Engineering Thermodynamics (DI03019011)

GTU Course Presentation

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Introduction and applications of Engineering thermodynamics. I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 1: Basic Concepts of Thermodynamics
- Lecture 1: Introduction and Applications of Engineering Thermodynamics

Agenda I

- 1. Definition and Origins of Thermodynamics
- 2. Macroscopic vs. Microscopic Viewpoints
- 3. Thermodynamic Systems, Boundaries, and Surroundings
- 4. Classification of Systems (Closed, Open, Isolated)
- 5. Piston-Cylinder Assembly as a Fundamental System
- 6. Thermodynamic Properties (Intensive vs. Extensive)
- 7. Engineering Applications of Thermodynamics

What is Thermodynamics? I

- Thermodynamics is the science that describes energy, its transformations, and its relation to states of matter.
- The term originates from the Greek words 'therme' (heat) and 'dynamis' (power), which originally related to early efforts to convert heat into mechanical work.
- It deals with the transition of energy from one form to another, specifically how thermal energy relates to mechanical, electrical, and chemical work.
- Governed by four fundamental laws: Zeroth Law (defines temperature), First Law (energy conservation), Second Law (entropy and process direction), and Third Law (absolute zero).

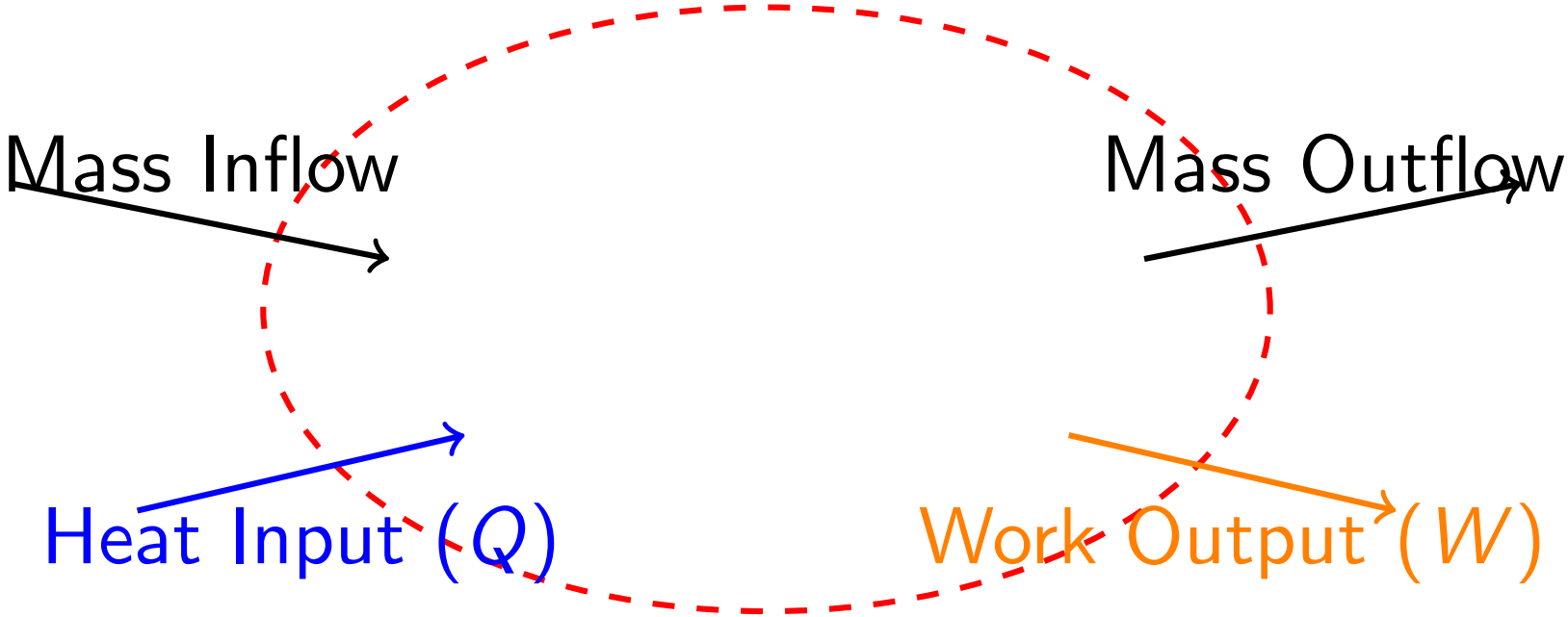
Historical Context of Thermodynamics I

- Developed primarily in the 19th century to address the operating efficiency of early industrial steam engines.
- Thomas Newcomen (1712) built the first atmospheric engine, which had low thermal efficiency (0.5 percent).
- James Watt (1769) introduced the separate condenser, dramatically improving efficiency and fueling the Industrial Revolution.
- Sadi Carnot (1824) published the foundations of the second law, defining the maximum possible efficiency limits of heat engines.
- Theoretical foundations were formalized by Rudolf Clausius (concept of entropy) and Lord Kelvin (absolute temperature scale) around 1850.

Macroscopic vs. Microscopic Viewpoints I

- Macroscopic (Classical) Thermodynamics: Focuses on the gross or average behavior of large quantities of matter.
- Requires no hypothesis about the detailed molecular structure of the substance; bulk properties are directly measurable.
- Variables of interest include Pressure (P), Volume (V), and Temperature (T). This is the primary approach used in engineering design.
- Microscopic (Statistical) Thermodynamics: Analyzes the behavior of individual molecules and quantum energy states.
- Uses statistical probability to relate molecular-level motion and interactions to bulk macroscopic properties.
- Essential for low-density gases (e.g., high-altitude aerodynamics), plasma physics, and nanotechnology.

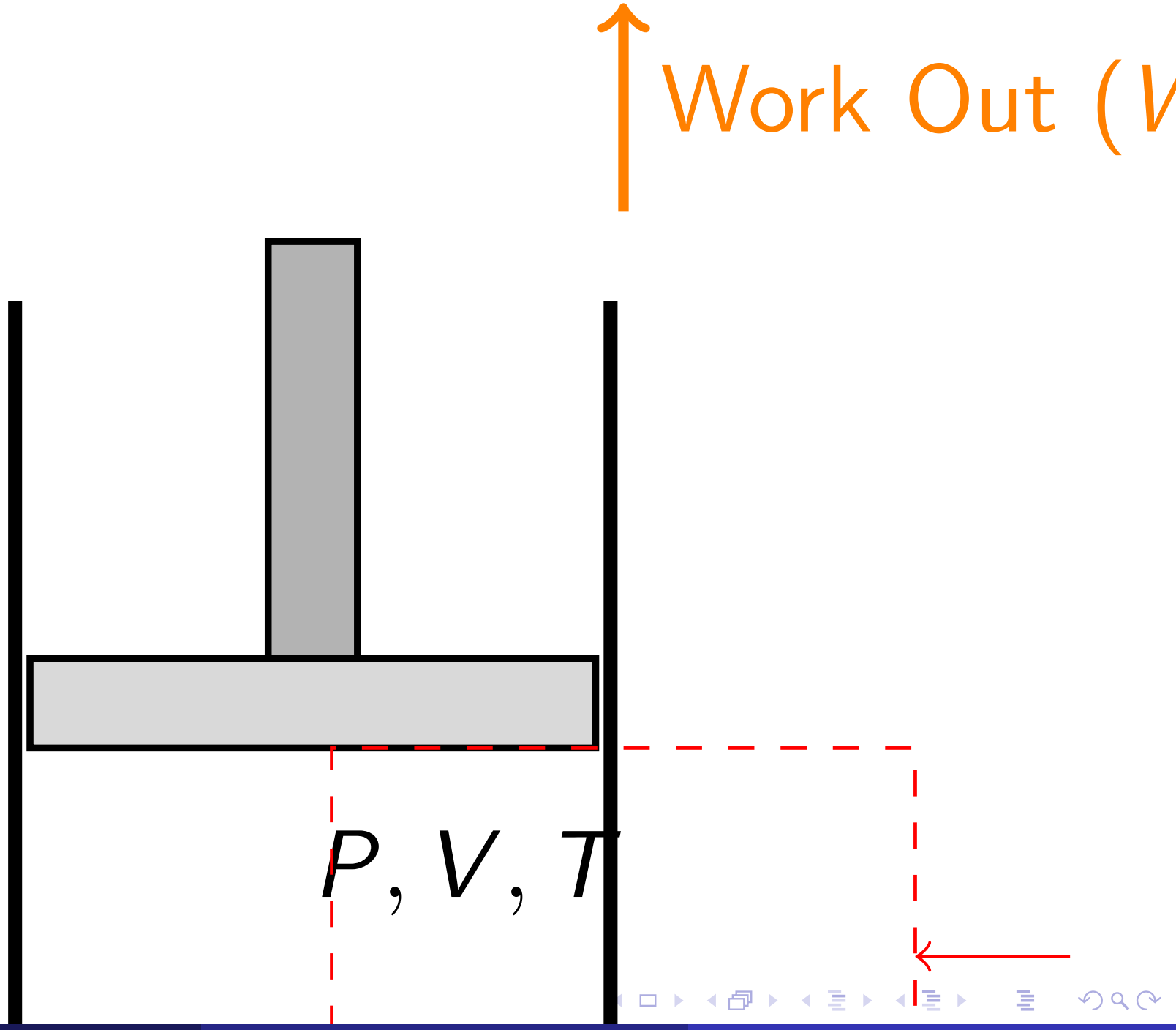
Thermodynamic System, Boundary, and Surroundings



Classification of Thermodynamic Systems I

- Closed System (Control Mass): Contains a fixed quantity of mass. No mass flow crosses the boundary, but heat and work can pass through.
- Open System (Control Volume): Mass, heat, and work can all cross the control surface boundary. Examples: turbines, compressors, nozzles.
- Isolated System: A system completely isolated from its surroundings. Neither mass nor energy (heat or work) can cross the boundary.
- Adiabatic System: A specialized system where no heat is transferred across the boundary ($Q = 0$); energy transfer can only occur as work.

The Piston-Cylinder Assembly



Thermodynamic Properties I

- A property is any macroscopic characteristic of a system in equilibrium to which a numerical value can be assigned.
- Intensive Properties: Independent of the mass or size of the system. Examples: Temperature (T), Pressure (P), and Density ($\$\rho$).
- Extensive Properties: Dependent on the size or extent of the system. Examples: Total Mass (m), Volume (V), and Total Internal Energy (U).
- Specific Properties: Extensive properties per unit mass, making them intensive. Examples: Specific Volume (v = V/m), Specific Enthalpy (h = H/m).

Engineering Applications of Thermodynamics I

- Power Generation: Steam power plants (Rankine Cycle), Gas turbine engines (Brayton Cycle), and Nuclear power systems.
- Propulsion Systems: Internal combustion engines (Otto and Diesel Cycles) in automobiles, turbojet/turbofan engines in aviation.
- Refrigeration and HVAC: Vapor compression cycles for household refrigerators, building air conditioners, and industrial heat pumps.
- Renewable Energy: Concentrated solar thermal power (CSP), geothermal heat recovery, and hydrogen fuel cells.
- Biological and Chemical Systems: Energy conversion in metabolic processes, chemical reactors, and life support systems.

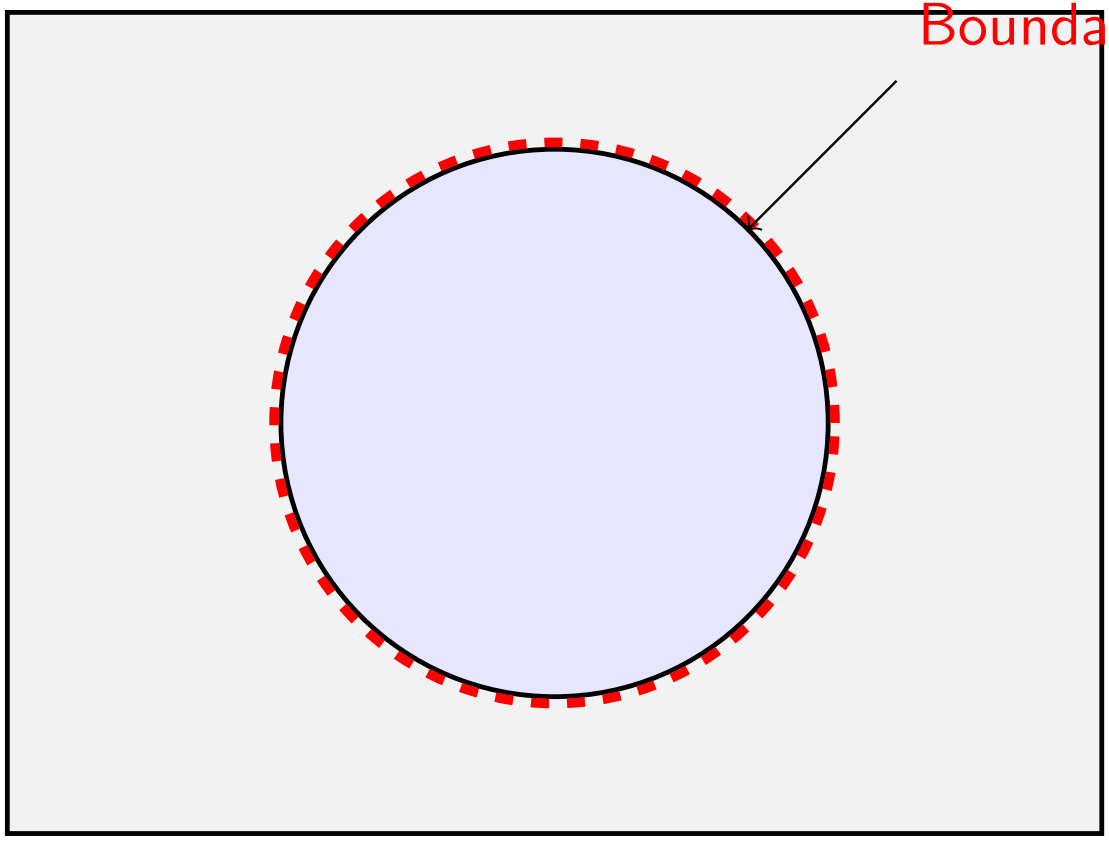
Basic thermodynamic Concepts. I

- Course: Engineering Thermodynamics
- Unit: Basic Concepts of Thermodynamics
- Lecture 2: Basic Thermodynamic Concepts
- An essential foundation for analyzing energy transfer, system boundaries, and state changes.

Agenda I

- 1. Definition of System, Boundary, Surroundings, and Universe
- 2. Classification of Systems: Closed, Open, and Isolated
- 3. Microscopic (Statistical) vs. Macroscopic (Classical) Thermodynamics
- 4. Thermodynamic Properties: Intensive vs. Extensive
- 5. Concepts of State, Process, Path, and Cycle
- 6. Thermodynamic Equilibrium and Quasi-Static Processes

The Thermodynamic System, Boundary, and Surroundings



Classification of Thermodynamic Systems I

- Closed System (Control Mass): Consists of a fixed amount of mass. No mass can cross the boundary, but energy (heat and work) can enter or leave.
- Open System (Control Volume): A properly selected region in space (e.g., compressor, turbine, nozzle). Both mass and energy can cross the control boundary (control surface).
- Isolated System: A system that is completely uninfluenced by its surroundings. Neither mass nor energy can cross the boundary (e.g., the universe as a whole, or a perfectly insulated rigid tank).
- Boundary Characteristics: Boundaries can be fixed or movable, and real or imaginary, depending on the system configuration.

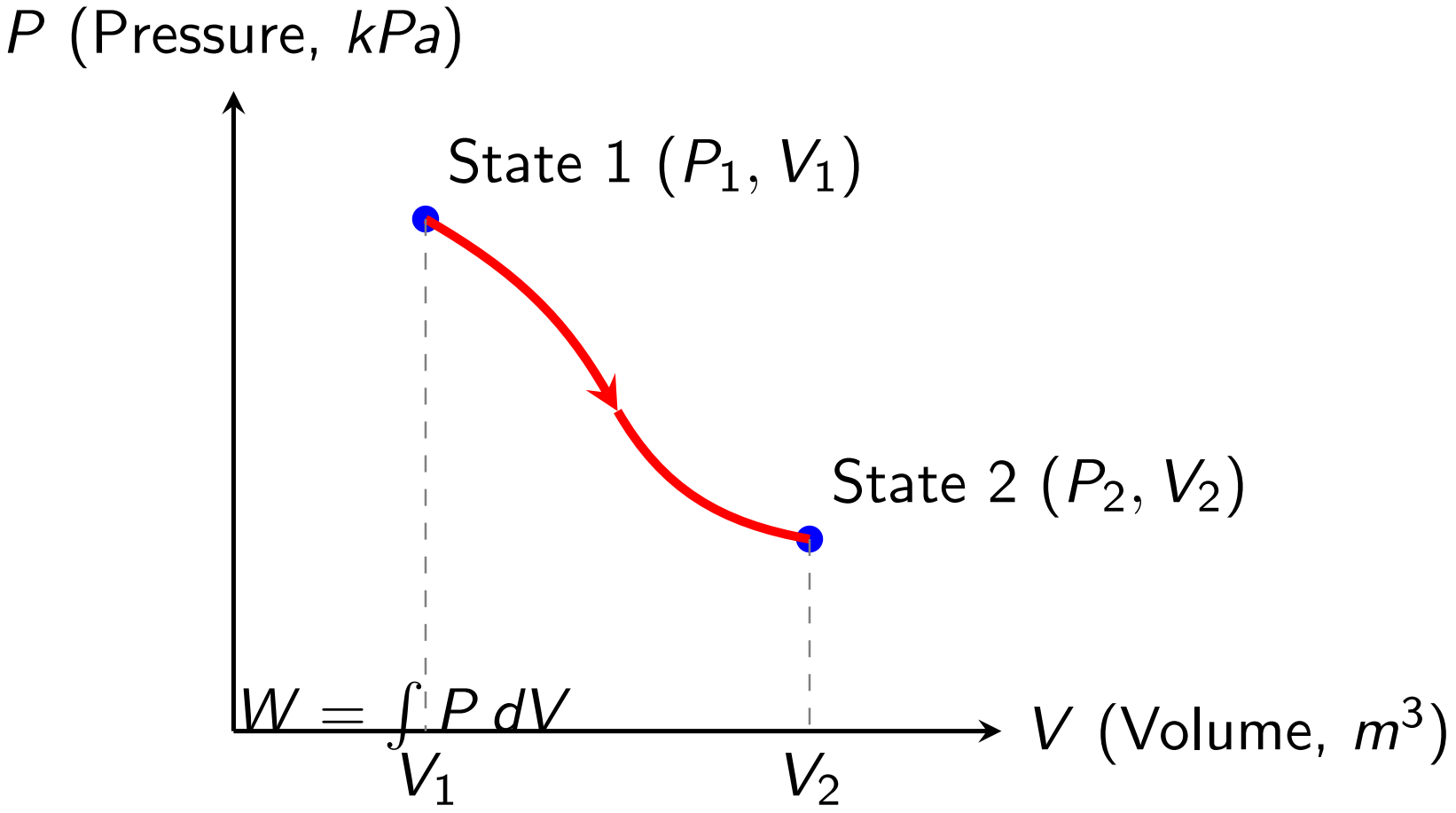
Macroscopic vs. Microscopic Viewpoints I

- Macroscopic Viewpoint (Classical Thermodynamics): Focuses on the gross or average effects of many molecules. Properties (like pressure and temperature) can be measured directly with instruments.
- Microscopic Viewpoint (Statistical Thermodynamics): Analyzes the average behavior of individual molecules and uses probability theory to relate molecular properties to macroscopic behavior.
- Classical vs. Statistical: Classical thermodynamics requires no hypothesis about the detailed structure of matter, making it simpler and highly practical for engineering applications.
- Continuum Concept: Matter is treated as continuous and homogeneous, ignoring molecular voids. This holds valid as long as the system size is much larger than the mean free path of the molecules.

Thermodynamic Properties: Intensive vs. Extensive I

- Thermodynamic Property: Any observable or measurable characteristic of a system in equilibrium (e.g., pressure, temperature, volume, mass).
- Intensive Properties: Independent of the mass or size of the system. Examples: Temperature (T), Pressure (P), Density (ρ), and Specific Volume (v).
- Extensive Properties: Directly proportional to the mass or size of the system. Examples: Total Mass (m), Total Volume (V), Total Internal Energy (U).
- Specific Properties: Extensive properties divided by mass, which makes them intensive (e.g., specific volume $v = V/m$, specific enthalpy $h = H/m$).
- State Postulate: The thermodynamic state of a simple compressible system is completely specified by two independent, intensive properties.

Thermodynamic States, Processes, and Paths



Thermodynamic Equilibrium I

- Equilibrium State: A state of balance where there are no unbalanced potentials or driving forces within the system, and no changes occur over time if isolated.
- Thermal Equilibrium: The temperature is uniform throughout the system. No heat transfer occurs within the system or with the surroundings if they are at the same temperature.
- Mechanical Equilibrium: No unbalanced forces exist within the system. The pressure at any point is constant over time (though it may vary with depth due to gravity in liquids).
- Phase Equilibrium: For a multi-phase system, the mass of each phase remains constant, indicating no net phase transition.
- Chemical Equilibrium: The chemical composition of the system does not change over time; no chemical reactions occur.

The Quasi-Static (Quasi-Equilibrium) Process I

- Definition: A process carried out sufficiently slowly such that the system deviates from thermodynamic equilibrium by only an infinitesimal amount at any instant.
- Path Representation: Because the system is always in virtual equilibrium, all intermediate states are well-defined. Thus, a quasi-static process can be drawn as a solid line on a property diagram (e.g., P-V diagram).
- Non-Quasi-Static Process: A rapid process where intermediate states are not in equilibrium. Properties vary spatially and cannot be represented by a single value or path line (typically shown as a dashed line).
- Engineering Significance: Quasi-static processes are idealized models that yield the maximum work output (for expansion) or minimum work input (for compression).

Summary and Key Takeaways I

- Core Terminology: System, Boundary, Surroundings, and Universe form the basic framework for thermodynamic analysis.
- Closed vs. Open: Closed systems exchange only energy; open systems exchange both mass and energy.
- Properties and State: State is defined by independent intensive properties. Intensive properties are mass-independent, while extensive properties scale with mass.
- Equilibrium: Complete thermodynamic equilibrium requires thermal, mechanical, chemical, and phase equilibrium.
- Processes: Quasi-static processes serve as the reversible, maximum-efficiency limits for real thermodynamic machinery.

State, System, Boundary and Surroundings. I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 1: Basic Concepts of Thermodynamics
- Topic: Thermodynamic Systems, Boundaries, Surroundings, and States

Agenda I

- Thermodynamic System Definition
- Surroundings and Boundaries
- Classification: Open, Closed, and Isolated Systems
- Thermodynamic State and Properties
- Equilibrium Types: Thermal, Mechanical, Phase, and Chemical
- The State Postulate

Defining the Thermodynamic System I

- System: A quantity of matter or a region in space chosen for study in thermodynamic analysis.
- Surroundings: The mass or region outside the system boundary that can affect or be affected by the system.
- Boundary: The real or imaginary surface that separates the system from its surroundings.
- Mathematical formulation: Universe = System + Surroundings.
- The boundary of a system can be fixed or movable, and real or imaginary.
- The boundary has zero thickness, contains no mass, and occupies no volume. Thus, no mass or energy can be stored within it.

(Quantity of matter
or region in space);
ode[align=center, font=] **at (3.5,0.8) SURROUNDINGS
(Everything outside
the system); [-¿, thick] (2.2, -1) – (1.5, -0.5);
ode[right] at (2.2, -1) Boundary; [thick, ¡-¿] (-3,-2.5) –
(3,-2.5) node[midway, below] UNIVERSE = SYSTEM +
SURROUNDINGS;**

Types of Systems: Closed, Open, and Isolated I

- Closed System (Control Mass): Consists of a fixed amount of mass. No mass can cross its boundary. Energy (heat and work) can cross the boundary.
- Open System (Control Volume): A region in space enclosing a device (e.g., compressor, turbine, nozzle) where mass and energy flow are present.
- Isolated System: A system where neither mass nor energy can cross its boundary (e.g., the universe or a perfectly insulated container).
- Control Surface: The physical or imaginary boundary of a control volume, which can be fixed or moving.

(Control Mass); [red,-¿,thick] (-3.8,0.5) – node[above, font=] Heat/Work (-3,0.5); [red,¡-,thick] (-3.8,-0.5) – node[below, font=] Work/Heat (-3,-0.5); [line width=1.5pt, double, red] (-1.75, -1.8) – (-1.75, -2.2) node[below, font=] No Mass Flow; [thick] (1,-1.5) rectangle (3.5,1.5);

ode[align=center, font=] at (2.25,0) Open System

(Control Volume); [blue,-¿,thick] (0.2,0.8) – node[above, font=] Mass In (1,0.8); [blue,-¿,thick] (3.5,-0.8) – node[below, font=] Mass Out (4.3,-0.8);

[red,-¿,thick] (2.25,2) – node[right, font=] Energy In/Out (2.25,1.5);

Concept of Thermodynamic State I

- Thermodynamic State: The condition of a system as described by its specific thermodynamic properties (e.g., pressure, temperature, volume).
- At a given state, all the properties of a system have fixed, unique values.
- If the value of even one property changes, the state of the system changes to a different one.
- Thermodynamics focuses on equilibrium states where no unbalanced potentials (thermal, mechanical, chemical) exist within the system.

