THERMODYNAMICS: FUNDAMENTALS OF HEAT AND TEMPERATURE

ADMINISTRATIVE ANNOUNCEMENTS

*   Problem Set Schedule: A new problem set will be assigned today, due next Wednesday. A final problem set with 2-3 problems will be assigned based on upcoming lecture material.

INTRODUCTION TO THERMODYNAMICS

*   Thermodynamics is a new topic, focusing on the study of heat, temperature, and related phenomena.
*   Some concepts may have been introduced in high school.
*   The next four lectures will delve deeply into these subjects.

TEMPERATURE

*   INTUITIVE NOTION OF TEMPERATURE

*   We all possess an intuitive understanding of "hot" and "cold." This intuition serves as a starting point, but physicists require more precise definitions.

*   THERMAL EQUILIBRIUM

*   DEFINITION: A system is in thermal equilibrium when its macroscopic properties (e.g., temperature) have stopped changing and remain constant over time. This implies isolation from the outside world.
*   ANALOGY: Similar to mechanical equilibrium where a system's mechanical properties (e.g., position, velocity) are stable.

*   EXAMPLES:

*   An isolated cup of hot water or cold water will settle into and maintain a uniform temperature.
*   If hot and cold water are mixed, the system is initially NOT in equilibrium (temperature varies spatially). Given sufficient time, the mixture will reach a uniform,

well-defined temperature, signifying a return to equilibrium.

*   Gas in a Piston: A gas confined by a piston with weights is in equilibrium if the piston is stationary and gas properties are uniform. If weights are suddenly removed, the piston moves, and the gas becomes turbulent with varying pressure and temperature. After some time, it settles into a new equilibrium state.

*   MACROSCOPIC VS. MICROSCOPIC:

*   At the macroscopic level, temperature and thermal equilibrium are only well-defined when the system has settled down.

*   At the microscopic level, individual atoms and molecules always have defined states (location, velocity), but these fine details are not what macroscopic temperature describes.

*   THE 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. In simpler terms, if A and C are at the same temperature, and B and C are at the same temperature, then A and B are at the same temperature.

*   SIGNIFICANCE: This law, though seemingly obvious, is fundamental. It allows for the global comparison and definition of temperature using a third system (like a thermometer). Without it, two objects that never directly interacted could not be reliably said to have the same temperature if they both read the same on a thermometer. It establishes temperature as a transitive property.

*   QUANTIFYING TEMPERATURE

*   The intuitive hot/cold distinction is insufficient; we need quantitative scales.
*   THERMOMETRIC PROPERTIES: We seek physical properties that reliably change with temperature.
*   Thermal Expansion: Materials expand when heated.
*   Solid Rods: A metal rod expands linearly with increasing temperature ($\Delta L = \alpha L_0 \Delta T$), where α is the coefficient of linear expansion (material-dependent).
*   Liquids: Liquids also expand when heated (e.g., fuel in a tank requires room for expansion).
*   Liquid-in-Glass Thermometers: Utilize the principle of liquid expansion. A large

reservoir of fluid (e.g., mercury, alcohol) expands into a very narrow evacuated tube, magnifying the change in volume into a visible change in height. This requires a prism for magnification.

* Gas Thermometers: Offer a more universal and precise method for temperature measurement.

* Principle: For a dilute gas at constant volume, the product of pressure (P) and volume (V) is linearly proportional to its temperature ($PV \propto T$).

* Advantage: Unlike liquids, which expand at different non-linear rates, different dilute gases, when calibrated at two points, will agree on temperatures across the entire range because their PV product is linear with temperature. This linearity makes gas thermometers highly reliable.

* TEMPERATURE SCALES

* CALIBRATION: Requires defining reference points that are universally reproducible.

* CENTIGRADE (CELSIUS) SCALE:

* Freezing Point of Water: Defined as 0 degrees Celsius (0 °C). This is the temperature at which water and ice coexist at standard atmospheric pressure.

* Boiling Point of Water: Defined as 100 degrees Celsius (100 °C). This is the temperature at which water and steam coexist at standard atmospheric pressure.

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

* LIMITATIONS: The boiling point of water varies with atmospheric pressure (e.g., lower at higher altitudes), making it less precise for universal standards.

* FAHRENHEIT SCALE: Another common scale with different reference points (32 °F for freezing water, 212 °F for boiling water), dividing the interval into 180 degrees.

* ABSOLUTE (KELVIN) SCALE:

* Derivation: Extrapolating the linear PV vs. T relationship of ideal gases to lower temperatures reveals that the PV product vanishes at a specific temperature: -273.15 °C. This temperature is identified as the absolute zero of temperature.

* ABSOLUTE ZERO (0 K): The theoretical lowest possible temperature, where all molecular motion ceases and pressure of an ideal gas becomes zero. It is a fundamental physical limit.

*   TRIPLE POINT OF WATER: To define the Kelvin scale precisely, a single, highly reproducible reference point is needed. This is the triple point of water—the unique temperature and pressure at which ice, liquid water, and water vapor can coexist in stable equilibrium. This point is defined as 273.16 Kelvin (273.16 K).

*   Relationship to Celsius: 0 °C = 273.15 K. Therefore, $T(K) = T(°C) + 273.15$.

*   Notation: Temperatures on the Kelvin scale are referred to as "Kelvin" (e.g., 100 Kelvin), not "degrees Kelvin."

HEAT (Q)

*   DEFINITION: Heat (Q) is the transfer of thermal energy between systems due to a temperature difference.

