in temperature (°C or K)

- \* Specific Heat of Water: By definition, the specific heat capacity of water is 1 cal/g°C (or 4186 J/kgK).

- \* Analogies: Similar to coefficients of thermal expansion (linear expansion coefficient  $\alpha$ , volume expansion coefficient  $\beta$ ), which describe how much a material's length or volume changes per degree of temperature change, and depend on the material.

- \* Calorimetry and Heat Exchange

- \* Principle: The total heat exchanged within an isolated system is zero (conservation of heat). Heat lost by hotter bodies equals heat gained by colder bodies.

- \* Formula:  $\sum (m * c * \Delta T) = 0$  for all components in the system.

- \* Example: Mixing hot lead pellets with cold water.

- \* Heat lost by lead:  $Q_{\text{lead}} = m_{\text{lead}} * c_{\text{lead}} * (T_{\text{final}} - T_{\text{initial\_lead}})$

- \* Heat gained by water:  $Q_{\text{water}} = m_{\text{water}} * c_{\text{water}} * (T_{\text{final}} - T_{\text{initial\_water}})$

- \* Conservation:  $Q_{\text{lead}} + Q_{\text{water}} = 0$ . This allows for the calculation of an unknown variable, such as the specific heat of lead ( $c_{\text{lead}}$ ) or the final temperature ( $T_{\text{final}}$ ).

- \* Note: Specific heat capacities are not truly constant but vary slightly with temperature. For basic problems, they are treated as constants over a given temperature range.

- \* Phase Changes and Latent Heat

- \* Phase Change: A transition of a substance from one physical state (solid, liquid, gas) to another. During a phase change, heat is absorbed or released without a change in temperature.

- \* Latent Heat: The energy absorbed or released per unit mass during a phase change.

- \* Latent Heat of Fusion ( $L_f$  or  $L_m$ ): Heat required to change a substance from solid to liquid (melting) or liquid to solid (freezing).

- \* Formula:  $Q = m * L_f$

- \* For water:  $L_f \approx 80 \text{ cal/g}$  (or  $334 \text{ kJ/kg}$ ).

- \* Latent Heat of Vaporization ( $L_v$ ): Heat required to change a substance from liquid to gas (boiling/evaporation) or gas to liquid (condensation).

- \* Formula:  $Q = m * L_v$

- \* For water:  $L_v \approx 540 \text{ cal/g}$  (or  $2260 \text{ kJ/kg}$ ).

- \* **Calorimetry Problems with Phase Changes:**

- \* These problems involve multiple steps as a substance undergoes temperature changes and potentially phase changes.

- \* Example: Heating ice from  $-30^\circ\text{C}$  to steam at  $100^\circ\text{C}$ .

1. Heat ice from  $-30^\circ\text{C}$  to  $0^\circ\text{C}$ :  $Q_1 = m_{\text{ice}} * c_{\text{ice}} * \Delta T$  (Note:  $c_{\text{ice}} \neq c_{\text{water}}$ ).

2. Melt ice at  $0^\circ\text{C}$ :  $Q_2 = m_{\text{ice}} * L_f$ .

3. Heat water from  $0^\circ\text{C}$  to  $100^\circ\text{C}$ :  $Q_3 = m_{\text{water}} * c_{\text{water}} * \Delta T$ .

4. Vaporize water at  $100^\circ\text{C}$ :  $Q_4 = m_{\text{water}} * L_v$ .

5. Heat steam above  $100^\circ\text{C}$ :  $Q_5 = m_{\text{steam}} * c_{\text{steam}} * \Delta T$ .

- \* Complex scenarios (e.g., ice-water mixtures): If the final state is unknown, one might assume a final state (e.g., all water above  $0^\circ\text{C}$ , all ice below  $0^\circ\text{C}$ , or a mixture at  $0^\circ\text{C}$ ) and check for consistency. If the calculation yields an impossible result (e.g., a negative final temperature when assuming all melted), the initial assumption is wrong, and another state must be considered.

#### 4. Modes of Heat Transfer

- \* Radiation

- \* Mechanism: Transfer of energy via electromagnetic waves.

- \* Does not require a medium; can travel through a vacuum.

- \* Example: Heat from the sun, heat from a glowing electric heater coil.

- \* Convection

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

- \* Process: Heated portions of the fluid become less dense and rise, while cooler, denser fluid sinks, creating convection currents that circulate heat.

- \* Example: Boiling water in a pot, atmospheric circulation (wind).

- \* Conduction

- \* Mechanism: Transfer of heat through direct contact between particles (atoms, molecules) within a material or between two materials in contact, without macroscopic movement of the material itself. Vibrational energy is passed from hotter to colder regions.

- \* Example: Heat moving along a metal spoon from a hot cup of tea.

- \* Heat Conduction Formula: The rate of heat flow ( $\Delta Q/\Delta t$ )

- \* Formula:  $\Delta Q/\Delta t = \kappa * A * (\Delta T/\Delta x)$

- \*  $\Delta Q/\Delta t$  = rate of heat transfer (Joules per second or Watts)

- \*  $\kappa$  (kappa) = thermal conductivity (material property, indicating how well a material conducts heat).

- \*  $A$  = cross-sectional area through which heat is flowing.

- \*  $\Delta T$  = temperature difference between the hot and cold ends.

- \*  $\Delta x$  = length or thickness of the material.

- \* Significance: Heat flows faster through materials with high thermal conductivity, through larger areas, and over steeper temperature gradients (large  $\Delta T$  over small  $\Delta x$ ).

- \* Thermal Reservoir: A body so large that its temperature does not change significantly when heat is added to or removed from it. Examples include large bodies of water (like an ocean) or the atmosphere (a room).

## 5. The Nature of Heat: Mechanical Equivalent

- \* Historical Problem: Inelastic collisions or friction often resulted in "lost" mechanical energy, yet the colliding bodies or objects experiencing friction would heat up. This suggested a link between mechanical energy and heat.

- \* Joule's Experiment

- \* Setup: A container of water with paddles immersed. Weights falling under gravity

turned the paddles, churning the water.

- \* Observation: The mechanical potential energy lost by the falling weights ( $Mgh$ ) caused the water to heat up.
- \* Conclusion: Joule demonstrated a direct proportionality between the lost mechanical energy (measured in Joules) and the heat gained by the water (measured in calories).
- \* Mechanical Equivalent of Heat: 1 calorie  $\approx$  4.184 Joules (often approximated as 4.2 J/cal). This established that heat is a form of energy.
  
- \* Heat as Microscopic Kinetic Energy
- \* Modern View: Heat is understood as the total random kinetic energy of the atoms and molecules that constitute a substance.
- \* **Distinction:**
- \* Macroscopic Kinetic Energy: The ordered motion of a whole object (e.g., a car moving).
- \* Thermal Energy (Heat): The disordered, random motion of individual molecules within the object.
- \* Reconciliation of Energy Conservation: When macroscopic kinetic energy is "lost" in an inelastic collision, it is converted into increased random kinetic energy of the constituent molecules, causing the object to heat up. The total energy (macroscopic kinetic + thermal + other forms) remains conserved.