Thermodynamic Equilibrium I

- Thermal Equilibrium: Temperature is uniform throughout the entire system.
- Mechanical Equilibrium: No change in pressure at any point in the system over time, and no unbalanced forces exist.
- Phase Equilibrium: For multi-phase systems, the mass of each phase remains constant and does not change with time.
- Chemical Equilibrium: The chemical composition of the system remains unchanged, meaning no chemical reactions are occurring.

The State Postulate I

- State Postulate: The state of a simple compressible system is completely specified by two independent, intensive properties.
- Simple Compressible System: A system in which electrical, magnetic, gravitational, motion, and surface tension effects are absent or negligible.
- Independent Properties: Properties where one can be varied while the other is held constant (e.g., temperature and specific volume).
- If a system is not simple compressible, additional properties (like velocity, elevation, magnetic field strength) must be specified.

Summary and Next Steps I

- System, boundary, and surroundings form the basic framework for thermodynamic analysis.
- Systems are classified as open, closed, or isolated based on mass and energy transfer.
- The state of a system is its condition defined by properties, and equilibrium implies a balance of potentials.
- The State Postulate allows us to define the state of a simple compressible system using exactly two independent intensive properties.
- Next Lecture: Processes, paths, cycles, and the Zeroth Law of Thermodynamics.

Types of Systems and boundaries. I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 1: Basic Concepts of Thermodynamics
- Instructor: Professor of Mechanical Engineering
- Topic 1.4: Detailed classification of thermodynamic systems and boundary characteristics.

Agenda I

- Defining the System, Boundary, and Surroundings
- Key Boundary Types: Real vs. Imaginary, Fixed vs. Movable, Adiabatic vs. Diathermanous
- Closed Systems (Control Mass) with examples
- Open Systems (Control Volume) with examples
- Isolated Systems and the Universe concept
- Summary and Comparison of System Behaviors

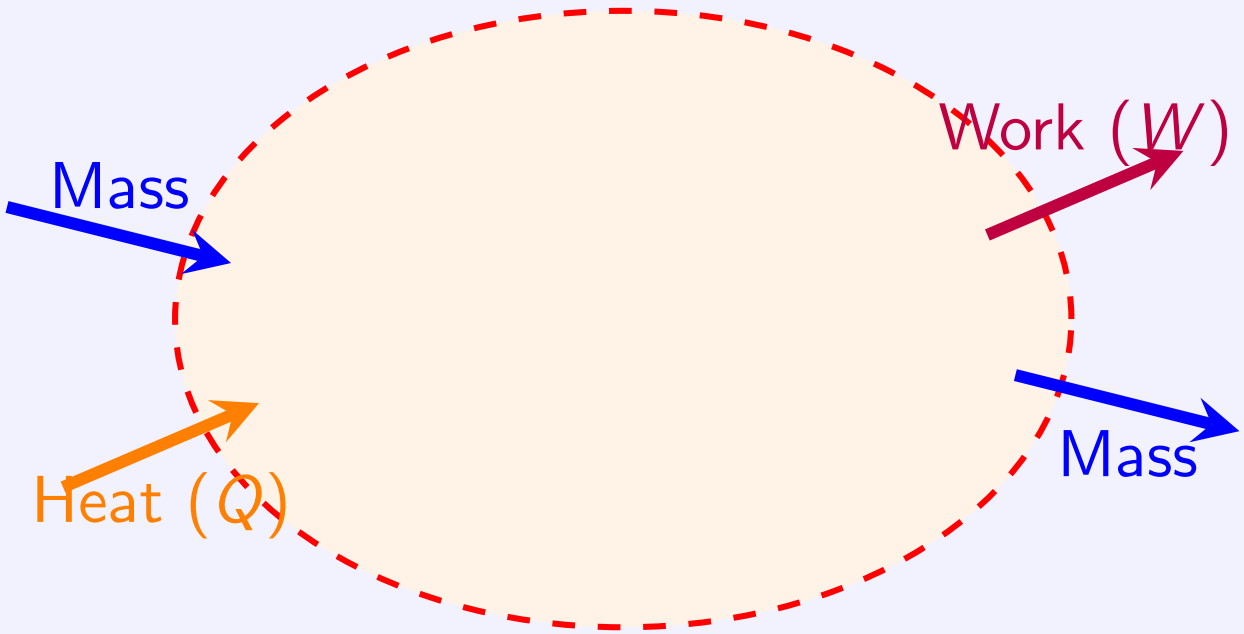
Thermodynamic System, Surroundings, and Boundary I

- A thermodynamic system is defined as a quantity of matter or a region in space chosen for study.
- The mass or region outside the system is called the surroundings.
- The real or imaginary surface that separates the system from its surroundings is called the boundary.
- The boundary of a system can be fixed or movable.
- Mathematically, the boundary has zero thickness, and thus it can neither contain any mass nor occupy any volume in space.

Characteristics of Boundaries I

- Real Boundary: A physical surface (e.g., the metal wall of a tank or cylinder).
- Imaginary Boundary: A conceptual surface chosen for analysis convenience (e.g., the open ends of a pipe or nozzle flow).
- Fixed Boundary: Does not move during a process; no boundary work (PdV work) can be performed.
- Movable Boundary: Can move, allowing the system to perform expansion or compression work.
- Adiabatic Boundary: Completely insulated; prevents heat transfer ($Q = 0$).
- Diathermanous Boundary: Conductive; allows heat exchange between system and surroundings.

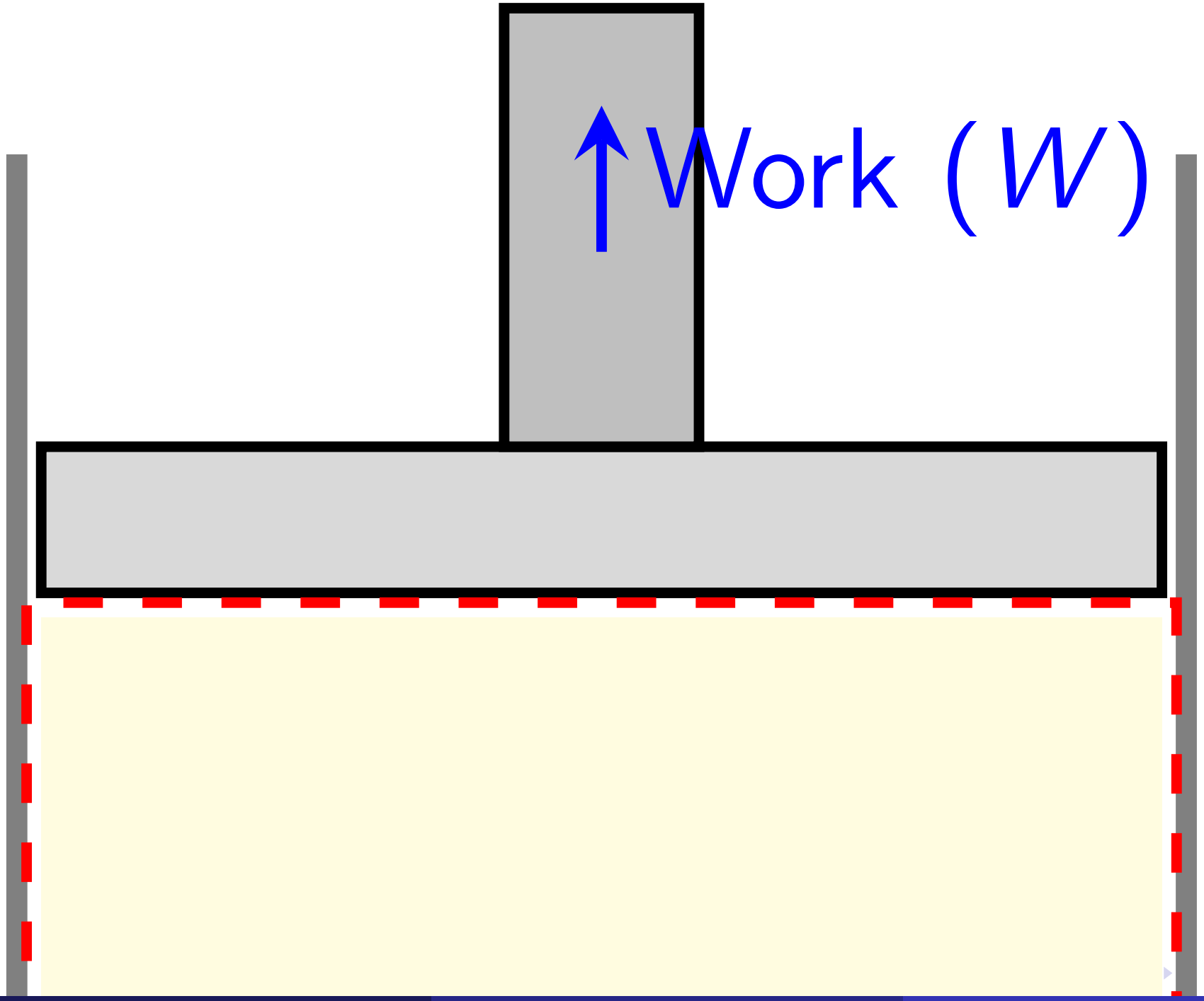
System, Boundary, and Surroundings Visualized



Closed System (Control Mass) I

- A closed system (also known as a control mass) consists of a fixed amount of mass.
- No mass can cross the system boundary (mass flow in/out = 0).
- Energy, in the form of heat and work, can cross the boundary.
- The volume of a closed system does not have to be fixed; it can expand or contract (e.g., piston-cylinder device).
- Example: A sealed container of gas or liquid being heated or compressed.

Closed System: Piston-Cylinder Assembly



Open System (Control Volume) I

- An open system (also known as a control volume) is a properly selected region in space.
- Both mass and energy (in the form of heat and work) can cross the system boundary.
- The boundary of a control volume is referred to as a control surface.
- Control surfaces can be real physical walls combined with imaginary open areas.
- Examples: Compressors, turbines, nozzles, heat exchangers, and pumps where fluid flows continuously.

Isolated Systems I

- An isolated system is a system that is completely uninfluenced by its surroundings.
- No mass is permitted to cross the system boundary (mass flow in/out $= 0$).
- No energy (neither heat nor work) is permitted to cross the system boundary ($Q = 0$, $W = 0$).
- Example: A perfectly insulated, rigid tank containing a substance.
- The Universe (system + surroundings) is considered a closed isolated system in global thermodynamic analysis.

Summary of System Types I

- Closed System (Control Mass): Mass is fixed (no flow), Energy can cross boundary. Volume can vary.
- Open System (Control Volume): Mass can flow across boundary, Energy can cross boundary. Volume is usually fixed in space.
- Isolated System: Neither mass nor energy can cross the boundary.
- Selecting the correct boundary and system type is the critical first step in solving any thermodynamics engineering problem.

Energy, Heat, Work, Power and Exergy I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 1: Basic Concepts of Thermodynamics
- Lecture 5: Energy, Heat, Work, Power and Exergy

Agenda I

- 1. Thermodynamic Definitions of Energy & Forms of Energy
- 2. Heat Transfer and Sign Conventions
- 3. Work Transfer and Path Functions
- 4. Mechanical Forms of Work & P-V Diagram
- 5. Power and Rate of Energy Transfer
- 6. Exergy (Availability): Concept and Definition
- 7. Exergy Balance and Irreversibility

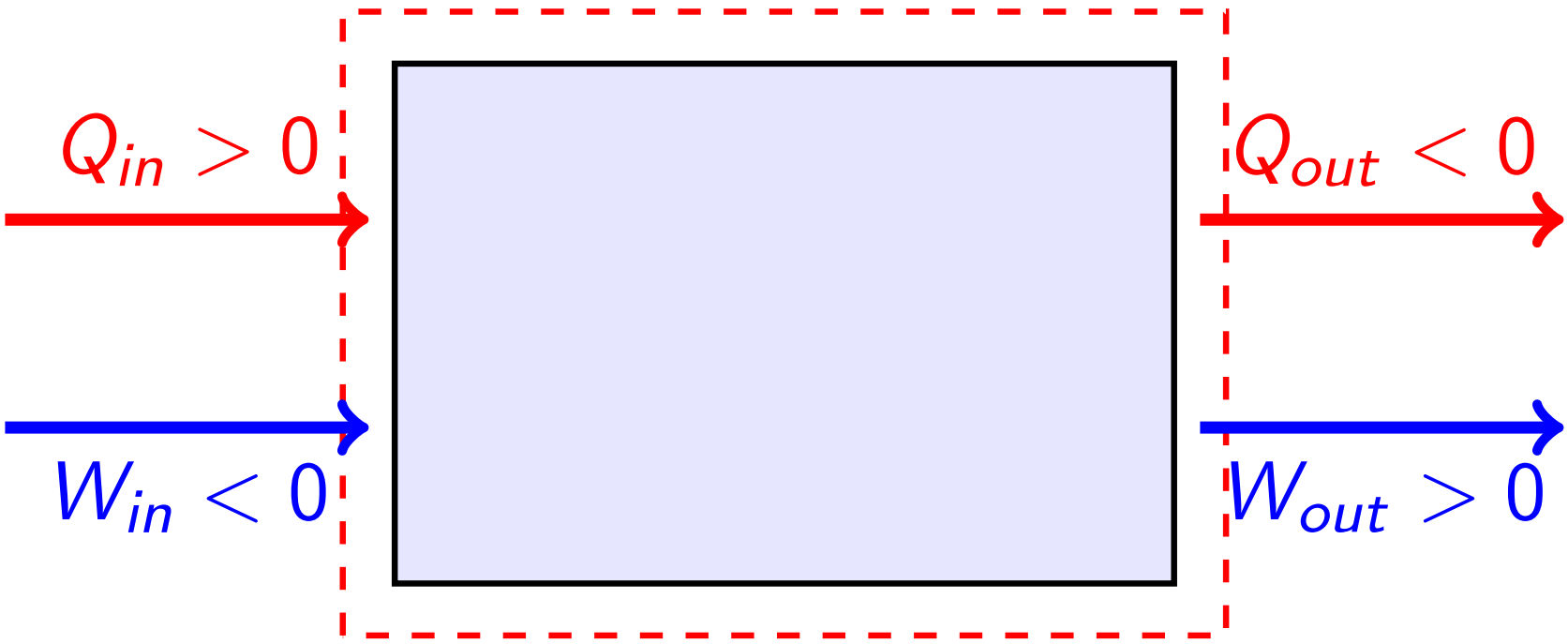
Forms of Energy: Macroscopic vs. Microscopic I

- Energy can exist in numerous forms: thermal, mechanical, kinetic, potential, electric, magnetic, chemical, and nuclear.
- Macroscopic forms of energy are those a system possesses as a whole with respect to some outside reference frame, such as kinetic ($E_k = \frac{1}{2}mV^2$) and potential ($E_p = mgz$) energies.
- Microscopic forms of energy are those related to the molecular structure of a system and the degree of the molecular activity, independent of outside reference frames. The sum of all microscopic forms of energy is called internal energy (U).
- Total energy of a system: $E = U + KE + PE = U + \frac{1}{2}mV^2 + mgz$. On a unit mass basis: $e = u + \frac{1}{2}V^2 + gz$.

Heat: Energy in Transition I

- Heat is defined as the form of energy that is transferred between two systems (or a system and its surroundings) by virtue of a temperature difference.
- Heat is energy in transition; it is recognized only as it crosses the system boundary. A system contains energy, but not heat.
- Heat transfer is a path function, meaning its value depends on the path followed during a process, not just the initial and final states. Its differential is inexact, denoted by δQ.
- Sign convention: Heat transfer to a system is positive ($Q > 0$), and heat transfer from a system is negative ($Q < 0$). An adiabatic process is one in which there is no heat transfer ($Q = 0$).

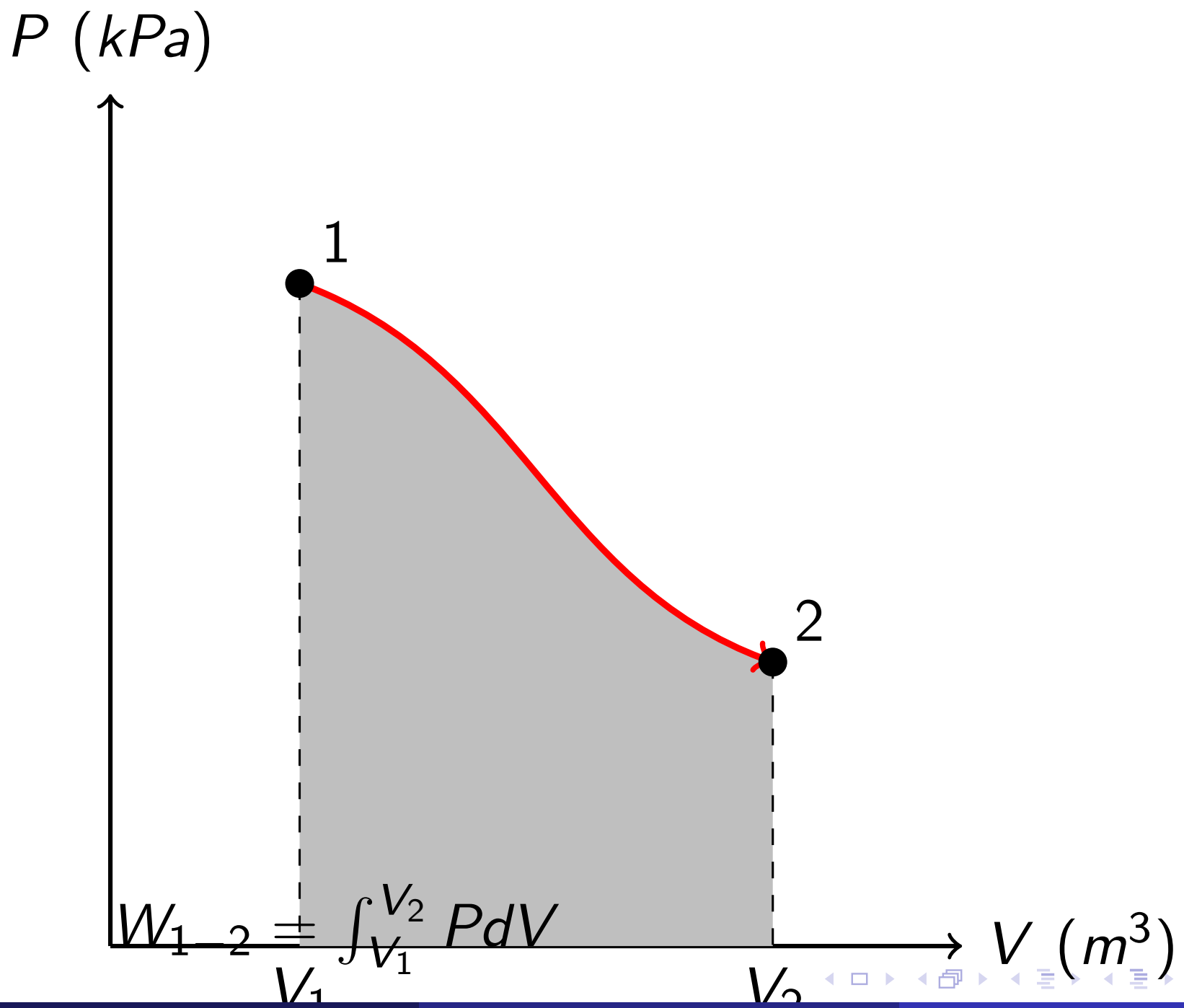
Thermodynamic Sign Conventions for Heat and Work



Work Transfer: Characteristics & Math representation I

- Work is the energy transfer associated with a force acting through a distance. In thermodynamics, work is defined as energy transfer across the system boundary not caused by a temperature difference.
- Like heat, work is a path function, not a state function. The work done depends on the process path between states 1 and 2: $W_{1-2} = \int_1^2 \delta W \neq W_2 - W_1$.
- Since work is a path function, its differential is inexact: δW. It cannot be expressed as a difference of system properties ($W_2 - W_1$ is mathematically incorrect; we write W_{1-2} or W instead).
- A system contains energy (internal, kinetic, potential) but it does not contain work. Work is an energy interaction that occurs only at system boundaries.

Mechanical Work: Boundary Work & P-V Diagram



Power: The Rate of Doing Work I

- Power is defined as the time rate of doing work or transferring energy. It is denoted by $\dot{W}$ (read as 'W dot').
- The unit of power is the Watt (1 W = 1 J/s). In engineering practice, kilowatts (kW), megawatts (MW), or horsepower (hp) are commonly used (1 hp $\approx$ 746 W).
- Mathematical definition of average power: $P_{avg} = \frac{\Delta W}{\Delta t}$. The instantaneous power is: $\dot{W} = \frac{\delta W}{dt}$.
- For mechanical work involving linear motion (force F and velocity V), power is given by: $\dot{W} = \vec{F} \cdot \vec{V}$. For rotational motion (torque T and angular velocity ω): $\dot{W} = T\omega$.

Concept of Exergy (Availability) I

- Exergy (also called availability) is the maximum useful work potential that can be obtained from a system as it undergoes a process to a specified dead state.
- Unlike energy, which is conserved according to the First Law of Thermodynamics, exergy is NOT conserved. It is destroyed during real, irreversible processes.
- The dead state represents a state of thermodynamic equilibrium between the system and its surroundings (usually at temperature $T_0 = 298.15$ K and pressure $P_0 = 101.3$ kPa). At the dead state, the exergy of the system is zero.
- Exergy represents the quality of energy. Higher-temperature thermal energy has a higher exergy content (greater potential to do work) than the same amount of low-temperature thermal energy.

Exergy vs. Energy & Irreversibility I

- Energy is a measure of quantity, whereas Exergy is a measure of quality or usefulness of that energy.
- Exergy destroyed ($X_{destroyed}$) is directly proportional to the entropy generated (S_{gen}) during a process, expressed by the Gouy-Stodola theorem: $X_{destroyed} = T_0 S_{gen}$.
- The exergy of a closed system at a state defined by U, V, and S relative to the dead state (U_0, V_0, S_0) is:
  $\Phi = (U - U_0) + P_0(V - V_0) - T_0(S - S_0) + KE + PE$.
- By analyzing exergy destructions (irreversibilities), mechanical engineers can pinpoint locations of thermodynamic inefficiency in power plants, engines, and refrigeration systems.

Thermodynamic properties and their units. I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 1: Basic Concepts of Thermodynamics
- Lecture 6: Thermodynamic Properties and Their Units

Agenda I

- Introduction to Thermodynamic Properties (Intensive vs. Extensive)
- The State Postulate & Independent Properties
- Mass-Volume Relations: Density, Specific Volume, and Specific Gravity
- Temperature: Zeroth Law and Temperature Scales
- Pressure: Absolute, Gauge, and Vacuum Pressures
- Pressure Measurement: U-Tube Manometer

Thermodynamic Properties: Definition and Classification I

- A property is any characteristic of a system in thermodynamic equilibrium that can be assigned a quantitative value without knowledge of the system's history.
- Properties must be macroscopically observable and independent of the path taken to reach the state (they are state functions, mathematical point functions).
- Intensive properties are independent of the mass or size of the system. Examples: Temperature (T), Pressure (P), and Density (ρ).
- Extensive properties depend directly on the size or extent of the system. Examples: Total Mass (m), Total Volume (V), and Total Internal Energy (U).

Thermodynamic Properties: Definition and Classification II

- Specific properties are extensive properties per unit mass, making them intensive. Examples: Specific volume ($v = V/m$ in m^3/kg) and Specific enthalpy ($h = H/m$ in kJ/kg).

Visualizing Intensive vs. Extensive Properties

$$mVTPm/2V/2TPm/2V/2TP$$

The State Postulate I

- The thermodynamic state of a simple compressible system is completely specified by two independent, intensive properties.
- A system is simple compressible in the absence of electrical, magnetic, gravitational, motion, and surface tension effects.
- If these effects are significant, an additional independent property must be specified for each effect.
- Independent properties: Two properties are independent if one can be varied while the other is held constant. Temperature (T) and specific volume (v) are always independent in single-phase regions.
- Crucial Exception: During phase change (e.g., boiling water), temperature and pressure are dependent properties (T_{sat} and P_{sat}) and cannot define the state alone.