*   HISTORICAL CONCEPT: Initially, heat was theorized as a fluid-like substance called "caloric fluid" that flowed from hotter to colder objects.

*   UNITS:

*   CALORIE (cal): Historically defined as the amount of heat required to raise the temperature of 1 gram of water by 1 degree Celsius.

*   KILOCALORIE (kcal or Calorie with a capital 'C'): 1000 calories, often used for food energy.

*   **SPECIFIC HEAT CAPACITY (c):**

*   DEFINITION: The amount of heat required to raise the temperature of 1 unit mass of a substance by 1 degree. It is a material-dependent property.

*   FORMULA: $Q = mc\Delta T$

* Q = heat transferred (calories or joules)

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

* c = specific heat capacity (e.g., cal/g°C or J/kgK)

* ΔT = change in temperature (°C or K)

*   WATER: By definition, the specific heat of water is 1 cal/g°C (or 4186 J/kgK).

*   CALORIMETRY PROBLEMS: These problems involve mixing substances at different temperatures and calculating the final temperature or an unknown specific heat. The principle of conservation of heat applies: total heat gained by one substance equals total heat lost by another ($\Sigma Q = 0$).

*   Example: Mixing a hot metal (e.g., lead) with cold water. Heat lost by lead = Heat

gained by water. m_lead * c_lead * (T_final - T_initial_lead) + m_water * c_water * (T_final - T_initial_water) = 0.

* PHASE CHANGES AND LATENT HEAT

* PHASE CHANGE: A transition of a substance from one state of matter to another (e.g., solid to liquid, liquid to gas). During a phase change, the temperature of the substance remains constant, even though heat is being added or removed.

* HEAT CURVE: If heat is continuously added to a substance (e.g., ice at -30°C), its temperature rises until a phase change occurs (e.g., melting at 0°C). During the phase change, the temperature plateaus. Once the phase change is complete, the temperature rises again.

* LATENT HEAT (L): The amount of heat required per unit mass to change the phase of a substance at a constant temperature.

* FORMULA: $Q = mL$

* Q = heat transferred

* m = mass of the substance

* L = latent heat (e.g., cal/g or J/kg)

* TYPES OF LATENT HEAT:

* Latent Heat of Fusion (L_f): For melting (solid to liquid) or freezing (liquid to solid). For water, $L_f \approx 80$ cal/g (334 kJ/kg).

* Latent Heat of Vaporization (L_v): For boiling/evaporation (liquid to gas) or condensation (gas to liquid). For water, $L_v \approx 540$ cal/g (2260 kJ/kg).

* COMPLEX PROBLEMS: Involve multiple steps of temperature change and phase change. One must account for all $Q = mc\Delta T$ and $Q = mL$ terms. Often requires assuming a final state (e.g., all liquid) and checking consistency (e.g., final temperature is within the assumed phase range).

HEAT TRANSFER MECHANISMS

Heat can be transferred by three primary mechanisms:

* RADIATION

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

* Medium: Does not require a medium; can occur through a vacuum (e.g., heat from

the Sun).

* Example: Heat from a space heater with glowing coils.

* CONVECTION

* Mechanism: Heat transfer through the bulk movement of fluids (liquids or gases).

* Process: When a fluid is heated, it expands, becomes less dense, and rises. Cooler, denser fluid sinks to take its place, creating convection currents.

* Example: Boiling water in a pot, atmospheric circulation.

* CONDUCTION

* Mechanism: Heat transfer through direct contact between particles, without macroscopic movement of the material itself. Energy is transferred via collisions between atoms/molecules.

* Rate of Heat Conduction (Fourier's Law): $Q/t = \kappa A(\Delta T/\Delta x)$

* Q/t = rate of heat flow (power, e.g., Joules/second or Watts)

* κ (kappa) = thermal conductivity of the material (material-dependent property; high for metals, low for insulators like wood or air).

* A = cross-sectional area through which heat flows.

* ΔT = temperature difference between the hot and cold ends.

* Δx = length/thickness of the material.

* Example: A metal spoon heating up when one end is placed in hot soup; a wooden handle on a skillet.

* RESERVOIR: A conceptual body so large that its temperature remains constant despite heat transfer to or from it (e.g., an ocean, a large room).

HEAT AS A FORM OF ENERGY

* Historical CONTEXT: Early physicists observed that mechanical energy often seemed to "disappear" in processes like inelastic collisions or friction, while simultaneously producing heat (e.g., colliding cars become hot, drilling generates heat).

* JOULE'S EXPERIMENT:

* SETUP: A weight attached to a paddle system submerged in water. As the weight falls, its gravitational potential energy is converted into mechanical work that churns the water.

*   OBSERVATION: The churning water heats up.

*   CONCLUSION: Joule demonstrated a direct proportionality between the amount of mechanical energy lost and the amount of heat generated.

*   MECHANICAL EQUIVALENT OF HEAT: Approximately 4.184 Joules (J) = 1 calorie (cal). This established heat as a form of energy.

*   MODERN UNDERSTANDING: Heat is the transfer of thermal energy, which at the microscopic level, corresponds to the random kinetic energy of atoms and molecules within a substance.

*   In an inelastic collision, the macroscopic kinetic energy of the colliding objects is transformed into the increased random kinetic energy of their constituent molecules, resulting in an increase in temperature (heat).

*   The Law of Conservation of Energy remains valid when heat is included as a form of energy.