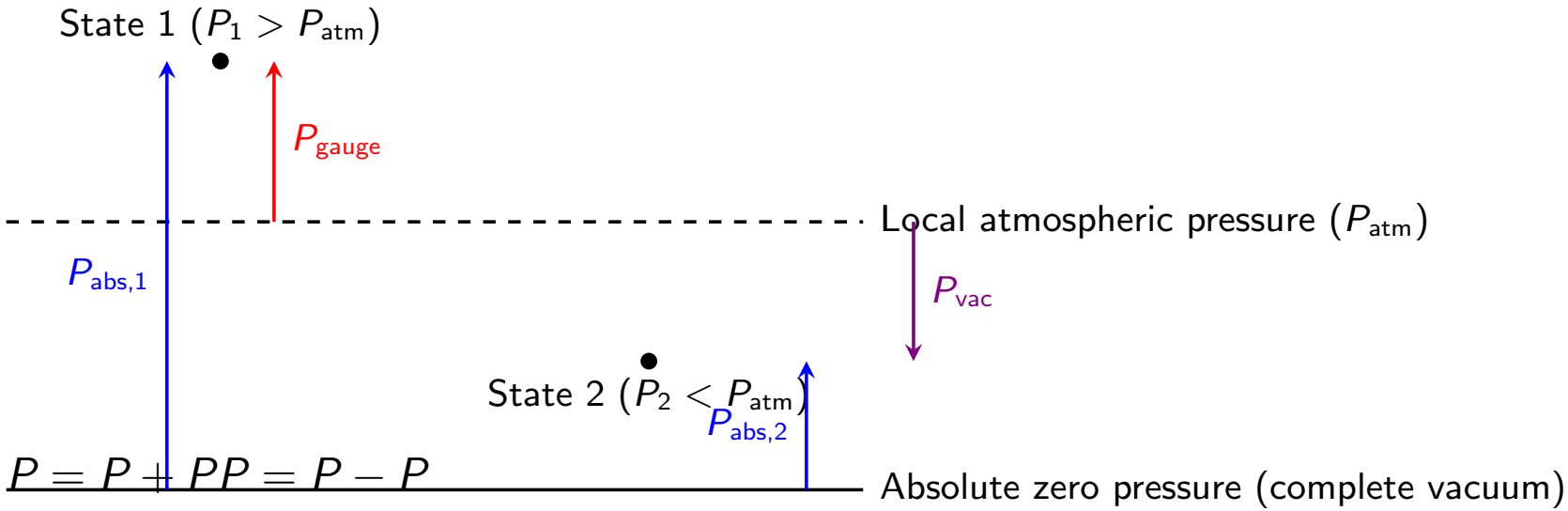
Mass-Volume Relations: Density & Specific Volume I

- Density (ρ) is defined as mass per unit volume: $\rho = m/V$ (SI unit: kg/m^3, English: lbm/ft^3).
- Specific Volume (v) is the reciprocal of density, representing the volume occupied by a unit mass: $v = V/m = 1/\rho$ (SI unit: m^3/kg, English: ft^3/lbm).
- Specific Gravity (SG) is the ratio of a substance's density to the density of water at 4°C ($\rho_{H_2O} = 1000$ kg/m^3): $SG = \rho/\rho_{H_2O}$.
- Specific Weight (γ) is the weight of a unit volume of a substance: $\gamma = \rho g$ (SI unit: N/m^3, English: lbf/ft^3), where g is local gravitational acceleration.

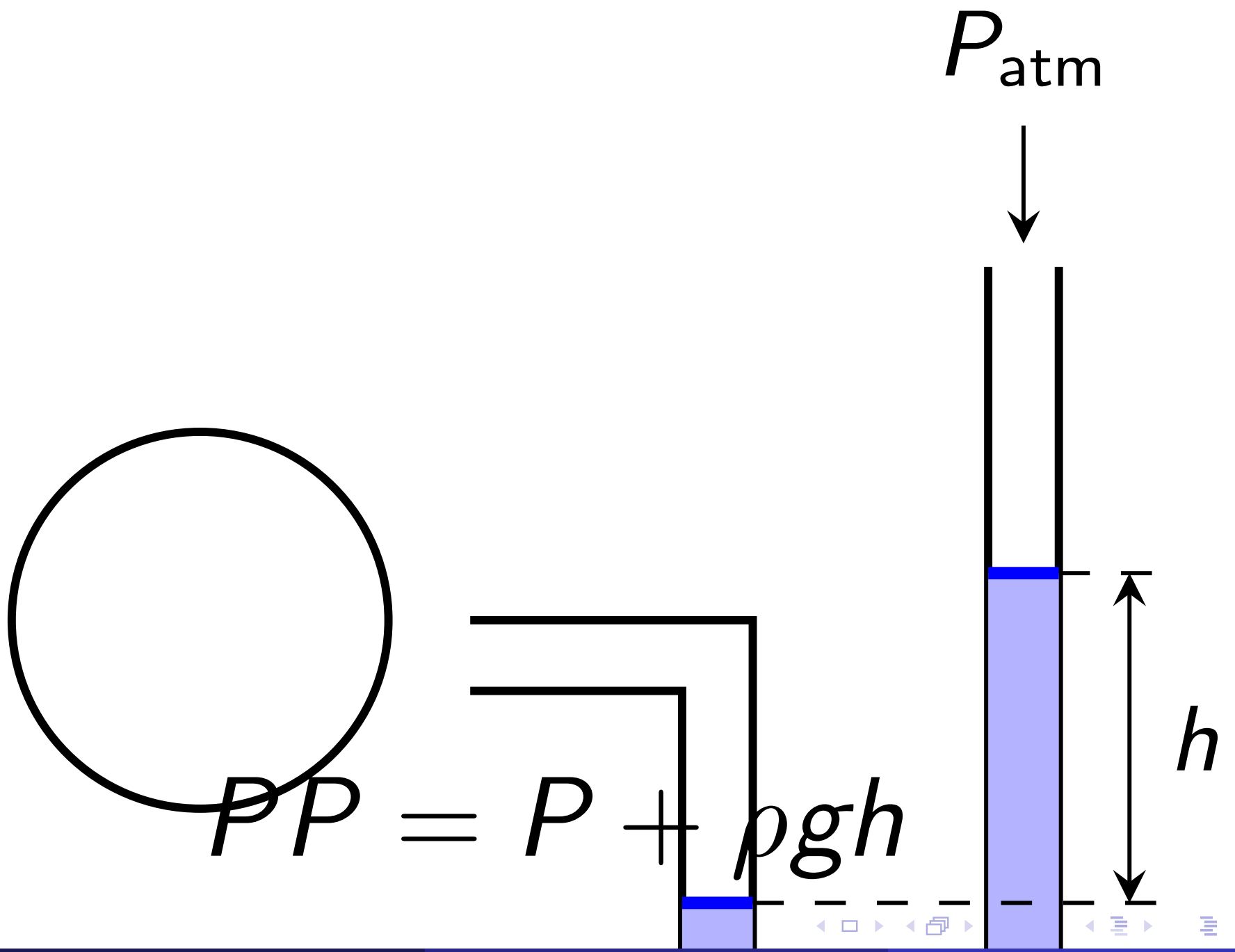
Temperature and the Zeroth Law of Thermodynamics I

- Temperature is a measure of the thermal energy (molecular translational kinetic energy) of a system.
- The Zeroth Law of Thermodynamics: If two bodies are in thermal equilibrium with a third body, they are in thermal equilibrium with each other.
- This law serves as the physical basis for temperature measurement, validating the use of thermometers.
- Thermodynamic temperature scale (absolute scale) is independent of the properties of any particular substance.
- Scale Relations: $T(\mathrm{K}) = T(^\circ\mathrm{C}) + 273.15$, $T(\mathrm{R}) = T(^\circ\mathrm{F}) + 459.67$, $T(\mathrm{R}) = 1.8\,T(\mathrm{K})$, and $T(^\circ\mathrm{F}) = 1.8\,T(^\circ\mathrm{C}) + 32$.

Thermodynamic Pressure Definitions



Pressure Measurement: The U-Tube Manometer



Summary of Thermodynamic Properties and Units I

- Pressure (P): SI units: Pascal ($1\ \text{Pa} = 1\ \text{N/m}^2$), bar ($1\ \text{bar} = 10^5\ \text{Pa}$), standard atmosphere ($1\ \text{atm} = 101,325\ \text{Pa}$). English units: lbf/in^2 (psi), lbf/ft^2 (psf).
- Temperature (T): SI units: Kelvin (K), Degree Celsius (°C). English units: Rankine (R), Degree Fahrenheit (°F).
- Volume (V) and Specific Volume (v): Volume in m^3 or L. Specific volume in m^3/kg or ft^3/lbm.
- Energy (E, U, H), Enthalpy, Entropy: Energy in Joule (J) or BTU. Specific energy in kJ/kg or Btu/lbm. Specific entropy in $kJ/kg \cdot K$ or $Btu/lbm \cdot R$.
- Note on calculations: Always use absolute values for temperature (T in K or R) and pressure (P in Pa or psi) in thermodynamic state relations and ideal gas laws.

Simple numerical based on above. I

- Course: Engineering Thermodynamics (DI03019011)
- Unit: 1 - Basic Concepts of Thermodynamics
- Lecture 7: Numerical Problems on Temperature Scales, Pressure, and Work
- Instructor: Professor of Mechanical Engineering
- Prerequisites: Zeroth Law, Thermometric Properties, Pressure Measurement, Quasi-static Work

Lecture Agenda I

- 1. Numerical on Temperature Scale Calibration (Zeroth Law application)
- 2. Numerical on Pressure Measurement using U-tube Manometer (Gauge vs Absolute Pressure)
- 3. Numerical on Quasi-Static Work Done during Polytropic Gas Expansion ($PV^n = C$)
- 4. Review of key thermodynamic equations and units conversion

Problem 1: Temperature Scale Calibration I

- Problem Statement: A new temperature scale °N is proposed. On this scale, the ice point (0 °C) is defined as -20 °N and the steam point (100 °C) is defined as 180 °N.
- Derive a linear relationship between the Celsius scale (t_C) and the new scale (t_N): $t_N = a \cdot t_C + b$.
- Using this relationship, determine the temperature in °N corresponding to a Celsius temperature of 37 °C (human body temperature).
- Also, find the temperature at which both scales read the same value.

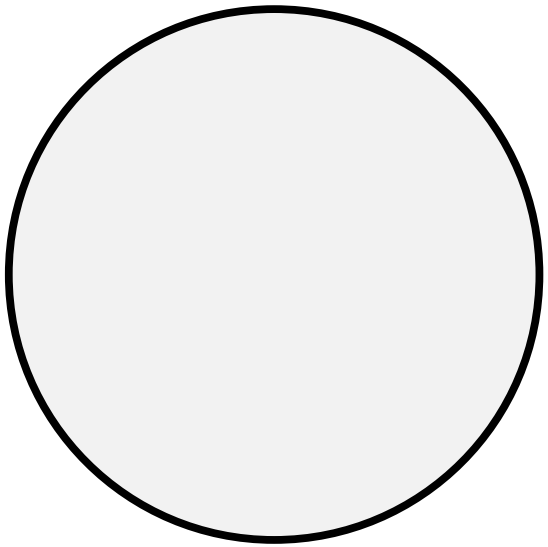
Solution to Problem 1: Calibration & Conversion I

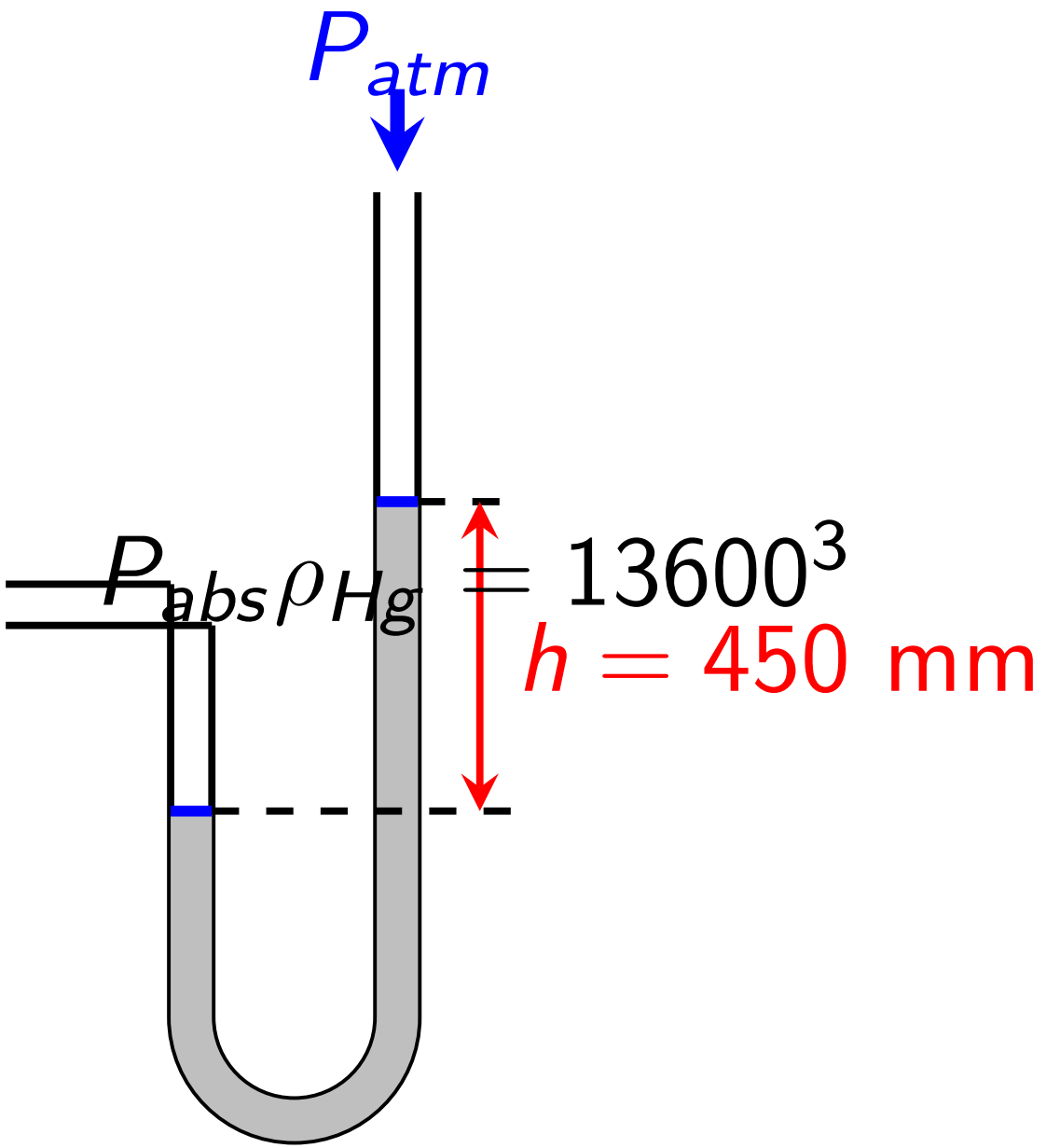
- Step 1: Set up the linear equation: $t_N = a \cdot t_C + b$
- Step 2: Apply ice point boundary condition: At $t_C = 0$ °C, $t_N = -20$ °N $\implies -20 = a(0) + b \implies b = -20$.
- Step 3: Apply steam point boundary condition: At $t_C = 100$ °C, $t_N = 180$ °N
  $\implies 180 = a(100) - 20 \implies 100a = 200 \implies a = 2$.
- Step 4: The conversion formula is: $t_N = 2t_C - 20$.
- Step 5: For $t_C = 37$ °C: $t_N = 2(37) - 20 = 74 - 20 = 54$ °N.
- Step 6: To find where they are equal, set $t_N = t_C = x$:
  $x = 2x - 20 \implies x = 20$. Both scales read the same value at 20 °C and 20 °N.

Problem 2: U-Tube Manometer Analysis I

- Problem Statement: A U-tube manometer containing mercury (density $\rho_m = 13600$ kg/m^3) is connected to a gas tank. The other end is open to the atmosphere.
- The atmospheric pressure is measured to be 101.3 kPa using a barometer.
- The difference in mercury levels between the two columns is $h = 450$ mm. The column connected to the gas tank has the lower mercury level.
- Calculate: (a) The gauge pressure of the gas in the tank in kPa.
- (b) The absolute pressure of the gas in the tank in kPa and bar.
- Assume local acceleration due to gravity $g = 9.81$ m/s^2.

Diagram: U-Tube Manometer Setup





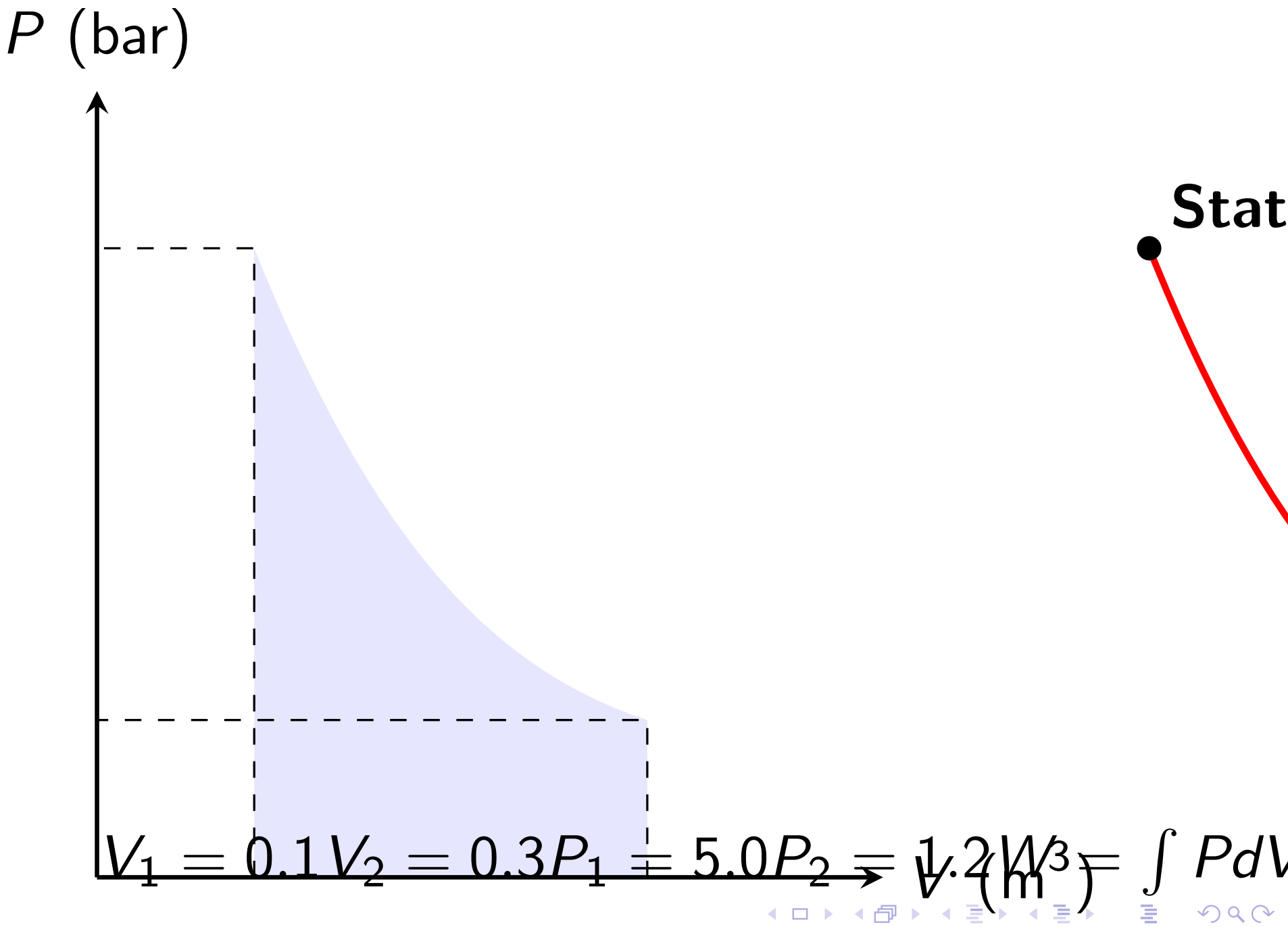
Solution to Problem 2: Manometer Equations I

- Step 1: Identify that the left column (connected to the tank) is lower, indicating that the tank pressure is higher than atmospheric pressure (gauge pressure is positive).
- Step 2: Formula for gauge pressure: $P_{gauge} = \rho_m \cdot g \cdot h$
- Step 3: Substitute the given values (converting h to meters): $P_{gauge} = 13600 \text{ kg/m}^3 \times 9.81 \text{ m/s}^2 \times 0.450 \text{ m}$
- Step 4: Calculate the gauge pressure: $P_{gauge} = 60037.2 \text{ N/m}^2 = 60.04 \text{ kPa}$.
- Step 5: Calculate absolute pressure: $P_{abs} = P_{atm} + P_{gauge} = 101.3 \text{ kPa} + 60.04 \text{ kPa} = 161.34 \text{ kPa}$.
- Step 6: Convert to bar: $P_{abs} = 161.34 \times 10^3 \text{ Pa} = 1.6134 \text{ bar}$ (since 1 bar $= 10^5$ Pa).

Problem 3: Polytropic Gas Expansion Work I

- Problem Statement: Air is contained in a piston-cylinder assembly and undergoes a quasi-static polytropic expansion process.
- State 1: $P_1 = 5.0$ bar $= 500$ kPa, $V_1 = 0.1$ m^3
- State 2: $P_2 = 1.2$ bar $= 120$ kPa, $V_2 = 0.3$ m^3
- The process path is governed by the polytropic relation: $P \cdot V^n = C$, where n is the polytropic index.
- Determine: (a) The polytropic index n.
- (b) The work done by the air during this expansion process in kJ.

Diagram: P-V Plot of Polytropic Expansion



Solution to Problem 3: Index and Work I

- Step 1: Write the polytropic relation for the two states: $P_1 V_1^n = P_2 V_2^n \implies \frac{P_1}{P_2} = \left(\frac{V_2}{V_1}\right)^n$.
- Step 2: Take natural logarithms on both sides: $\ln\left(\frac{P_1}{P_2}\right) = n \cdot \ln\left(\frac{V_2}{V_1}\right)$.
- Step 3: Calculate the index n: $n = \frac{\ln(5.0/1.2)}{\ln(0.3/0.1)} = \frac{\ln(4.167)}{\ln(3)} = \frac{1.4271}{1.0986} \approx 1.30$.
- Step 4: The work done formula for a polytropic process ($n \neq 1$) is: $W_{1-2} = \frac{P_1 V_1 - P_2 V_2}{n-1}$.
- Step 5: Convert pressures to kPa (500 kPa and 120 kPa) to get work directly in kJ: $W_{1-2} = \frac{(500 \times 0.1) - (120 \times 0.3)}{1.30-1}$.
- Step 6: Compute the final value: $W_{1-2} = \frac{50-36}{0.30} = \frac{14}{0.30} = 46.67$ kJ. Since $W_{1-2} > 0$, work is done by the system.

Thermodynamic equilibrium. I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 1: Basic Concepts of Thermodynamics
- Lecture 8: Thermodynamic Equilibrium
- Instructor: Professor of Mechanical Engineering

Agenda I

- Introduction to Thermodynamic Equilibrium: Definition & Criteria
- Thermal Equilibrium & The Zeroth Law of Thermodynamics
- Mechanical Equilibrium & Pressure Uniformity
- Phase Equilibrium & Chemical Potential Equality
- Chemical Equilibrium & Gibbs Free Energy Minimization
- The State Postulate & Thermodynamic State Description

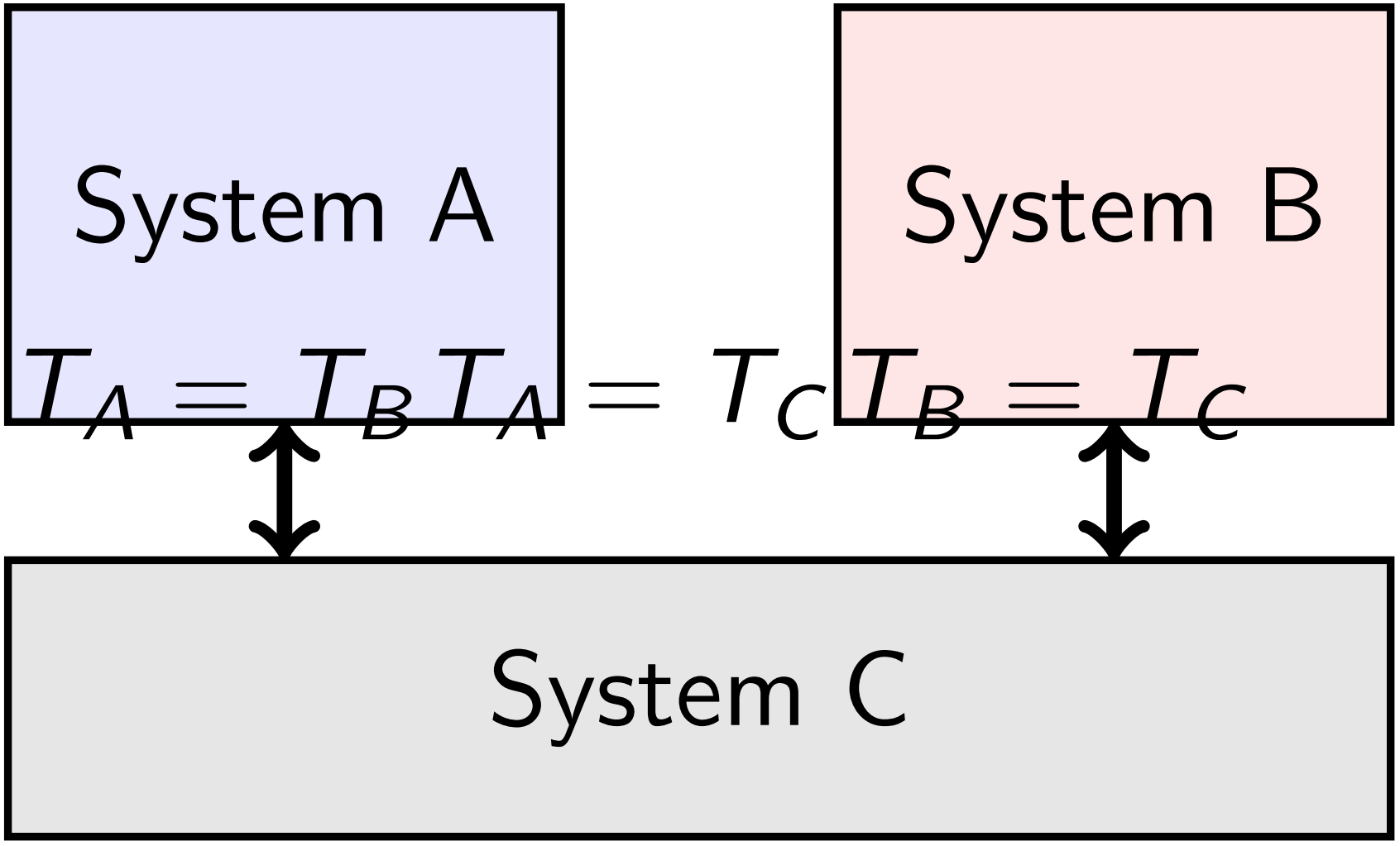
Introduction to Thermodynamic Equilibrium I

- A system is in thermodynamic equilibrium if its intensive and extensive properties do not change with time when isolated from its surroundings.
- No unbalanced potentials (driving forces) exist within the system; there are no spatial gradients of temperature, pressure, or chemical potential.
- To achieve complete thermodynamic equilibrium, a system must satisfy four conditions simultaneously: thermal, mechanical, phase, and chemical equilibrium.
- If any of these conditions are violated, spontaneous internal processes will occur to drive the system toward equilibrium, producing entropy.

Thermal Equilibrium & The Zeroth Law I

- Thermal equilibrium is reached when the temperature is uniform throughout the system and equal to the temperature of its surroundings.
- Under thermal equilibrium, the temperature gradient is zero ($dT/dx = 0$), and there is no net heat transfer within the system or across its boundary.
- The Zeroth Law of Thermodynamics: If two bodies are in thermal equilibrium with a third body, they are in thermal equilibrium with each other.
- This law validates the concept of temperature measurement, allowing the third body to act as a calibrated thermometer.

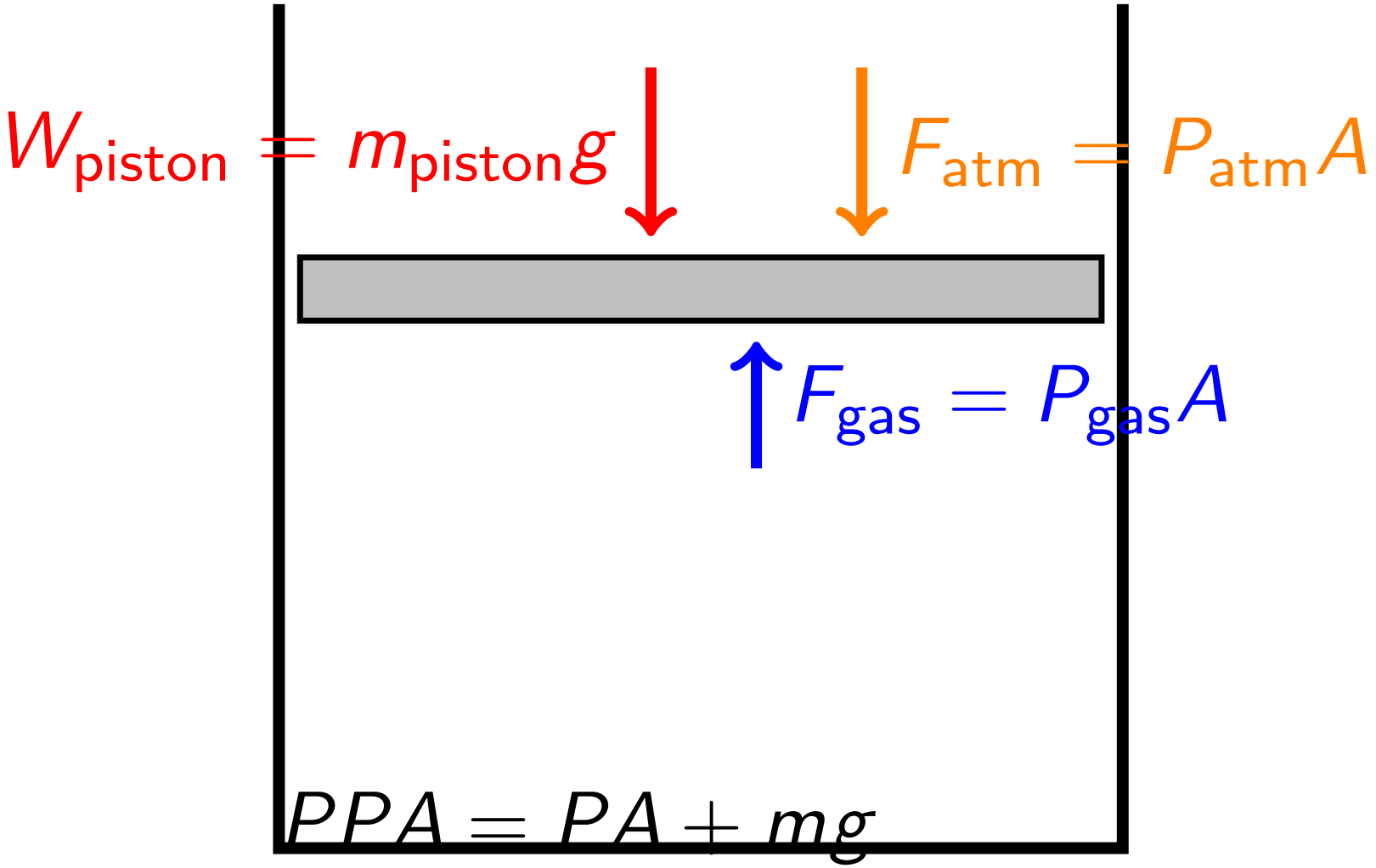
Thermal Equilibrium & Zeroth Law Illustration



Mechanical Equilibrium & Pressure Uniformity I

- Mechanical equilibrium requires that there are no unbalanced forces acting within the system or between the system and its surroundings.
- The pressure throughout the system must be spatially uniform, except for negligible static pressure gradients due to gravity in small-scale systems.
- If a pressure gradient exists, fluid acceleration and shear stresses will occur to redistribute mass and equalize pressure.
- For a simple piston-cylinder device, mechanical equilibrium is achieved when the internal gas pressure balances the external forces and the piston weight.

Force Balance in Mechanical Equilibrium



Phase Equilibrium & Chemical Potential I

- Phase equilibrium occurs when a multi-phase system experiences no net mass transfer of any chemical species between its coexisting phases.
- For a system containing multiple phases of a pure substance (e.g., liquid and vapor), the mass of each phase remains constant over time.
- Thermodynamically, phase equilibrium is defined by the equality of the chemical potential of each species in all phases: $\mu\_i\hat{}\alpha = \mu\_i\hat{}\beta$.
- Any gradient in chemical potential ($\Delta\mu$) acts as a driving force for mass transfer (evaporation, condensation, or diffusion).

Chemical Equilibrium & Reaction Affinity I

- Chemical equilibrium is attained when the chemical composition of a system remains constant over time and no net chemical reactions take place.
- At this state, the rate of the forward chemical reaction equals the rate of the reverse reaction, yielding a net reaction rate of zero.
- Thermodynamically, a system at constant temperature and pressure reaches chemical equilibrium when its Gibbs free energy is at a minimum ($dG = 0$).
- The chemical affinity of all possible reactions is zero, indicating that there is no remaining thermodynamic affinity or driving force for reaction.

Summary & The State Postulate I

- Thermodynamic equilibrium is a restrictive state requiring simultaneous thermal, mechanical, phase, and chemical equilibrium.
- The State Postulate: The thermodynamic state of a simple compressible system is completely specified by two independent, intensive properties.
- A system must be in thermodynamic equilibrium for its properties to be macroscopically defined and for the State Postulate to apply.
- Actual engineering processes are modeled as quasi-equilibrium (or quasi-static) processes, which deviate only infinitesimally from equilibrium.

Zeroth law of thermodynamics and its application. I

- Course: Engineering Thermodynamics (DI03019011)
- Lecture 9: Zeroth Law of Thermodynamics and its Application
- Instructor: Professor of Mechanical Engineering
- Unit 1: Basic Concepts of Thermodynamics

Lecture Agenda I

- 1. The Concept of Thermal Equilibrium
- 2. Statement and Origin of the Zeroth Law of Thermodynamics
- 3. Physical Significance: The Concept of Temperature
- 4. Thermometric Properties and Thermometric Substances
- 5. Temperature Scales (Celsius, Kelvin, Fahrenheit)
- 6. Reference Points and the Triple Point of Water
- 7. Constant-Volume and Constant-Pressure Gas Thermometers
- 8. Practical Applications and Summary

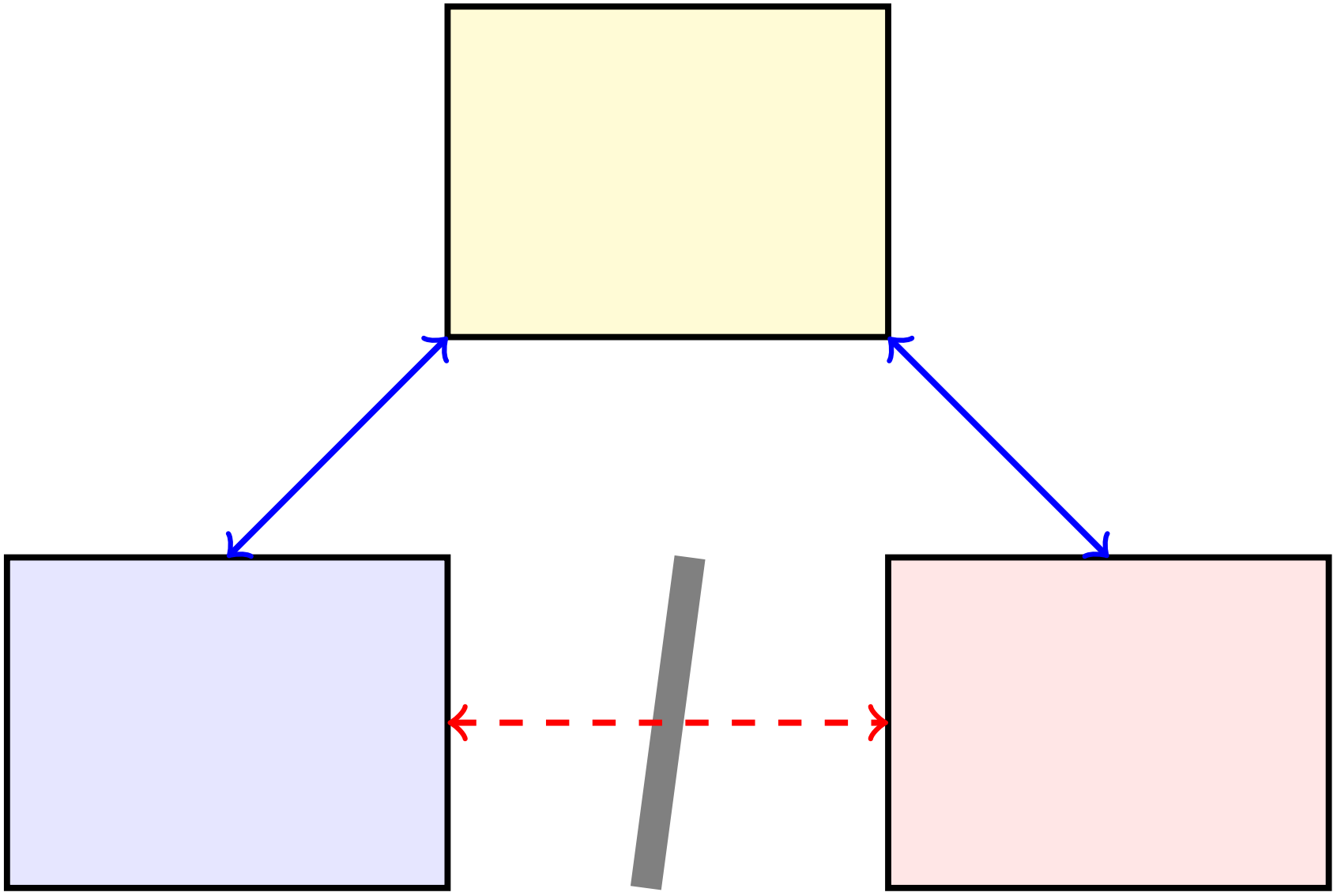
Understanding Thermal Equilibrium I

- Thermal equilibrium is a state in which there is no net flow of thermal energy (heat) between two systems when they are connected by a path permeable to heat.
- If two systems are in thermal equilibrium, they share a common thermodynamic state variable called temperature.
- Diathermic Wall: A boundary that allows heat transfer (e.g., thin copper plate).
- Adiabatic Wall: A boundary that prevents heat transfer (e.g., thick fiberglass or vacuum insulation).
- When two bodies at different temperatures are placed in contact via a diathermic wall, heat flows from the hotter to the colder body until they reach thermal equilibrium.

The Zeroth Law of Thermodynamics I

- Formal Statement: If two thermodynamic systems are each in thermal equilibrium with a third system, then they are in thermal equilibrium with each other.
- Historical Context: Formulated by Ralph H. Fowler in 1931, long after the First and Second Laws were established.
- Why 'Zeroth'?: It is more fundamental than the First and Second Laws because it defines the concept of temperature, which is assumed by the other laws.
- Mathematical logic: If System A is in thermal equilibrium with System C ($T_A = T_C$), and System B is in thermal equilibrium with System C ($T_B = T_C$), then System A is in thermal equilibrium with System B ($T_A = T_B$).
- This law provides the logical basis for the design and validation of all thermometers.

Schematic Representation of the Zeroth Law



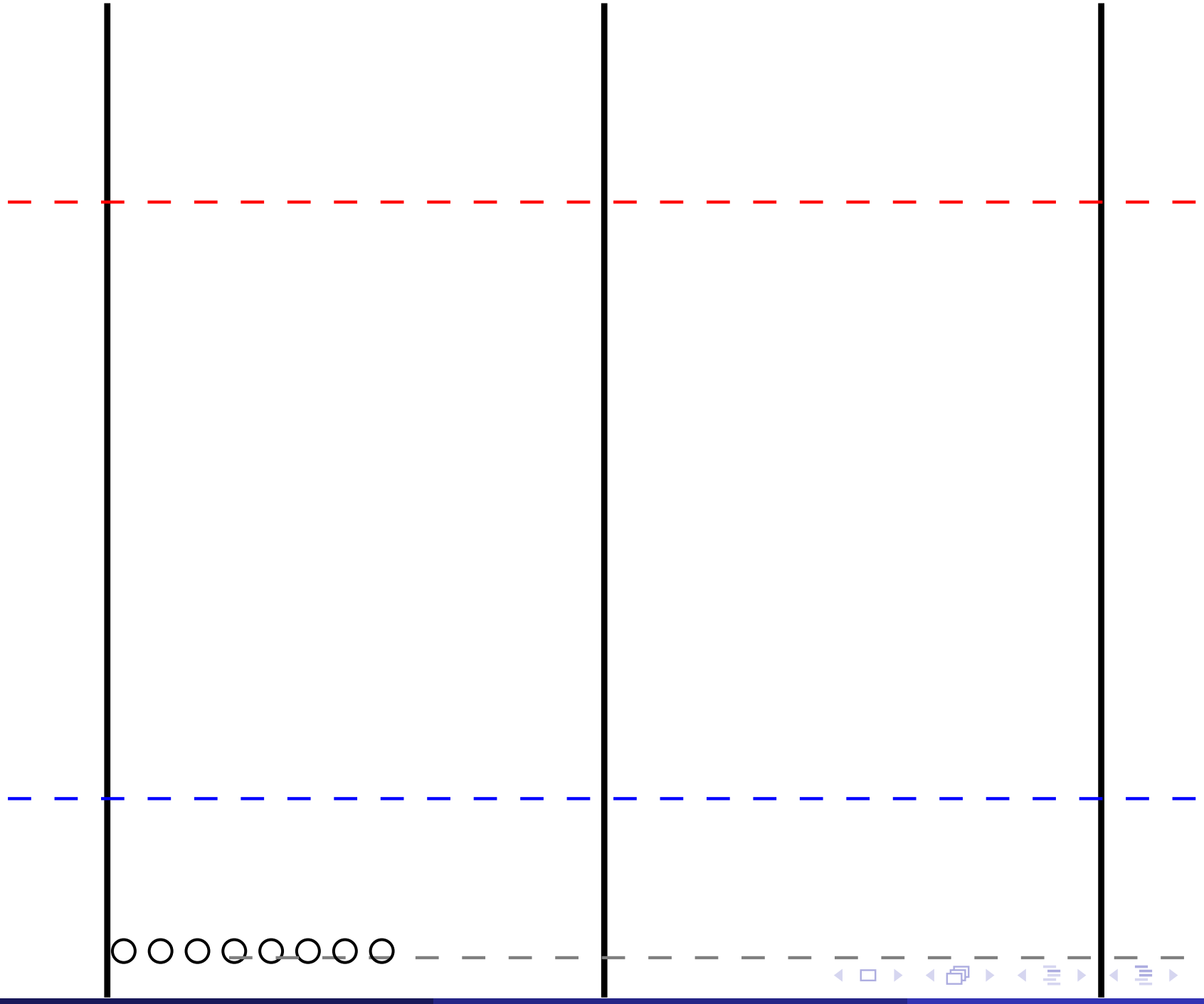
Concept of Temperature & Thermometric Properties I

- Temperature is a scalar measure of the thermal state of a system, representing its 'hotness' or 'coldness' and determining the direction of heat transfer.
- To measure temperature, we utilize a 'Thermometric Substance' and its 'Thermometric Property' (X): a physical property that varies monotonically with temperature.
- Examples of Thermometric Substances and their Properties:
- 1. Liquid-in-glass thermometer: Volume of liquid (mercury/alcohol)
- 2. Constant-volume gas thermometer: Pressure of gas (P)
- 3. Constant-pressure gas thermometer: Volume of gas (V)
- 4. Electric resistance thermometer: Electrical resistance (R)
- 5. Thermocouple: Thermoelectric electromotive force (E)
- 6. Optical Pyrometer: Monochromatic radiation intensity

Temperature Scales and Calibration I

- Calibration requires establishing a relationship between temperature (T) and the thermometric property (X).
- Linear relationship assumption: T(X) = a * X (for single-point calibration) or T(X) = a * X + b (for two-point calibration).
- Traditional Two-Point Calibration: Uses two easily reproducible reference states under standard atmospheric pressure (1.01325 bar):
- - Ice Point (Freezing point of air-saturated water): 0°C / 32°F
- - Steam Point (Boiling point of pure water): 100°C / 212°F
- Modern Single-Point Calibration: Uses the Triple Point of water (where solid, liquid, and vapor phases coexist in equilibrium) defined as exactly 273.16 K (0.01°C) at 611.65 Pa.
- Formula with Triple Point: T = 273.16 * (X / X_tp), where X_tp is the property value at the triple point.

Comparison of Temperature Scales



Constant-Volume Gas Thermometer I

- A constant-volume gas thermometer is the standard instrument used to define the thermodynamic temperature scale and calibrate other thermometers.
- Operation: A bulb filled with a low-density gas (e.g., hydrogen, helium) is placed in the system. The volume of the gas is kept constant by adjusting the height of a mercury column (manometer) to a reference index.
- Thermometric Property: The pressure (P) of the gas, measured via the manometer height difference: P = P_atm + ρ * g * h.

Constant-Volume Gas Thermometer II

- Thermodynamic Temperature: Defined as T = 273.16 * $\lim\_P\_tp \rightarrow 0(P/P\_tp)where P\_tp is the pressure at the triple point$.
- By taking the limit as the pressure of the gas approaches zero, the gas behaves ideally, making the temperature measurement independent of the specific gas used.

Engineering Applications & Summary I

- 1. Standard Thermometry: Every temperature sensor (RTD, Thermocouple, Thermistor) operates by reaching thermal equilibrium with the target system and using the Zeroth Law to infer temperature.
- 2. Calibration of Industrial Sensors: High-precision baths are used as a reference system (System C) to calibrate industrial temperature sensors (System A) against a standard thermometer (System B).
- 3. HVAC and Climate Control: Environmental sensors monitor building spaces by assuming the sensor is in thermal equilibrium with the room air.
- 4. Medical Thermometers: Clinical thermometers require a certain time in contact with the body to ensure thermal equilibrium is established before reading.

Engineering Applications & Summary II

- Summary: Without the Zeroth Law, temperature could not be defined as a universal state variable, and temperature measurements would be logically invalid.

Heat and work relation with Joule's Experiment. I

- Course: Engineering Thermodynamics (DI03019011)
- Topic: 2.1 Heat and work relation with Joule's Experiment
- Unit 2: First Law of Thermodynamics
- Understanding the equivalence of mechanical work and thermal energy.

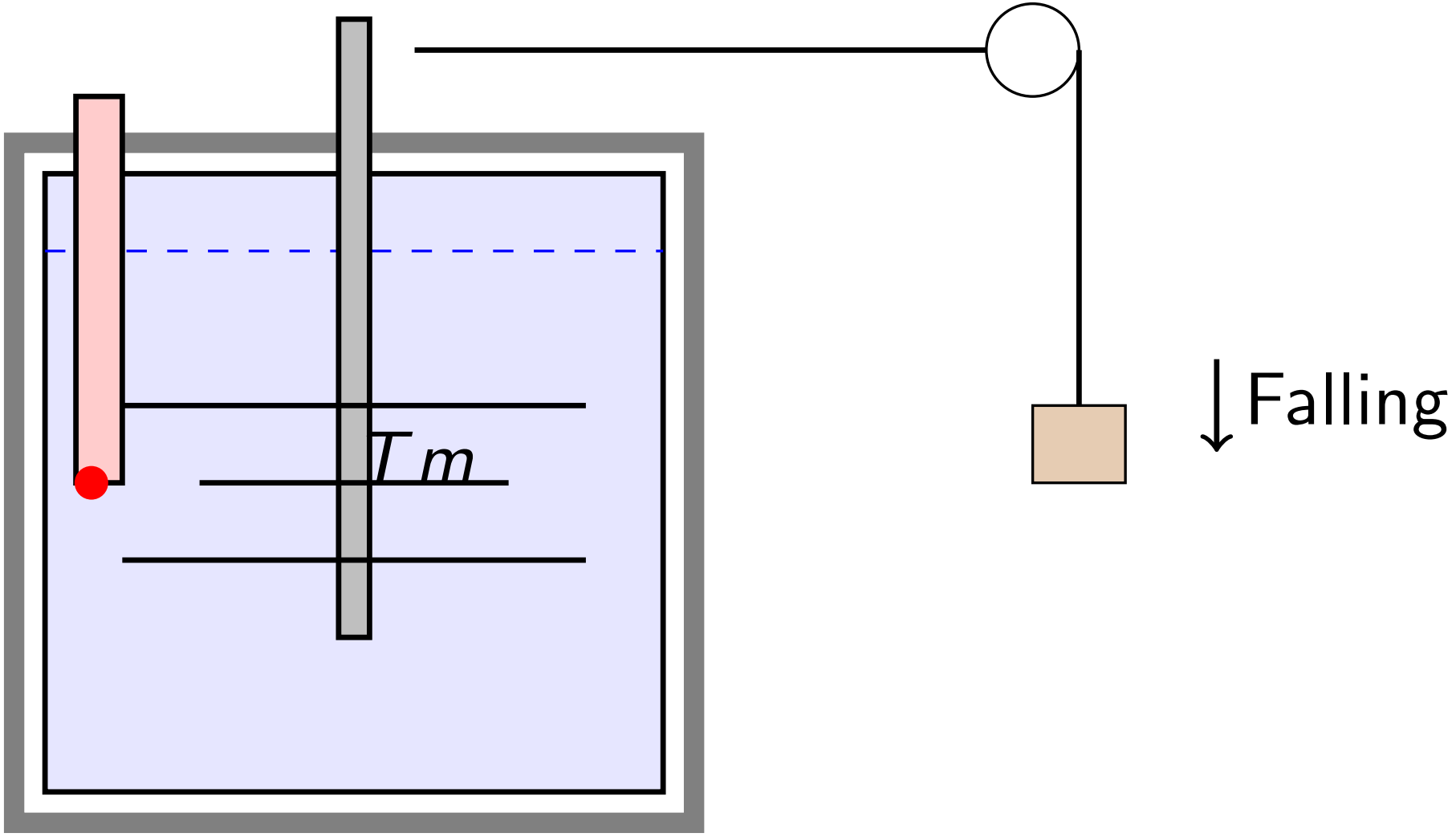
Lecture Agenda I

- Thermodynamic definitions of Heat (Q) and Work (W) as path functions.
- The historical shift from the Caloric theory of heat to the Mechanical theory.
- Joule's Paddle Wheel Experiment: Detailed setup and physical parameters.
- Experimental observations, calculations, and the Mechanical Equivalent of Heat (J).
- Formulating the First Law of Thermodynamics for closed systems undergoing a cycle.
- Sign conventions, mathematical expressions, and engineering implications.

Thermodynamic Heat and Work I

- Work (W) is an energy interaction where the sole effect on things external to the system can be reduced to the raising of a weight.
- Heat (Q) is defined as the form of energy transfer across the boundary of a system due to a temperature difference.
- Both heat and work are boundary phenomena; they are only recognized as they cross the boundary of the system.
- Both are path functions, meaning their values depend on the specific path/process followed, not just the initial and final states.
- Mathematically, they are inexact differentials represented by $\delta Q and \delta W$.

Joule's Paddle Wheel Experiment Setup



Joule's Experiment: Execution & Observations I

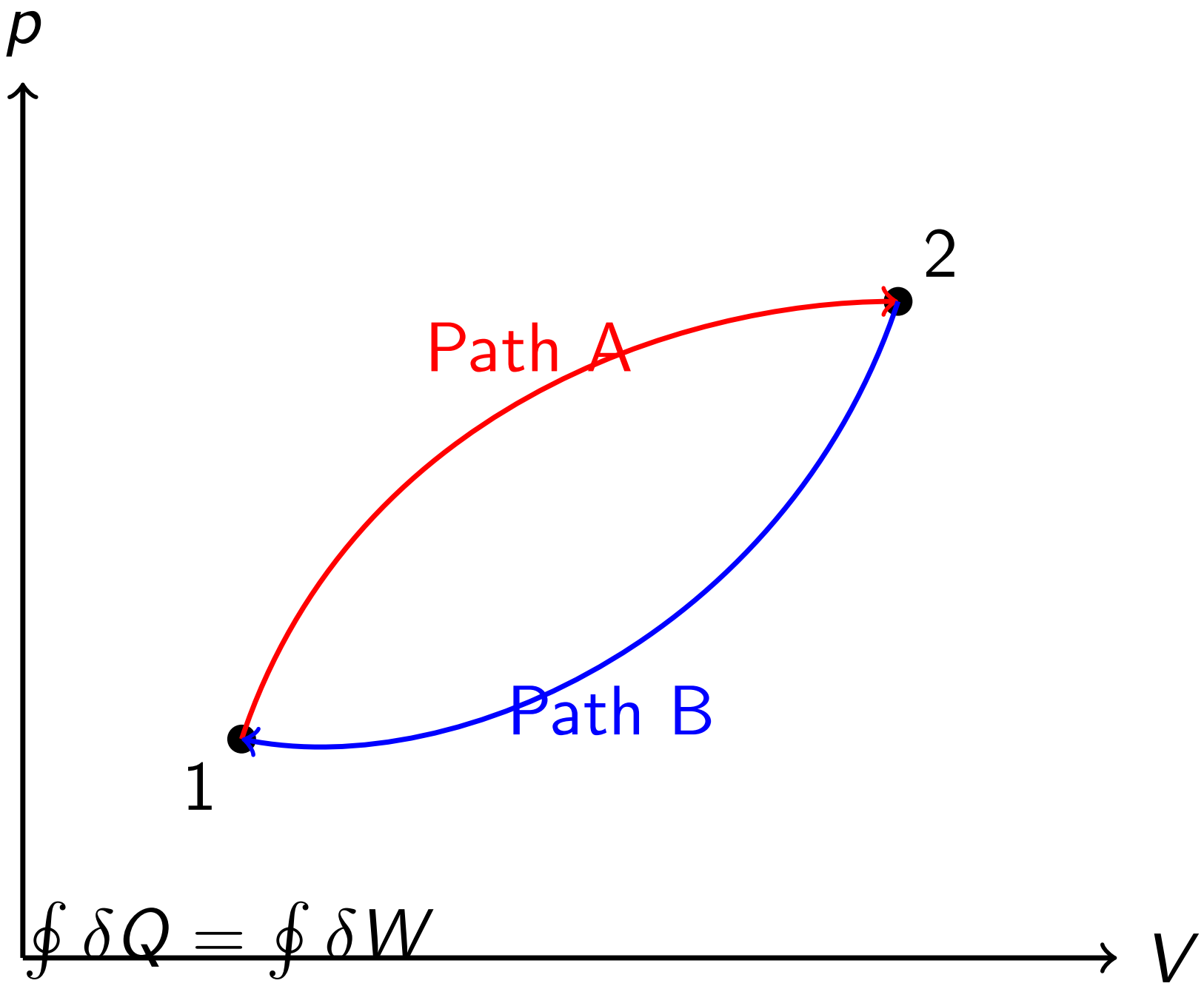
- A known mass (m) is allowed to fall slowly through a height (h), driving a paddle wheel immersed in water.
- The loss of potential energy of the mass represents the mechanical work input: $W = mgh$.
- The paddle wheel stirs the water, converting the work input into random kinetic energy via fluid friction (viscosity).
- The thermometer registers a rise in temperature from T1 to T2, indicating a change in the state of the fluid.
- The thermal insulation is then removed, and the fluid is cooled back to T1 by bringing it in contact with a cold water bath.
- The heat transfer (Q) from the system to the bath is measured using standard calorimetry techniques.

The Mechanical Equivalent of Heat (J) I

- Joule's repeated experiments showed that a fixed amount of work input (W) always corresponds to a proportional amount of heat output (Q).
- Mathematically, this relationship is expressed as: $W \propto Q \implies W = J \cdot Q$.
- The constant of proportionality, J, is called the Mechanical Equivalent of Heat.
- In traditional units: $J \approx 4.186$ Joules per Calorie($or 778$ ft$\cdot$ lbf/Btu).

- In modern SI units, both heat and work are measured in Joules (J), which yields $J = 1.0$.
- This experimental proof disproved the Caloric theory and established heat as energy in transition.

First Law for a Cyclic Process



Mathematical Formulation of the First Law I

- For a closed system undergoing a thermodynamic cycle, the cyclic integral of heat is equal to the cyclic integral of work.
- Expressed mathematically as:
  $\oint \delta Q = \oint \delta W, when both the terms are measured in the same units.$

- This means that the net heat transfer to the system during the cycle is equal to the net work transfer from the system.
- If a cycle consists of two processes, A and B: Q_A + Q_B = W_A + W_B.
- This law is an empirical statement of the Conservation of Energy Principle applied to thermodynamic systems.

Sign Conventions and Constraints I

- Heat transfer TO the system is treated as positive (+Q); heat transfer FROM the system is negative (-Q).
- Work done BY the system is treated as positive (+W); work done ON the system is negative (-W).
- A Perpetual Motion Machine of the First Kind (PMM1) is a device that would produce net work output without any energy input.
- PMM1 violates the First Law of Thermodynamics ($\oint \delta W = \oint \delta Q$) *and is therefore physically impossible*.

- Although the First Law establishes the equivalence of work and heat, it does not specify the direction of heat transfer.

Summary and Key Takeaways I

- Joule's experiment proved that work and heat are mutually convertible and equivalent forms of energy.
- The Mechanical Equivalent of Heat (J) acts as the conversion factor between mechanical and thermal units.
- For any closed system completing a cycle, the cyclic integrals satisfy: $\oint \delta Q = \oint \delta W$.

- This relationship serves as the foundational empirical validation for the First Law of Thermodynamics.
- The concept sets the stage for defining internal energy (U) as a thermodynamic state property in non-cyclic processes.

Heat and work relation with Joule's Experiment. I

- Course Code: DI03019011
- Course Title: Engineering Thermodynamics
- Lecture 11: Heat and Work Relation with Joule's Experiment
- Instructor: Professor of Mechanical Engineering
- Unit 2: First Law of Thermodynamics

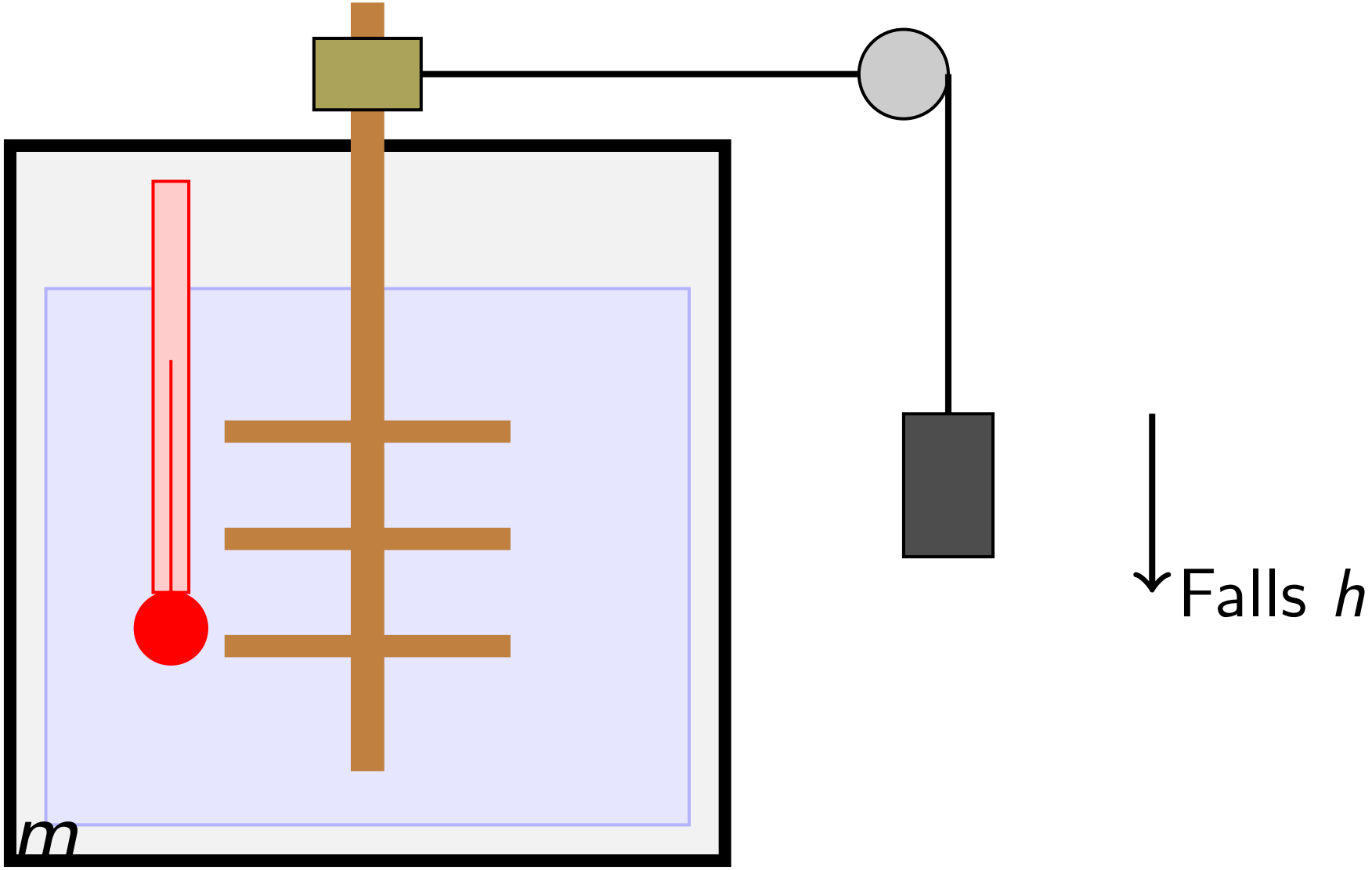
Today's Agenda I

- Distinction between Thermodynamic Heat and Work
- Historical Context: Fall of the Caloric Theory
- Joule's Paddle-Wheel Experimental Setup and Methodology
- The Mechanical Equivalent of Heat (J)
- First Law of Thermodynamics for a Closed System undergoing a Cycle
- Physical and Mathematical Equivalence of Heat and Work

Thermodynamic Heat vs. Work I

- Both heat (Q) and work (W) represent energy in transit across the boundaries of a system.
- Work is defined as energy transfer where the sole effect on things external to the system could be reduced to the raising of a weight.
- Heat is defined as energy transfer across a system boundary solely due to a temperature difference between the system and its surroundings.
- Neither Q nor W are properties of the system; they are path functions, meaning their values depend on the specific process path.
- Their differentials are inexact, denoted by δQ and δW rather than dQ and dW.

Joule's Paddle-Wheel Experiment Setup



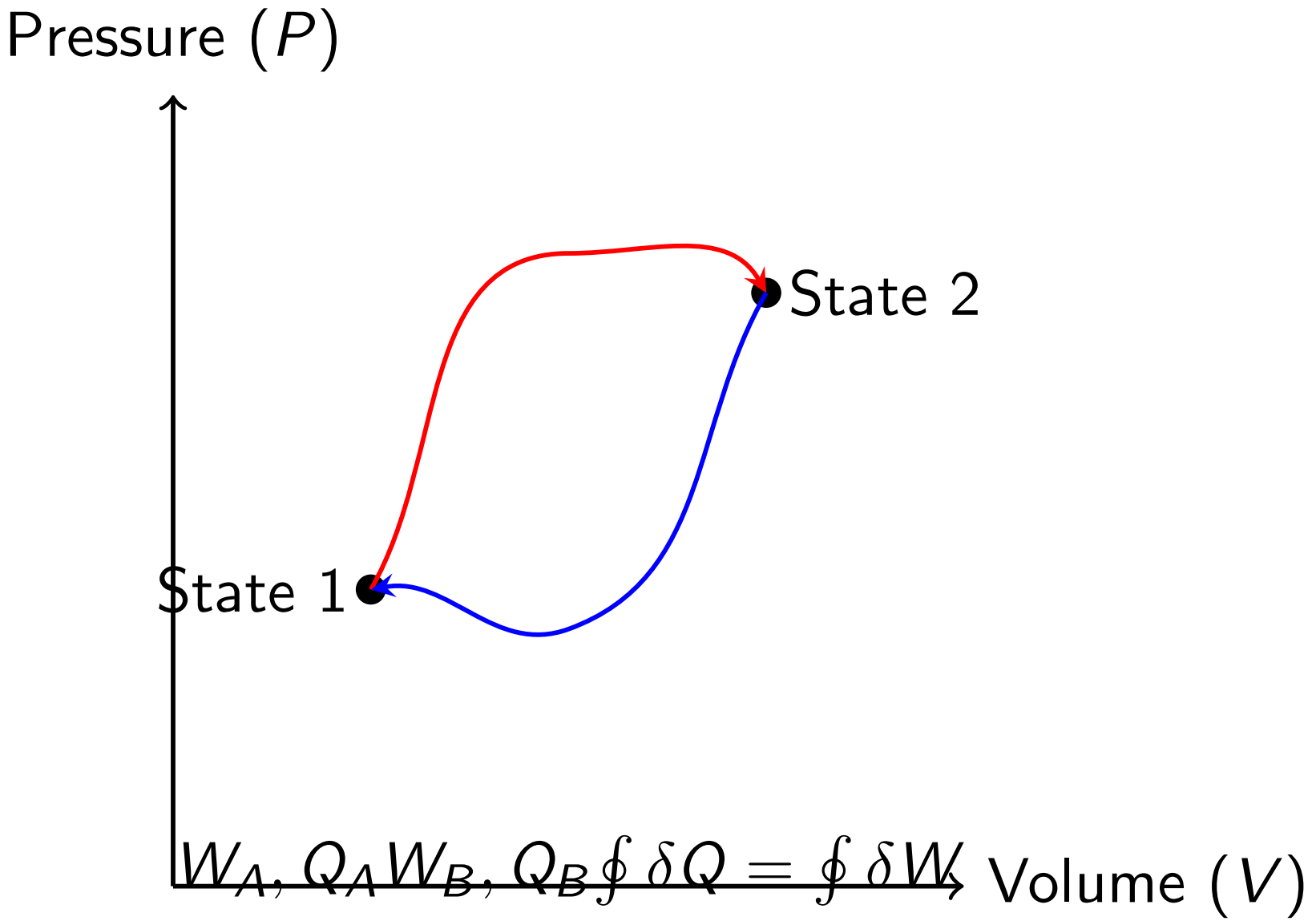
Operation & Physics of the Experiment I

- A known mass (m) is allowed to fall through a measured height (h), doing mechanical work on the water.
- The potential energy loss of the falling weights is $PE_{loss} = mgh$. This energy is transferred via the pulley and drum to rotate the paddle wheel.
- The rotation of the paddle wheel is resisted by the fluid's viscosity, converting the mechanical work into thermal internal energy.
- The insulation around the vessel ensures the process is adiabatic ($Q = 0$), preventing heat loss to the surroundings.
- The temperature rise of the water (ΔT) is measured precisely using a highly sensitive thermometer.

Quantitative Results & Mechanical Equivalent of Heat I

- Joule's experiments established that a fixed amount of mechanical work (W) always produces a proportional rise in temperature equivalent to a fixed quantity of heat (Q).
- The mechanical equivalent of heat (J) is defined by the relation: $W = J \cdot Q$.
- In historical units, Joule determined that 778 ft-lb$_f$ of work is equivalent to 1 Btu of heat, or 4.184 Joules $=$ 1 calorie.
- This discovery led directly to the abandonment of the Caloric Theory, which viewed heat as an indestructible fluid.
- In modern SI units, work and heat are both measured in Joules (J), making the conversion factor $J = 1$ J/J.

First Law of Thermodynamics for a Cycle



Mathematical Formulation of the First Law I

- For any system undergoing a thermodynamic cycle, the cyclic integral of heat is proportional to the cyclic integral of work.
- Mathematically represented as: $\oint \delta Q = \oint \delta W$ in SI units.
- This cyclic relation can be expanded for a cycle consisting of multiple processes: $\sum Q_{net} = \sum W_{net}$.
- Sign convention: Heat transfer INTO the system ($Q_{in} > 0$) and work done BY the system ($W_{out} > 0$) are positive.
- Heat transfer OUT OF the system ($Q_{out} < 0$) and work done ON the system ($W_{in} < 0$) are negative.

Joule's Alternative Verification Experiments I

- To ensure the equivalence was a general law, Joule performed several other experiments using different forms of work.
- Electrical Work: Passing an electric current through a resistor immersed in a fluid to measure temperature rise relative to electrical energy input (I^2Rt).
- Compression of Gas: Measuring the work done in compressing a gas and the corresponding heat rejected to a surrounding water bath.
- Free Expansion of Gas: Letting a gas expand into an evacuated vessel without performing work or transferring heat, proving no change in temperature for an ideal gas.
- These diverse setups consistently confirmed that mechanical, electrical, and thermal energies are mutually convertible.

Key Takeaways and Summary I

- Heat and work are different modes of energy transfer across system boundaries, not stored properties.
- Joule's experiment proved that heat is a form of energy, leading directly to the formulation of the First Law of Thermodynamics.
- For a closed system executing a cycle, the net heat transfer equals the net work done: $\oint \delta Q = \oint \delta W$.
- This relation lays the groundwork for the concept of 'Internal Energy' (U) as a thermodynamic state property.
- Understanding the equivalence of heat and work is crucial for designing power plants, engines, and refrigeration systems.

Law of conservation of energy. I

- Course: Engineering Thermodynamics
- Unit: First Law of Thermodynamics

Agenda I

- Introduction to the First Law of Thermodynamics
- Concept of Energy Conservation
- Closed Systems vs. Open Systems
- Mathematical Formulation of the First Law
- Internal Energy, Heat, and Work Transfer
- Applications to Steady-Flow Processes
- Interactive System Diagram
- Energy Balance for Control Volumes
- Cycle Analysis and Perpetual Motion Machines
- Key Takeaways and Summary

The Principle of Conservation of Energy I

- Energy can neither be created nor destroyed; it can only change from one form to another.
- The total energy of an isolated system remains constant over time.
- Energy interactions across system boundaries occur solely in the forms of heat transfer (Q) and work transfer (W).
- For any process in a closed system, the net change in system energy is equal to the net heat added minus the net work done by the system: $\Delta E = Q - W$ (using sign convention: heat in is positive, work out is positive).

Mass = Constant
$E_1 \rightarrow E_2$; [-¿, ultra thick, red!80!black] (-1.5,2) – node[above] Heat In (Q) (0,2); [-¿, ultra thick, blue!80!black] (4,1) – node[above] Work Out (W) (5.5,1); [dashed, red, thick] (-0.3,-0.3) rectangle (4.3,3.3); ode[anchor=north] at (2,-0.5) $\Delta E = Q - W$;

Mathematical Representation of the First Law I

- The general energy balance equation is: E_in - E_out = ΔE_system
- For a simple compressible system, the total energy (E) consists of Internal Energy (U), Kinetic Energy (KE), and Potential Energy (PE): $E = U + KE + PE$
- Therefore, $\Delta E = \Delta U + \Delta KE + \Delta PE = Q - W$
- In differential form: $dE = dQ - dW$, where dE is a state function (exact differential), while dQ and dW are path functions (inexact differentials).

Forms of Energy: U, KE, and PE I

- Internal Energy (U): Microscopic energy associated with molecular translation, rotation, vibration, and intermolecular forces. Depends primarily on temperature for ideal gases.
- Kinetic Energy (KE): Macroscopic energy associated with system motion relative to a reference frame. KE = 0.5 * m * V^2, hence ΔKE = 0.5 * m * (V_2^2 - V_1^2).
- Potential Energy (PE): Macroscopic energy associated with system elevation in a gravitational field. PE = m * g * z, hence ΔPE = m * g * (z_2 - z_1).

(CV); [-¿, ultra thick, green!60!black] (-1.5,0.5) – node[above] $\dot{m}_{in}, h_{in}$ (0,0.5); [-¿, ultra thick, orange!80!black] (4,0.5) – node[above] $\dot{m}_{out}, h_{out}$ (5.5,0.5); [-¿, ultra thick, red!80!black] (1,3) – node[right] $\dot{Q}_{in}$ (1.5,1.5); [-¿, ultra thick, blue!80!black] (2.5,1.2) – node[right] $\dot{W}_{out}$ (3.5,-0.5); [dashed, blue, thick] (-0.3,-1.7) rectangle (4.3,2.5);

Steady-Flow Energy Equation (SFEE) I

- For steady-flow systems, the mass and energy within the control volume remain constant: dM_cv/dt = 0 and dE_cv/dt = 0.
- Conservation of mass requires: Sum of m_dot_in = Sum of m_dot_out.
- The Steady-Flow Energy Equation (SFEE) on a rate basis is: Q_dot - W_dot = Sum[m_dot_out * (h_out + V_out^2/2 + g*z_out)] - Sum[m_dot_in * (h_in + V_in^2/2 + g*z_in)].
- Enthalpy (h = u + Pv) represents the sum of internal energy and flow work (Pv) required to push mass across the control surface.

Perpetual Motion Machines of the First Kind I

- A Perpetual Motion Machine of the First Kind (PMM1) is a device that violates the First Law of Thermodynamics.
- It produces net work continuously without any net energy input (i.e., W_out ¿ 0 with Q_in = 0, or creating energy from nothing).
- According to the Law of Conservation of Energy, such a device is physically impossible to construct.
- Any machine claiming to deliver more energy than it consumes is a PMM1 and violates the thermodynamic foundation of engineering.

Lecture Summary & Homework I

- Energy is a thermodynamic property; its value changes only through boundary interactions of Heat and Work.
- The First Law provides a quantitative balance between energy input, energy output, and internal change.
- For closed systems: Q - W = ΔU + ΔKE + ΔPE.
- For open systems, flow work is integrated into the enthalpy term (h = u + Pv), yielding the SFEE.
- Homework: Complete problems 2.15 to 2.22 on closed system analysis and SFEE for nozzles and turbines.

Statement of the first law of thermodynamics. I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 2: First Law of Thermodynamics
- Lecture 13: Statement of the first law of thermodynamics.
- Instructor: Professor of Mechanical Engineering

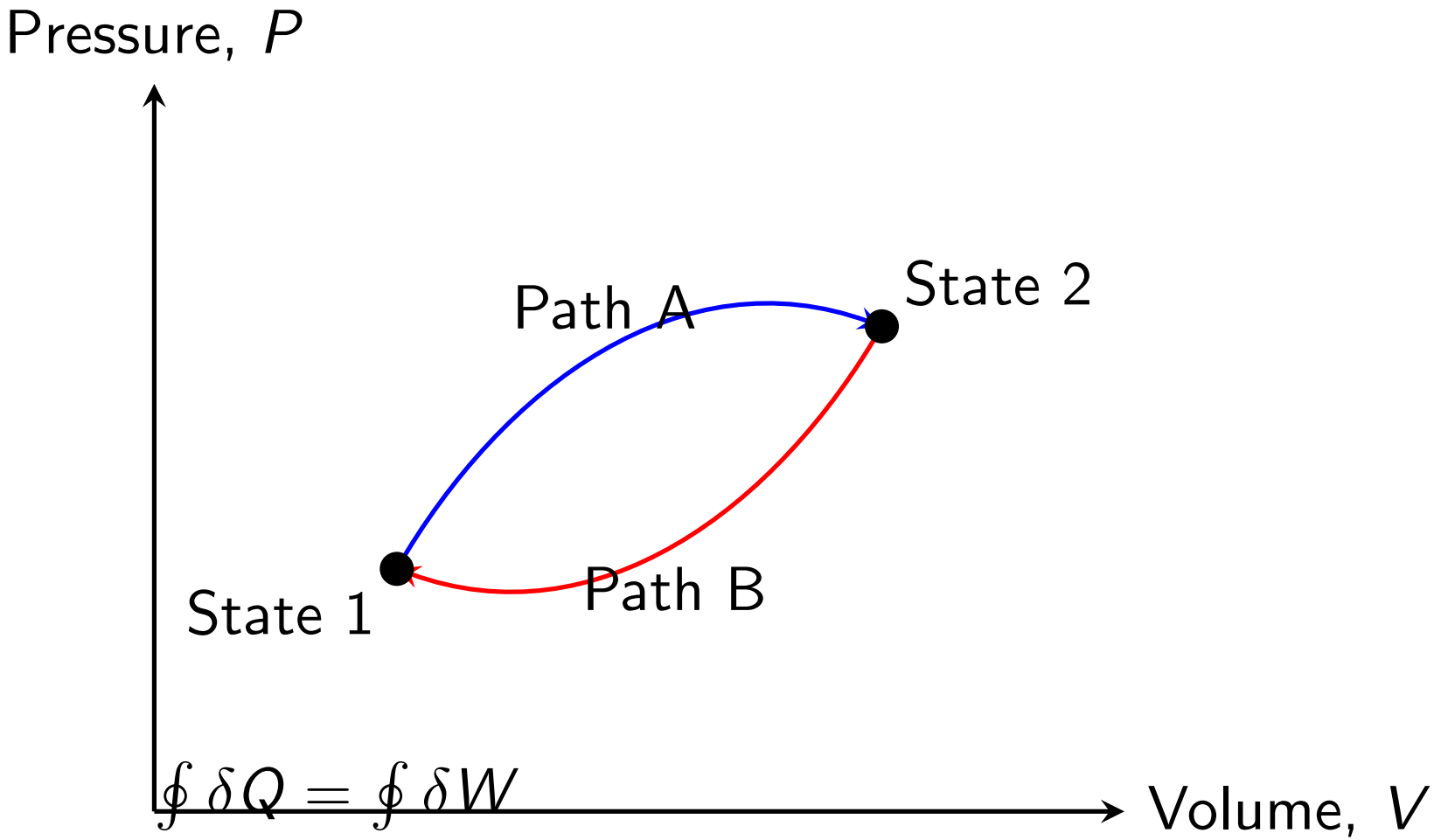
Agenda I

- Historical Context & Joule's Classical Experiments
- First Law of Thermodynamics for a Cyclic Process
- Mathematical Derivation of Energy as a State Property
- First Law for a Change of State (Non-Cyclic Process)
- Components of Total Energy (Internal, Kinetic, Potential)
- Perpetual Motion Machine of the First Kind (PMM1)

Historical Context & Joule's Experiments I

- Prior to the 1840s, heat was treated as 'caloric', a weightless fluid that flowed from hot to cold bodies.
- James Prescott Joule (1840-1849) disproved the caloric theory through a series of rigorous experiments.
- Joule's paddle-wheel apparatus demonstrated that mechanical work done by falling weights on a fluid increases its temperature.
- He established the equivalence of heat and work: a fixed amount of work always produces a corresponding amount of heat.
- Joule's mechanical equivalent of heat ($J \approx 4.184$ J/cal) is now the foundation of SI energy units.

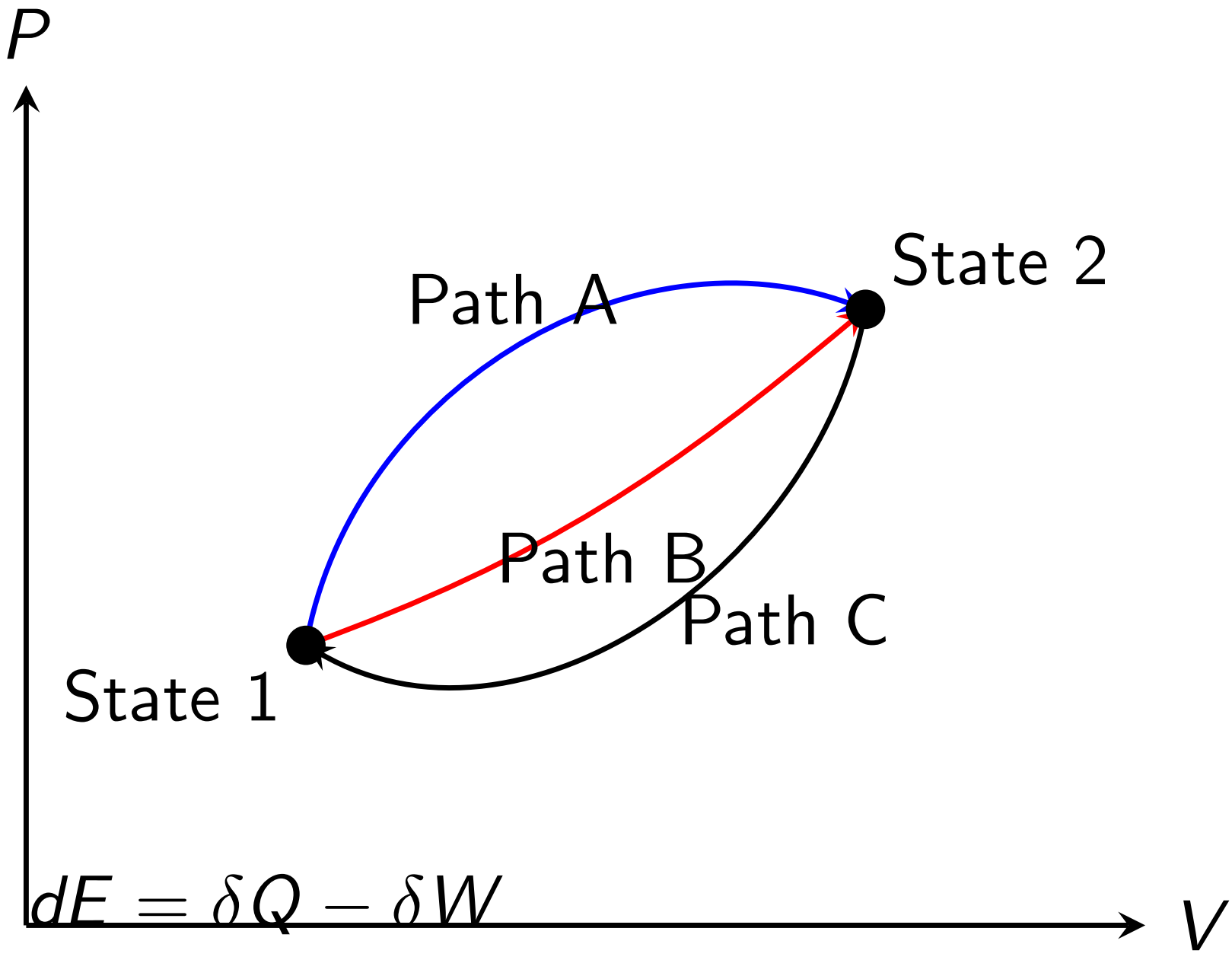
First Law for a Cyclic Process



First Law Statement: Cyclic Process I

- For a closed system undergoing a thermodynamic cycle, the net heat transfer is equal to the net work done.
- Mathematically: $\oint \delta Q = \oint \delta W$, where the integration is carried out over the entire cycle.
- Both heat (Q) and work (W) must be expressed in the same energy units (e.g., Joules).
- The cyclic integral $\oint$ denotes that the system returns exactly to its initial thermodynamic state.
- This equivalence forms the empirical basis of the conservation of energy principle.

Derivation of Energy as a State Property



First Law Statement: Change of State I

- By considering two different cycles 1-A-2-C-1 and 1-B-2-C-1, we find that the integral of $(\delta Q - \delta W)$ is identical for both Path A and Path B.
- Therefore, the value of $\int_1^2 (\delta Q - \delta W)$ is independent of the path and depends only on the initial and final states.
- This proves that $(\delta Q - \delta W)$ represents the change in a thermodynamic property of the system.
- We define this property as the total energy E of the system: $dE = \delta Q - \delta W$.
- For a finite process: $Q - W = \Delta E = E_2 - E_1$.

Energy as a State Property I

- Although heat transfer (Q) and work transfer (W) are path functions, their difference ($Q - W$) is a point function.
- The differential dE is an exact differential, meaning its integration over a cycle is zero: $\oint dE = 0$.
- Energy E is an extensive property of the system, meaning its value scales with the mass of the system.
- The conservation of energy principle states that energy can change form but the total quantity remains constant in an isolated system.

Components of Total Energy (E) I

- The total energy E of a system is the sum of macroscopic energy (kinetic, potential) and microscopic energy (internal).
- Expression: $E = U + KE + PE = U + \frac{1}{2}mV^2 + mgz$, where U is internal energy.
- Internal energy U represents the energy associated with molecular translation, rotation, vibration, and chemical bonds.
- For a stationary closed system, the change in kinetic and potential energy is negligible: $\Delta KE \approx 0$, $\Delta PE \approx 0$.
- Thus, the first law for a stationary closed system is: $Q - W = \Delta U$.

Perpetual Motion Machine of the First Kind (PMM1) I

- A Perpetual Motion Machine of the First Kind (PMM1) is a device that would continuously produce work without any energy input.
- This violates the First Law of Thermodynamics because it violates the conservation of energy principle (creating energy from nothing).
- Mathematically, a PMM1 implies $\oint \delta W > 0$ when $\oint \delta Q = 0$, leading to $\oint dE > 0$, which is impossible.
- No machine can operate continuously and deliver work without consuming an equivalent amount of energy from a source.
- The impossibility of a PMM1 serves as a negative statement and validation of the First Law of Thermodynamics.

Application of the first law of thermodynamics: I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 2: First Law of Thermodynamics
- Lecture 14: Application of the first law of thermodynamics:
- Department of Mechanical Engineering
- Focus: Applying energy conservation principles to closed (non-flow) and open (flow) systems.

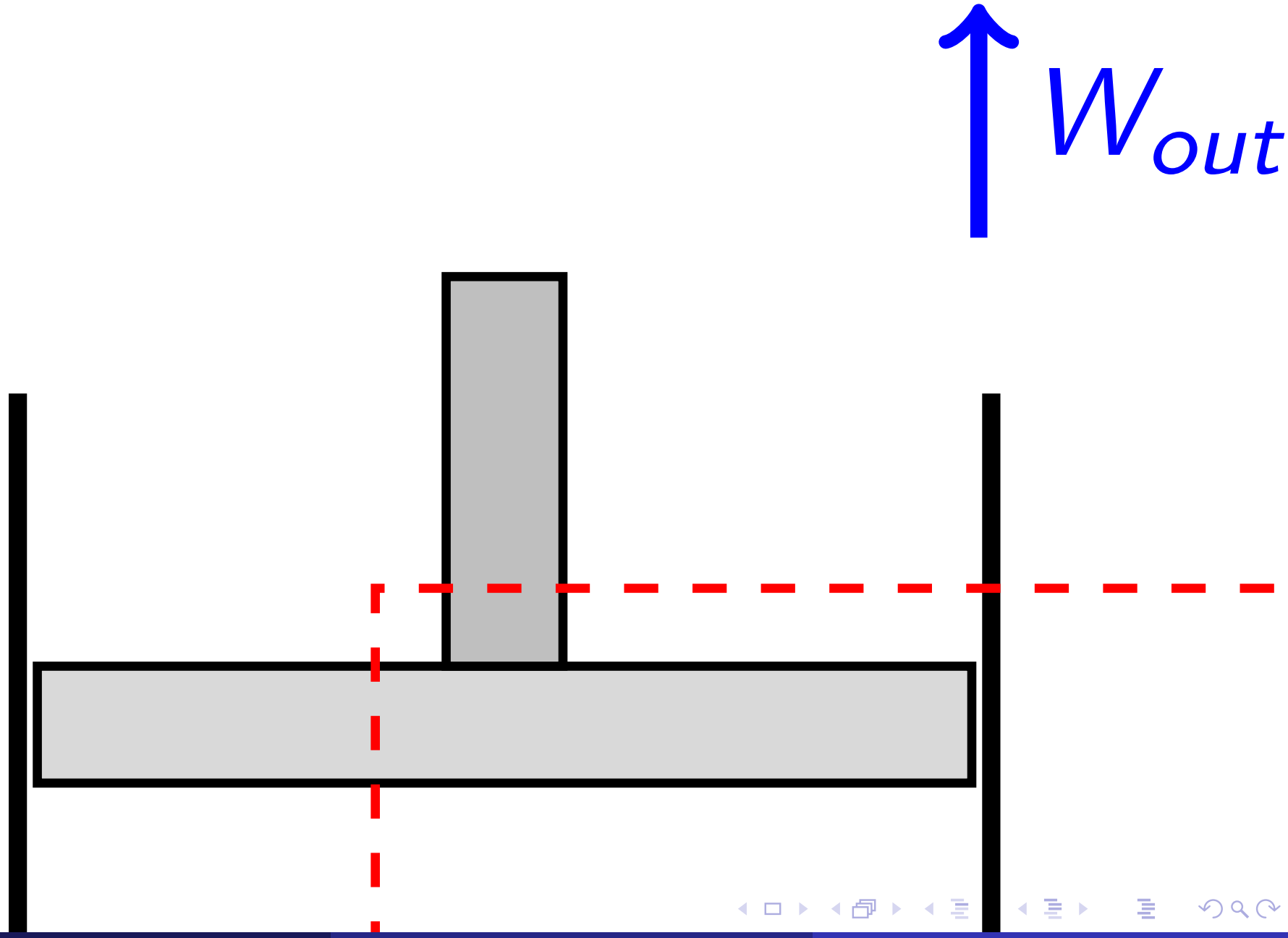
Agenda I

- First Law of Thermodynamics for Closed Systems (Non-flow processes)
- Analysis of specific non-flow processes: Isochoric, Isobaric, Isothermal, Adiabatic, and Polytropic
- First Law of Thermodynamics for Open Systems (Control volumes)
- Derivation and meaning of the Steady Flow Energy Equation (SFEE)
- Application of SFEE to engineering devices: Nozzles, Diffusers, Turbines, Compressors, and Throttling valves

First Law for Closed Systems (Non-Flow) I

- A closed system is a thermodynamic system where mass remains constant within the boundaries, but energy transfer (heat and work) is allowed.
- The first law of thermodynamics for a change of state in a closed system is expressed as: $Q - W = \Delta E$ (where $E = U + KE + PE$).
- For most stationary engineering applications, changes in kinetic energy (KE) and potential energy (PE) are negligible, simplifying the relation to: $Q - W = \Delta U = m(u_2 - u_1)$.
- Here, Q is the net heat transfer into the system, W is the net work done by the system, and U is the internal energy of the system.
- Sign conventions: Heat added to the system ($Q > 0$) and work done by the system ($W > 0$) are positive.

Diagram: Energy Interactions in a Closed Piston-Cylinder System



Thermodynamic Processes in Closed Systems I

- Constant Volume (Isochoric) Process: Since $dV = 0$, displacement work $W = \int PdV = 0$. The first law reduces to $Q = \Delta U = mC_v(T_2 - T_1)$.
- Constant Pressure (Isobaric) Process: Displacement work $W = P(V_2 - V_1)$. The first law becomes $Q = \Delta U + P\Delta V = \Delta H = mC_p(T_2 - T_1)$, where $H = U + PV$ is enthalpy.
- Constant Temperature (Isothermal) Process: For an ideal gas, $\Delta U = 0$, hence heat transfer equals boundary work: $Q = W = P_1V_1 \ln(V_2/V_1) = mRT \ln(P_1/P_2)$.
- Adiabatic Process ($Q = 0$): No heat exchange. The first law gives $W = -\Delta U = -mC_v(T_2 - T_1)$. For reversible adiabatic process: $PV^\gamma = \text{constant}$.

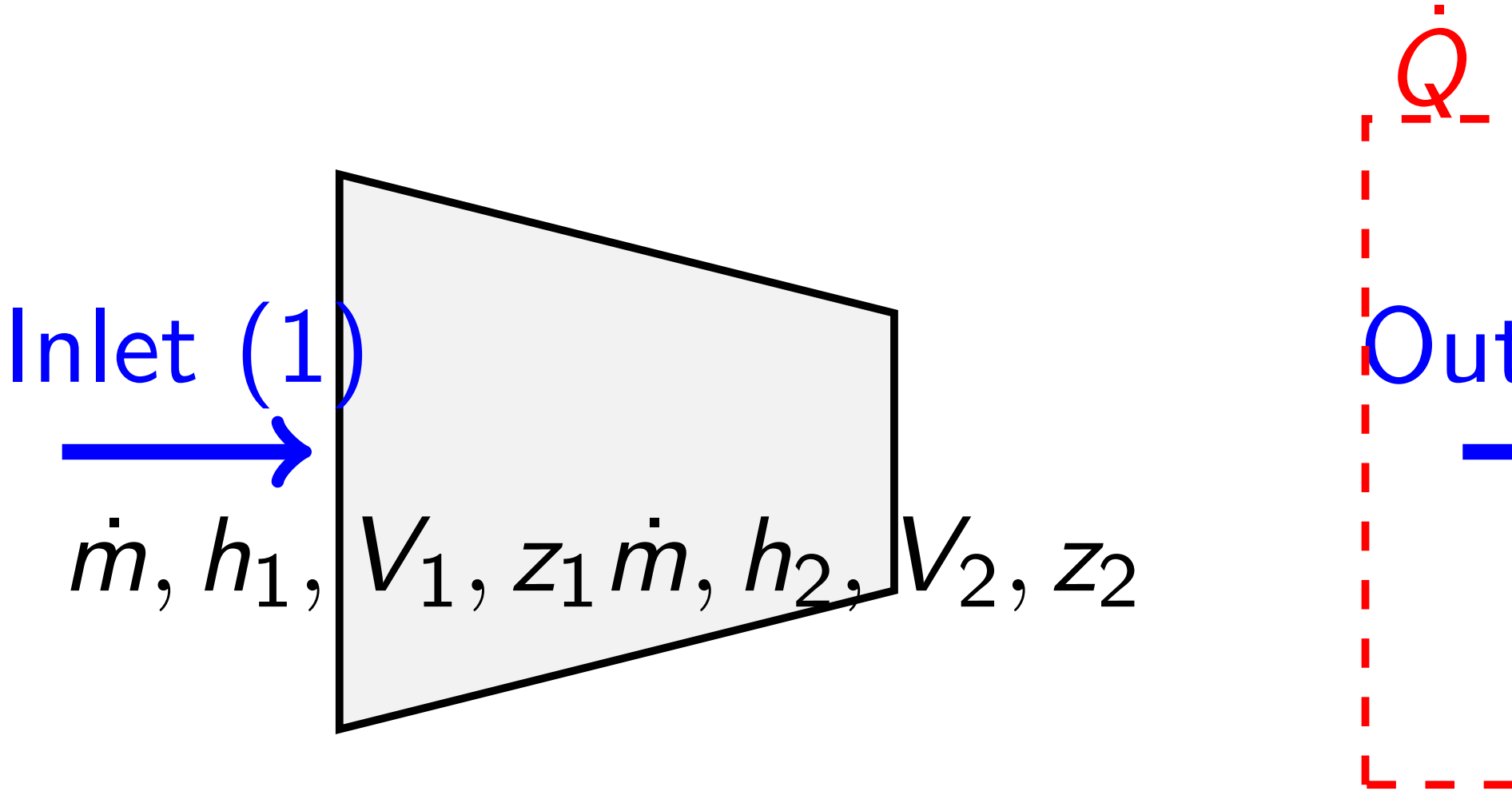
Thermodynamic Processes in Closed Systems II

- Polytropic Process ($PV^n = \text{constant}$): Work done is $W = (P_1V_1 - P_2V_2)/(n-1)$. Heat transfer is $Q = W \times (\gamma - n)/(\gamma - 1)$.

First Law for Open Systems (Control Volumes) I

- An open system, or control volume, allows both mass and energy (heat and work) to cross its boundaries.
- To analyze an open system, we establish a spatial boundary called the control surface (CS), through which fluid flows.
- Conservation of Mass: The rate of change of mass inside the control volume equals the net mass flow rate entering minus leaving: $d(m_{cv})/dt = \sum \dot{m}_{in} - \sum \dot{m}_{out}$.
- Conservation of Energy: The rate of change of energy inside the control volume is given by the net heat transfer rate, net shaft work rate, and the energy carried by fluid streams: $d(E_{cv})/dt = \dot{Q} - \dot{W}_{control\ volume} + \sum \dot{m}_{in}(h + \frac{V^2}{2} + gz)_{in} - \sum \dot{m}_{out}(h + \frac{V^2}{2} + gz)_{out}$.
- Enthalpy ($h = u + Pv$) is used here to account for internal energy (u) and the flow work (Pv) required to push the fluid into and out of the control volume.

Diagram: Steady Flow Control Volume (SFEE Application)



The Steady Flow Energy Equation (SFEE) I

- Under steady flow conditions, the mass inside the control volume remains constant ($d(m_{cv})/dt = 0$) and the energy inside the control volume remains constant ($d(E_{cv})/dt = 0$).
- For a single inlet and single outlet system, the mass flow rate is constant: $\dot{m}_{in} = \dot{m}_{out} = \dot{m} = \rho_1 A_1 V_1 = \rho_2 A_2 V_2$.
- The Steady Flow Energy Equation (SFEE) on a mass rate basis is: $\dot{Q} - \dot{W}_{shaft} = \dot{m}\left[(h_2 - h_1) + \frac{V_2^2 - V_1^2}{2} + g(z_2 - z_1)\right]$.
- On a unit mass basis (divided by $\dot{m}$), the equation is: $q - w = (h_2 - h_1) + \frac{V_2^2 - V_1^2}{2000} + \frac{g(z_2 - z_1)}{1000}$ (where kinetic and potential terms are scaled to kJ/kg).
- This equation is the foundation for analyzing steady-state engineering devices such as turbines, compressors, nozzles, and heat exchangers.

SFEE Applied to Common Engineering Devices I

- Nozzles and Diffusers: Designed to vary velocity. No work is done ($w = 0$), and they are usually insulated ($q = 0$). The SFEE reduces to: $h_1 + V_1^2/2 = h_2 + V_2^2/2$.
- Turbines: Expand high-pressure fluid to produce shaft work. Heat loss is usually negligible ($q \approx 0$) and kinetic/potential energy changes are small. SFEE simplifies to: $w_{out} = h_1 - h_2$.
- Compressors and Pumps: Consume shaft work to increase fluid pressure. Assuming adiabatic operation and negligible KE/PE changes: $w_{in} = h_2 - h_1$.
- Throttling Devices (Valves/Capillaries): Restrict flow causing a pressure drop without work ($w = 0$) or heat transfer ($q = 0$). KE and PE changes are negligible. Thus, $h_1 = h_2$ (isenthalpic process).

SFEE Applied to Common Engineering Devices II

- Boilers and Condensers: Heat exchangers where phase change occurs. Work $w = 0$, and changes in KE and PE are negligible. The SFEE simplifies to: $q = h_2 - h_1$.

Summary and Analytical Methodology I

- Identify the System: Determine if the system is closed (control mass) or open (control volume).
- Determine the State Changes: Use thermodynamic tables or ideal gas equations to find properties (internal energy u, enthalpy h, specific volume v).
- Identify Energy Interactions: Identify boundary work, shaft work, and heat transfer. Note whether the process is adiabatic ($Q = 0$), isothermal ($T = c$), etc.
- Apply Conservation Equations: Use the simplified first law ($Q - W = \Delta U$) for closed systems, or SFEE ($q - w = \Delta h + \Delta KE + \Delta PE$) for steady-flow systems.
- Consistency of Units: Always ensure units match. Convert kinetic energy ($V^2/2$ in m^2/s^2) and potential energy (gz in m^2/s^2) to kJ/kg by dividing by 1000.

Definition of the flow process, control volume and flow work. I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 2: First Law of Thermodynamics
- Lecture 15: Definition of the flow process, control volume and flow work.
- Instructor: Professor of Mechanical Engineering

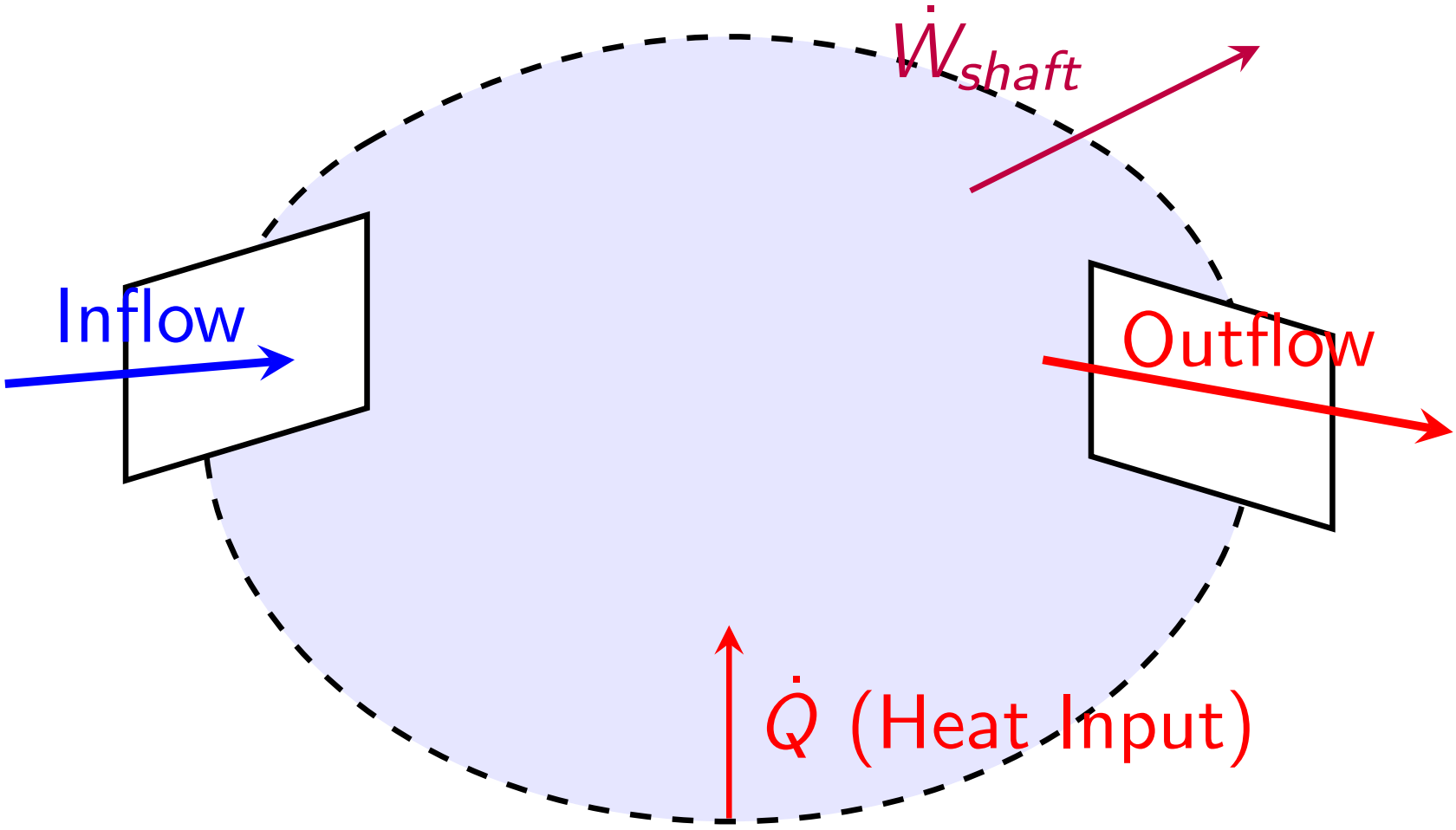
Lecture Agenda I

- Introduction to Open Systems (Control Volumes)
- Concept of Control Surfaces and Boundaries
- Mass Conservation in Control Volumes (Continuity)
- Energy Transfer by Mass Flow
- Concept of Flow Work (Flow Energy)
- Mathematical Derivation of Flow Work
- Total Energy of a Flowing Fluid
- First Law of Thermodynamics for a Control Volume
- Summary and Practical Engineering Applications

What is a Control Volume? I

- A control volume (CV) is a selected region in space chosen for thermodynamic analysis of open systems.
- Unlike a closed system (control mass), mass can cross the boundary of a control volume.
- The boundary of a control volume is called the control surface (CS), which can be real (physical) or imaginary.
- The control volume can be fixed in size and shape (e.g., a nozzle), or it can involve moving boundaries (e.g., a piston-cylinder device with fluid inlet/outlet).
- Examples of control volumes include pumps, compressors, turbines, heat exchangers, and nozzles.

Schematic of a Control Volume



Conservation of Mass (Continuity Equation) I

- The principle of conservation of mass states that mass cannot be created or destroyed.
- For a control volume, the rate of change of mass within the CV must equal the net mass flow rate across the boundary.
- Mathematical formula: $\text{dM\_cv/dt} = \sum\text{\_}in - \sum\text{\_}out$

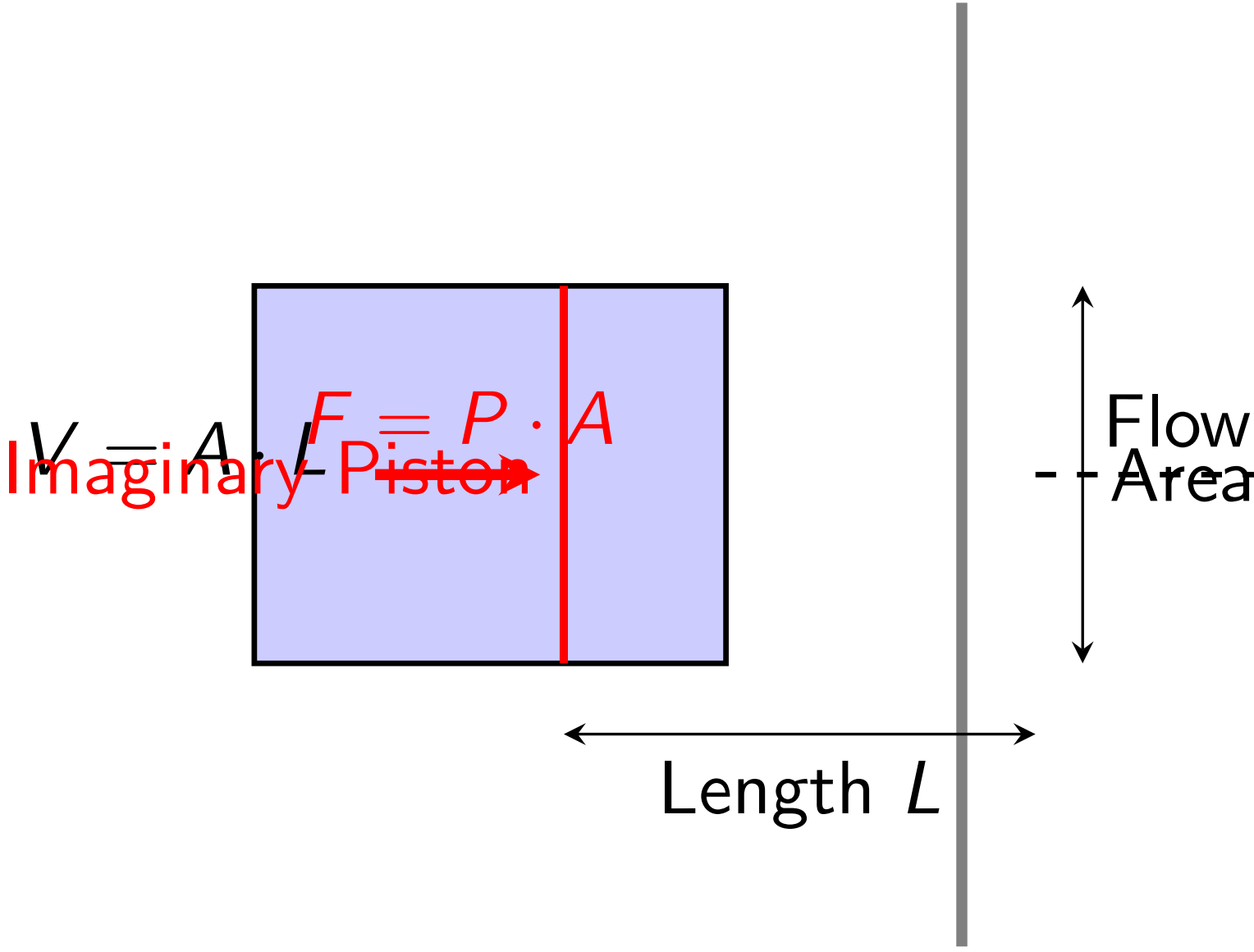
- For steady-state processes, the mass within the control volume remains constant ($\text{dM\_cv/dt} = 0$), so $\sum\text{\_}in = \sum\text{\_}out$.

- Mass flow rate ($\dot{m}$) can be expressed as $\dot{m} = \rho \cdot \text{A} \cdot \text{V\_avg}$, where ρ is density, A is cross-sectional area, and V_avg is average velocity.

The Concept of Flow Work I

- Unlike closed systems, open systems involve mass entering and leaving the control volume.
- Work is required to push fluid into or out of the control volume; this is called flow work (or flow energy).
- Flow work is necessary to maintain a continuous flow of fluid through the control volume.
- Consider a fluid element of volume V and pressure P entering the CV; the force acting on the fluid element is $F = P \cdot A$.
- If the fluid element of length L is pushed into the CV, the work done is $W\_flow = F \cdot L = P \cdot A \cdot L = P \cdot V$.

Flow Work Derivation Model



Flow Work and Enthalpy I

- On a unit mass basis, specific flow work is: w_flow $=$ P $\cdot$ v, where v is the specific volume (v $=$ V/m).
- Flow work is expressed per unit mass as P $\cdot$ v, which is a state function since both P and v are thermodynamic properties.
- The total energy of a non-flowing fluid is internal energy: e $=$ u $+$ ke $+$ pe.
- The total energy of a flowing fluid includes flow work: θ $=$ e $+$ w_flow $=$ u $+$ P$\cdot$ v $+$ ke $+$ pe.
- Since u $+$ P$\cdot$ v is defined as enthalpy (h), the total energy of a flowing fluid simplifies to: θ $=$ h $+$ V^2/2 $+$ g$\cdot$ z.

Energy Transport by Mass Flow I

- Mass flow carries energy across the control surface in three forms: internal energy, kinetic energy, and potential energy, in addition to flow work.
- Rate of energy transport by mass: $\dot{E}$_mass $= \dot{m} \cdot \theta = \dot{m} \cdot$ (h + V^2/2 + g· z).
- Enthalpy (h = u + Pv) naturally combines the internal energy of the fluid and the flow work needed to move it.
- For steady flow, the energy entering the control volume via heat, work, and mass must equal the energy exiting through these same mechanisms.
- This formulation forms the basis for the Steady-Flow Energy Equation (SFEE) used in analyzing real engineering systems.

Key Summary and Applications I

- A control volume is essential for open system thermodynamic analysis where mass flow occurs.
- Flow work (Pv) represents the energy required to push fluid across the control surface boundaries.
- Integrating flow work with internal energy yields Enthalpy (h = u + Pv), which simplifies open system energy equations.
- The total energy of a flowing fluid per unit mass is θ = h + V^2/2 + g· z.
- These principles are critical for designing and analyzing thermodynamic devices like steam turbines, gas compressors, and refrigeration cycles.

Definition of the flow process, control volume and flow work. I

- Course: Engineering Thermodynamics
- Unit 2: First Law of Thermodynamics for Open Systems
- Lecture 16: Definition of the Flow Process, Control Volume, and Flow Work

Lecture Agenda I

- Introduction to Flow Processes (Open Systems)
- Definition and Boundary of a Control Volume
- Physical Meaning of Flow Work (Flow Energy)
- Mathematical Derivation of Flow Work
- Total Energy of a Flowing Fluid
- Significance of Enthalpy in Flow Processes

Introduction to Flow Processes I

- Many engineering devices involve the flow of fluid through them (e.g., pumps, turbines, nozzles, compressors, heat exchangers).
- Such devices cannot be analyzed using closed-system (control mass) relations because mass crosses the system boundaries.
- These systems are classified as open systems or flow processes, where thermodynamic properties are analyzed in a region of space.
- The analysis requires accounting for mass transport as well as heat and work interactions across the boundary.

(CV); [¡-] (3.8,0.3) – (5,1) node[right] Control Surface (CS); [ultra thick, -¿, blue] (-1,-0.5) – (0.3,-0.5) node[midway, above] Inlet;
ode at (-1,-1.2) [align=left] $\dot{m}_{in}$
p_1, v_1, T_1; [ultra thick, -¿, red] (4,-1) – (5.3,-1) node[midway, above] Outlet;
ode at (5.5,-1.7) [align=left] $\dot{m}_{out}$
p_2, v_2, T_2; [thick, -¿, orange!80!black] (2,1.8) – (2,0.8) node[midway, left] $\dot{Q}$; [thick, -¿, black] (3,-2) – (3,-3) node[midway, right] $\dot{W}_{shaft}$;

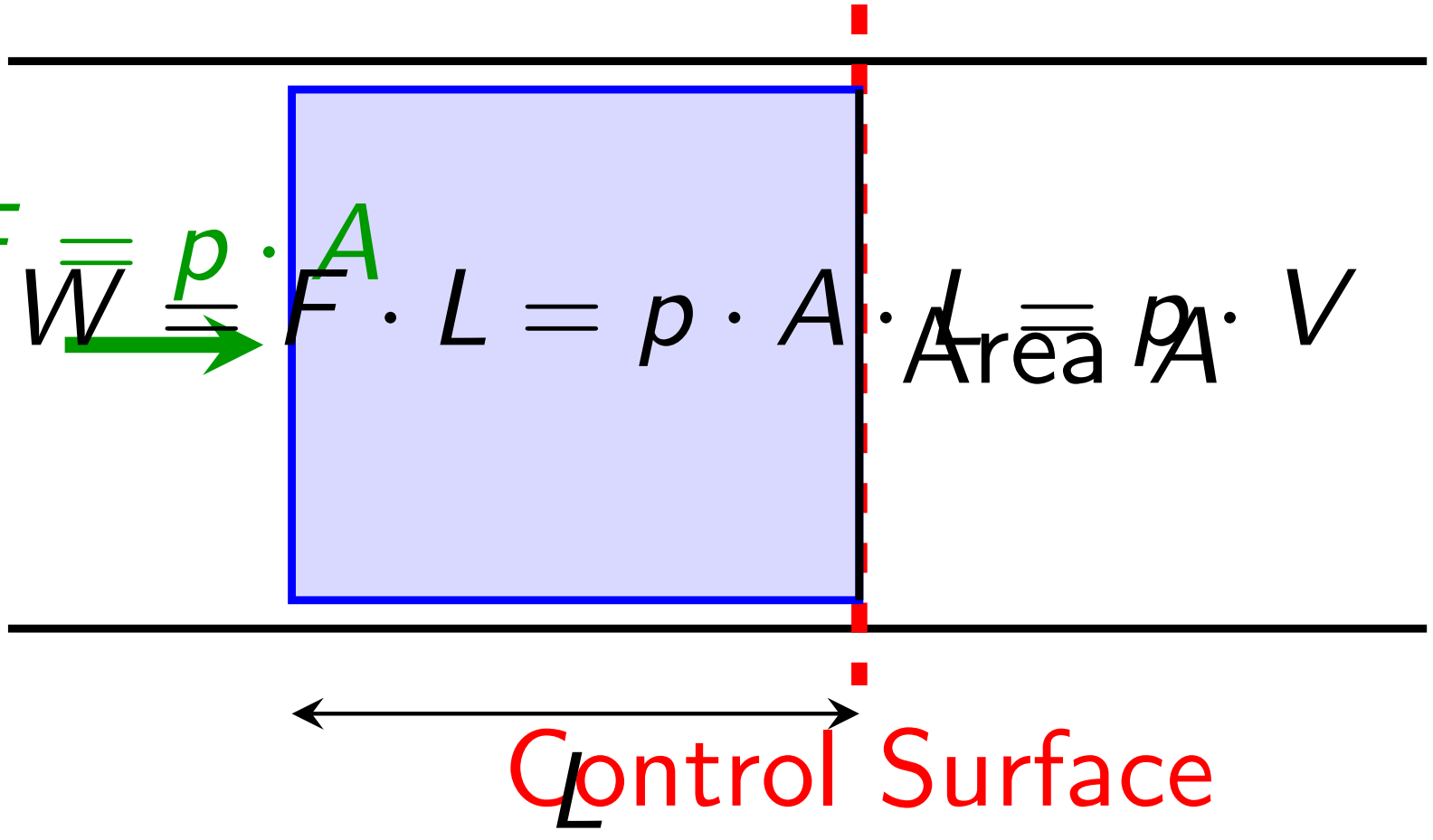
Control Volume and Control Surface I

- A control volume (CV) is a selected region in space through which mass and energy can flow.
- The boundary of a control volume is called the control surface (CS), separating the CV from its surroundings.
- Control surfaces can be real physical walls (such as turbine casing) or imaginary planes (such as inlet and outlet flow areas).
- Control volumes can be fixed in shape and position (typical for steady flow devices) or moving and deforming.

Concept of Flow Work I

- Flow work (or flow energy) is the energy required to push a fluid element across the control surface.
- Unlike closed systems where work is done by moving boundaries, flow work is done by the fluid upstream to push mass in, or by the system to push mass out.
- For a stationary fluid element, there is no flow work; it is a dynamic quantity that only exists when fluid is flowing across a boundary.
- It is a vital component when accounting for the total energy transport associated with mass flow.

Visualization and Derivation of Flow Work



Mathematical Representation of Flow Work I

- Consider a fluid element of pressure p, cross-sectional area A, and length L crossing the control surface.
- The force acting on the element by the upstream fluid is $F = pA$, and the distance it travels to cross the CS is L.
- The work done to push the element across the boundary is $W_{flow} = FL = (pA)L = pV$, where V is the volume of the element.
- On a unit mass basis, specific flow work is $w_{flow} = pv$ (where $v = V/m$ is the specific volume of the fluid).
- Flow work is positive at the inlet (work done on the CV) and negative at the outlet (work done by the CV).

Total Energy of a Flowing Fluid I

- A flowing fluid carries internal energy (u), kinetic energy ($V^2/2$), and potential energy (gz) per unit mass.
- Additionally, the fluid carries the flow work (pv) required to cross the control surface.
- The total energy of a flowing fluid per unit mass is: $\theta = u + pv + V^2/2 + gz$.
- Since enthalpy is defined as $h = u + pv$, the total energy can be rewritten as: $\theta = h + V^2/2 + gz$.
- This explains why enthalpy is a convenient thermodynamic property to use in open-system (control volume) energy equations.

Summary & Key Takeaways I

- An open system is analyzed using a Control Volume (CV) enclosed by a Control Surface (CS).
- Flow work (pv) is the energy needed to push mass across the control surface, critical for open systems.
- Enthalpy ($h = u + pv$) combines internal energy and flow work, simplifying thermodynamic analysis of flow processes.
- The total energy of a flowing fluid per unit mass is given by $\theta = h + V^2/2 + gz$.

Steady and unsteady flow processes. I

- Course: Engineering Thermodynamics
- Unit: First Law of Thermodynamics

Agenda I

- Concept of Control Volume and Control Surface
- Conservation of Mass: Continuity Equation
- Steady Flow Processes and Steady Flow Energy Equation (SFEE)
- Thermodynamic Analysis of Steady Flow Engineering Devices
- Introduction to Unsteady (Transient) Flow Processes
- Analysis of Transient Processes: Charging and Discharging of Vessels

Control Volume and Control Surface Concepts I

- A Control Volume (CV) is a region in space selected for thermodynamic study of systems involving mass flow.
- The boundary of a control volume is called the Control Surface (CS), which can be real, imaginary, fixed, or moving.
- Mass, heat, and work can cross the control surface during a process.
- Steady Flow: Fluid properties at any point within the control volume do not change with time.
- Unsteady (Transient) Flow: Fluid properties and mass/energy storage within the CV vary with time.

m_{CV}, E_{CV}; [-¿, ultra thick, blue!80!black] (-0.5,1) – (1.2,1);
ode at (-0.5,1.4) $\dot{m}_i, \theta_i$; [-¿, ultra thick, red!80!black] (5.8,-0.5) – (7.5,-0.5);
ode at (7.5,-0.1) $\dot{m}_e, \theta_e$; [-¿, thick, red, double] (2.5,2.8) – (3.2,1.8) node[midway, above right] $\dot{Q}$; [-¿, thick, brown!80!black, double] (4.5,-1.2) – (5.5,-2.2) node[midway, below left] $\dot{W}_{sh}$;
ode[draw, dashed, inner sep=2pt, fill=white] at (2.2, 1.8) CS;

Conservation of Mass: Continuity Equation I

- Mass conservation states that the net rate of mass accumulation within a control volume equals the net mass flow rate in minus out.
- General rate equation: $dm_{CV}/dt = \sum \dot{m}_i - \sum \dot{m}_e$, where $\dot{m}$ is mass flow rate.
- Mass flow rate expression: $\dot{m} = \rho AV = AV/v$, with density ρ, area A, velocity V, and specific volume v.
- For steady flow, mass accumulation is zero ($dm_{CV}/dt = 0$), yielding: $\sum \dot{m}_i = \sum \dot{m}_e$.
- For a single-stream steady flow device, the continuity equation simplifies to: $\dot{m}_1 = \dot{m}_2 \implies \rho_1 A_1 V_1 = \rho_2 A_2 V_2$.

Steady Flow Energy Equation (SFEE) I

- Under steady-flow conditions, the total energy content of the control volume remains constant ($dE_{CV}/dt = 0$).
- The first law rate equation:
  $\dot{Q} - \dot{W} = \sum \dot{m}_e(h_e + V_e^2/2 + gz_e) - \sum \dot{m}_i(h_i + V_i^2/2 + gz_i)$.
- Enthalpy ($h = u + pv$) represents the sum of internal energy (u) and flow work (pv) required to push fluid across the boundaries.
- For a single-stream steady flow device:
  $q - w = (h_2 - h_1) + (V_2^2 - V_1^2)/2 + g(z_2 - z_1)$, where $q = \dot{Q}/\dot{m}$ and $w = \dot{W}/\dot{m}$.
- In many thermodynamic applications, potential energy changes and kinetic energy changes are negligible relative to enthalpy changes.

Thermodynamic Analysis of Engineering Devices I

- Nozzles & Diffusers: Work $w = 0$ and heat $q \approx 0$. Nozzles increase velocity ($V_2 > V_1$) while diffusers decrease it. SFEE: $h_1 + V_1^2/2 = h_2 + V_2^2/2$.
- Turbines: Produce work from fluid expansion. Assuming adiabatic ($q \approx 0$) and negligible KE/PE: $w_{out} = h_1 - h_2$.
- Compressors & Pumps: Consume work to increase pressure. Assuming adiabatic ($q \approx 0$) and negligible KE/PE: $w_{in} = h_2 - h_1$.
- Throttling Devices: Restrict flow causing a pressure drop without work or heat exchange. Process is isenthalpic: $h_1 = h_2$.
- Heat Exchangers: Transfer heat between fluid streams without work ($w = 0$). Energy balance: $\sum \dot{m}_i h_i = \sum \dot{m}_e h_e$.

Unsteady Flow Processes (Transient Analysis) I

- Unsteady flow processes involve states and properties within the control volume that change continuously over time.
- Examples include charging a rigid cylinder from a supply line, or discharging/emptying a pressurized vessel.
- Uniform-State, Uniform-Flow (USUF) model assumes uniform state in the CV at any instant, and uniform properties at inlets/exits.
- USUF Mass Balance: $m_2 - m_1 = \sum m_i - \sum m_e$, where 1 and 2 are the initial and final states of the control volume.
- USUF Energy Balance: $Q - W = (m_2 u_2 - m_1 u_1)_{CV} + \sum m_e h_e - \sum m_i h_i$ (neglecting kinetic and potential energies).

Volume $V =$
textconst
Initial: m_1, T_1, u_1
Final: m_2, T_2, u_2; [-¿, blue, ultra thick] (-0.8, 0) – (0.2, 0)
node[midway, below] $\dot{m}_i$; [-¿, blue, ultra thick] (1.2, 0) – (1.8, 0);
[-¿, red, thick] (3.25, 1.5) – (3.25, 2.2) node[midway, right] Q;

Summary of Flow Processes I

- Steady Flow: Mass and energy content within the CV remain constant; inflow equals outflow for mass and energy.
- Unsteady Flow: Inflow and outflow rates do not balance, leading to accumulation or depletion of mass and energy over time.
- Thermodynamic properties of the flowing fluid are accounted for using enthalpy ($h = u + pv$) due to the associated flow work.
- Properties stored inside the control volume are evaluated using internal energy (u).
- Adiabatic charging of an evacuated rigid tank ($m_1 = 0$, $Q = 0$) with an ideal gas yields $u_2 = h_i$, which results in a temperature rise $T_2 = \gamma T_i$.

Steady Flow Energy Equations (SFEE) and its applications in Nozzle, Diffuser, Boiler, Turbine, Compressor, Condenser, and throttling devices. I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 2: First Law of Thermodynamics for Open Systems
- Focus: Steady Flow Energy Equation (SFEE) formulation and application to core engineering devices including nozzles, diffusers, boilers, turbines, compressors, condensers, and throttling valves.
- Instructor: Professor of Mechanical Engineering

Agenda I

- Steady Flow Systems vs. Unsteady Flow Systems and fundamental assumptions
- Derivation of the Steady Flow Energy Equation (SFEE) from energy conservation principles
- Analysis of Nozzles and Diffusers: Acceleration and pressure rise dynamics
- Analysis of Boilers and Condensers: Constant pressure heat addition and rejection
- Analysis of Turbines and Compressors: Shaft work extraction and consumption
- Analysis of Throttling Devices: Constant enthalpy pressure drop processes

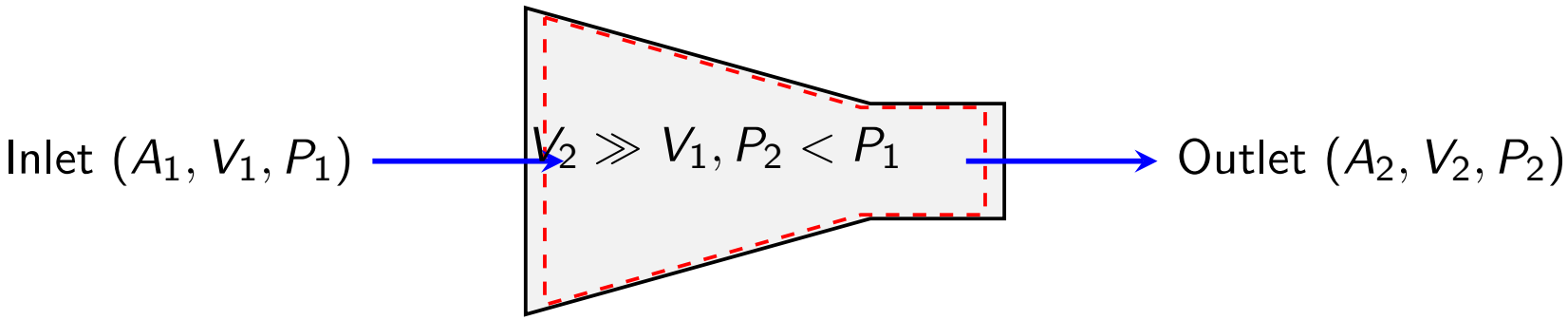
Understanding Steady Flow Systems I

- Definition: An open system in which fluid flows at a constant rate and fluid properties at any point within the control volume do not change over time.
- Conservation of Mass: $\dot{m}_{in} = \dot{m}_{out} =$ constant, which means $\rho_1 A_1 V_1 = \rho_2 A_2 V_2$ for one-dimensional inlet and exit.
- Zero Accumulation: Mass accumulation $(dm_{CV}/dt = 0)$ and energy accumulation $(dE_{CV}/dt = 0)$ inside the control volume are zero.
- Constant State Properties: State of fluid at inlet, outlet, and any point inside the control volume is invariant with time.
- Heat and Work Rates: Energy transfers as heat $(\dot{Q})$ and work $(\dot{W})$ across the system boundary occur at constant rates.

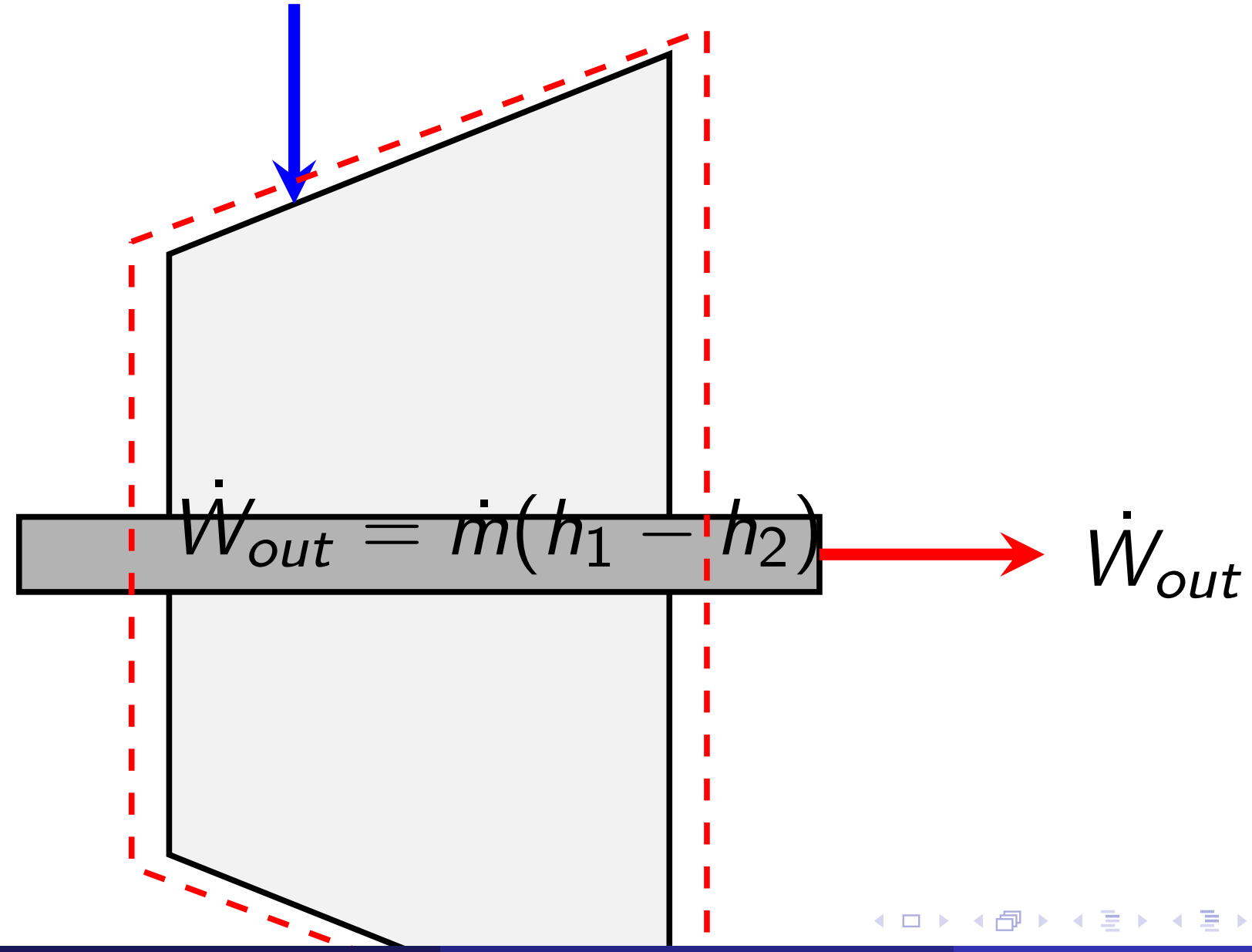
Deriving the Steady Flow Energy Equation I

- Mathematical formulation from First Law of Thermodynamics for open systems.
- General equation:
  $\dot{Q} - \dot{W}_{shaft} = \dot{m}[(h_2 - h_1) + \frac{V_2^2 - V_1^2}{2} + g(z_2 - z_1)]$.
- Enthalpy term: Enthalpy $h = u + Pv$ combines internal energy u and flow work Pv required to sustain fluid motion across boundaries.
- Per unit mass basis:
  $q - w_{shaft} = (h_2 - h_1) + \frac{V_2^2 - V_1^2}{2} + g(z_2 - z_1)$, where $q = \dot{Q}/\dot{m}$ and $w_{shaft} = \dot{W}_{shaft}/\dot{m}$.
- Dimensional Consistency: Enthalpy is typically in kJ/kg, so kinetic energy ($V^2/2000$) and potential energy ($gz/1000$) must be scaled by 1000.

Nozzles and Diffusers: Velocity and Pressure Transformers



Turbines: Extracting Mechanical Work



Compressors and Pumps: Adding Work to Fluids I

- Compressor & Pump: Devices designed to elevate fluid pressure by consuming external mechanical power.
- Gas vs. Liquid: Compressors compress gaseous media (significant volume change), while pumps pressurize incompressible liquids (volume remains nearly constant).
- SFEE Assumptions: Negligible kinetic and potential energy changes ($\Delta ke \approx 0$, $\Delta pe \approx 0$). Adiabatic ($q = 0$) for fast processes, or isothermal/polytropic for cooled compressors.
- Work input equation: $\dot{W}_{in} = \dot{m}(h_2 - h_1)$ or $w_{in} = h_2 - h_1$. Shaft work enters system ($w < 0$, so $w_{in} = -w_{shaft}$).
- Pump work simplification: Since liquid density is constant ($v \approx$ constant), the enthalpy difference is $h_2 - h_1 \approx v(P_2 - P_1)$, so $w_{pump} = v(P_2 - P_1)$.

Boilers and Vapor Generators: Heat Addition I

- Boiler: A steam-generating heat exchanger where heat from hot combustion gases is transferred to feed water at constant pressure.
- Vapor Generator: General term for devices performing phase transition from liquid to gas via external thermal energy input.
- SFEE Assumptions: No shaft work ($\dot{W}_{shaft} = 0$), negligible kinetic and potential energy changes ($\Delta ke \approx 0$, $\Delta pe \approx 0$).
- Heat Input Equation: $\dot{Q}_{in} = \dot{m}(h_2 - h_1)$ or $q_{in} = h_2 - h_1$. Pressure is assumed constant ($P_1 \approx P_2$).
- Phase changes: Liquid water enters as subcooled liquid and exits as wet, dry saturated, or highly superheated steam ($h_2 \gg h_1$).

Condensers and Heat Exchangers: Heat Rejection I

- Condenser: A shell-and-tube heat exchanger where exhaust vapor rejects heat to cooling water and condenses into a liquid state.
- SFEE Assumptions: No work interaction ($\dot{W}_{shaft} = 0$), negligible kinetic and potential energy changes ($\Delta ke \approx 0$, $\Delta pe \approx 0$).
- Heat Rejection Equation: $\dot{Q}_{out} = \dot{m}(h_1 - h_2)$ or $q_{out} = h_1 - h_2$ where fluid state transitions from vapor to liquid.
- Energy Balance: For two independent streams (hot condensing fluid and cold cooling fluid), $\dot{m}_{hot}(h_{hot,in} - h_{hot,out}) = \dot{m}_{cold}(h_{cold,out} - h_{cold,in})$.
- System Importance: Completes the Rankine steam cycle by condensing working fluid, reducing the specific volume to enable low work input in the pump.

Throttling Devices: Constant Enthalpy Expansion I

- Throttling definition: Irreversible expansion process through a restriction (partially closed valve, orifice, capillary tube, or porous plug) with a large pressure drop.
- SFEE Assumptions: Highly localized and rapid process, meaning it is adiabatic ($q = 0$), no shaft work ($w = 0$), and negligible potential energy change ($\Delta z = 0$).
- Isenthalpic Process: Outlet velocity equals inlet velocity in a steady flow duct, yielding $h_1 = h_2$. Thus, throttling is a constant enthalpy process.
- Joule-Thomson Coefficient: $\mu_{JT} = \left(\frac{\partial T}{\partial P}\right)_h$. If $\mu_{JT} > 0$, temperature drops; if $\mu_{JT} < 0$, temperature rises; for ideal gases, $\mu_{JT} = 0$ (no temperature change).
- Applications: Thermal expansion valves in vapor compression refrigeration cycles, and throttling calorimeters used to determine steam quality.

Simple numerical examples based on the above. I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 2: First Law of Thermodynamics
- Lecture 19: Simple numerical examples based on the above.
- Instructor: Professor of Mechanical Engineering
- Focus: Practical mathematical application of First Law of Thermodynamics equations to closed (non-flow) and open (steady-flow) engineering systems.

Agenda I

- Recap of key governing thermodynamic equations
- Example 1: Isobaric closed-system expansion of an ideal gas
- Diagrammatic representation of constant-pressure heating
- Example 2: Polytropic process (n=1.3) work and heat calculation
- Diagrammatic representation of polytropic expansion on P-V coordinates
- Example 3: Steady Flow Energy Equation (SFEE) applied to a steam turbine
- Example 4: Velocity calculation of gas flowing through an insulated nozzle
- Critical checklist for unit conversion, assumptions, and sign conventions

Governing Equations Recap I

- First Law for Closed Systems (Non-flow): Q - W = Delta U, where Q is net heat transfer, W is boundary work, and Delta U is internal energy change.
- Boundary Work (Non-flow): W = Integral of P*dV. For Constant Pressure (Isobaric): W = P*(V2 - V1). For Polytropic (PV^n = C): W = (P1*V1 - P2*V2)/(n-1).
- First Law for Open Systems (Steady Flow Energy Equation): m_dot * [h1 + (C1^2)/2000 + g*z1/1000] + Q_dot = m_dot * [h2 + (C2^2)/2000 + g*z2/1000] + W_dot_sh.
- Sign Conventions: Heat transfer INTO the system (+Q) and work done BY the system (+W) are positive. Heat rejected (-Q) and work done ON the system (-W) are negative.
- Ideal Gas Relations: Change in internal energy Delta U = m * C_v * Delta T. Change in enthalpy Delta H = m * C_p * Delta T.

Governing Equations Recap II

- Units Warning: Kinetic energy (C^2/2000) and potential energy (g*z/1000) terms must be divided by 1000 to convert Joules/kg to kJ/kg to align with Enthalpy (h).

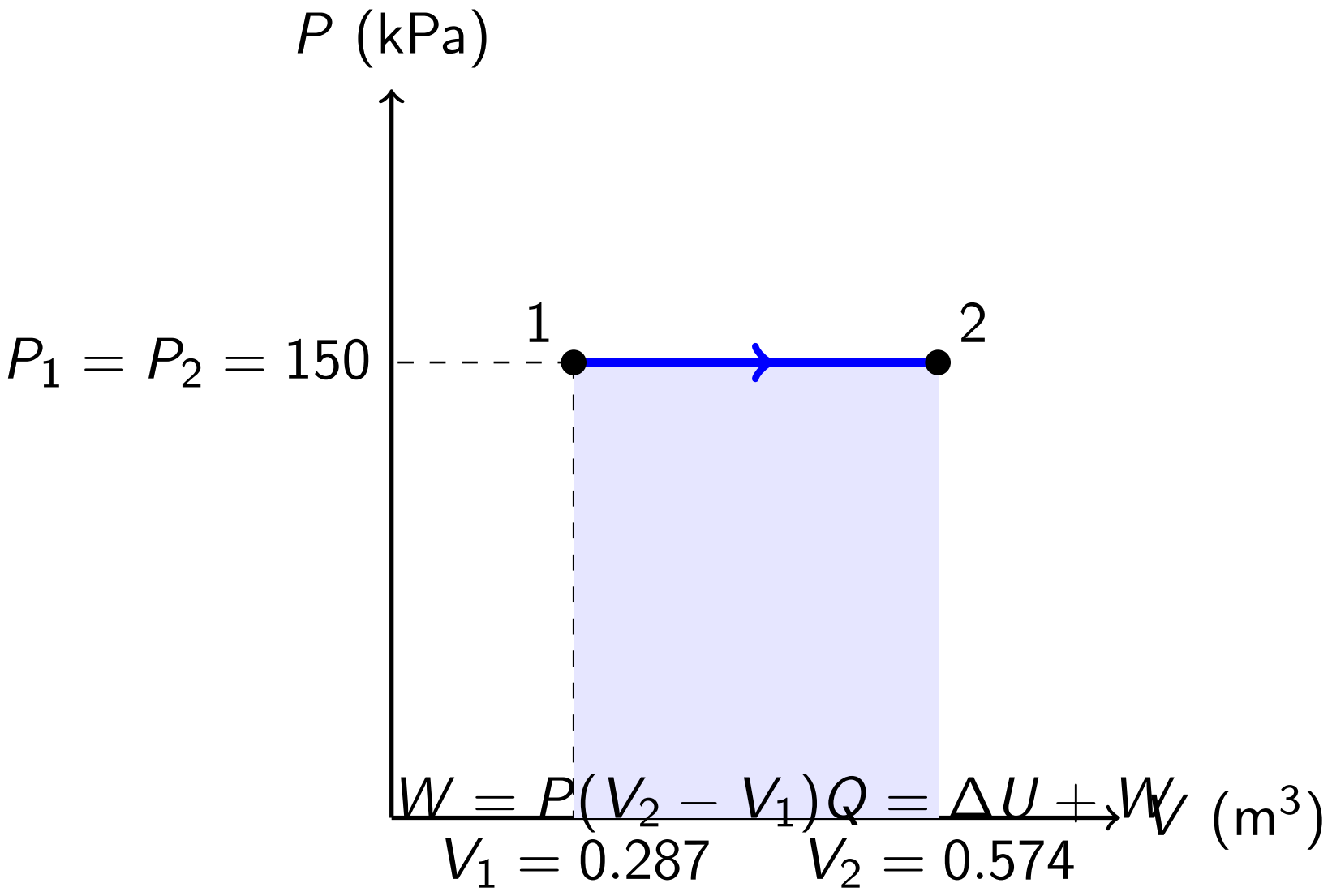
Example 1: Isobaric Gas Heating I

- Problem: 0.5 kg of air at 150 kPa and 300 K is heated at constant pressure in a piston-cylinder device until its volume doubles. Find W, Q, and Delta U.
- Constants for Air: Specific heat at constant pressure C_p = 1.005 kJ/kg-K, Gas constant R = 0.287 kJ/kg-K, Specific heat at constant volume C_v = 0.718 kJ/kg-K.
- Step 1 (Find Volume): Use ideal gas equation to find initial volume V1. V1 = (m * R * T1) / P1 = (0.5 kg * 0.287 kJ/kg-K * 300 K) / 150 kPa = 0.287 m^3. Thus, final volume V2 = 2 * V1 = 0.574 m^3.
- Step 2 (Find final temperature T2): For constant pressure, Charles's Law applies: T2 = T1 * (V2 / V1) = 300 K * 2 = 600 K.

Example 1: Isobaric Gas Heating II

- Step 3 (Boundary Work W): W = P * (V2 - V1) = 150 kPa * (0.574 - 0.287) m^3 = 43.05 kJ (positive work done BY system).
- Step 4 (Heat Transfer Q): Q = m * C_p * (T2 - T1) = 0.5 kg * 1.005 kJ/kg-K * (600 - 300) K = 150.75 kJ (positive heat entering system).
- Step 5 (Internal Energy Change Delta U): Using 1st Law, Delta U = Q - W = 150.75 kJ - 43.05 kJ = 107.70 kJ. (Check: Delta U = m * C_v * (T2 - T1) = 0.5 * 0.718 * 300 = 107.70 kJ).

P-V Diagram: Constant Pressure Expansion



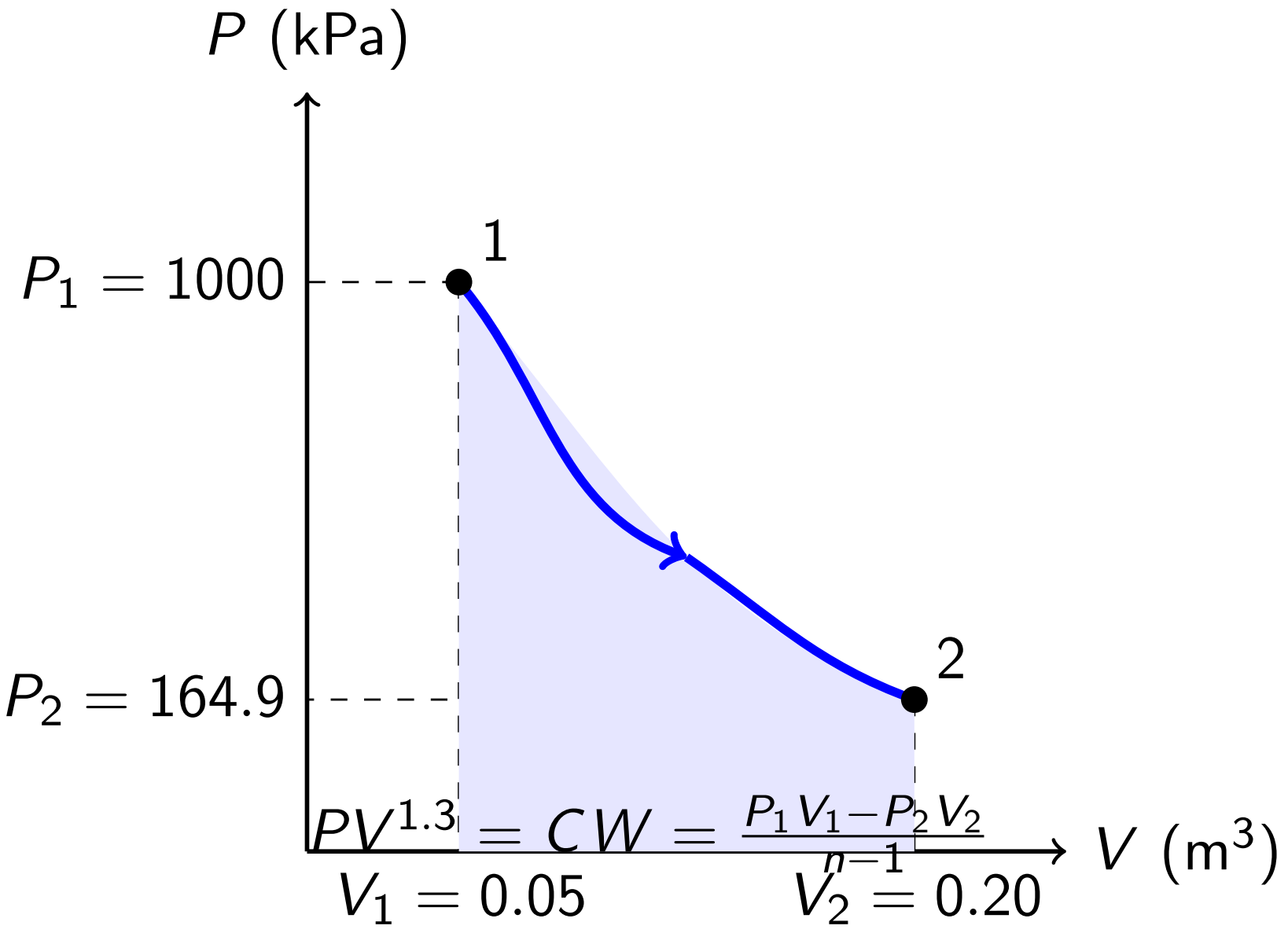
Example 2: Polytropic Expansion I

- Problem: A gas expands polytropically (PV^1.3 = C) inside a piston-cylinder. Initial conditions: P1 = 1.0 MPa, V1 = 0.05 m^3. Final volume V2 = 0.2 m^3. Internal energy decreases by 40 kJ. Find W and Q.
- Given Parameters: P1 = 1000 kPa, V1 = 0.05 m^3, V2 = 0.2 m^3, polytropic index n = 1.3, Delta U = -40 kJ.
- Step 1 (Find P2): Using relation P1*V1^n = P2*V2^n, we get P2 = P1 * (V1 / V2)^n = 1000 kPa * (0.05 / 0.2)^1.3 = 1000 * (0.25)^1.3 = 164.94 kPa.
- Step 2 (Calculate Work W): W = (P1*V1 - P2*V2) / (n - 1) = (1000 kPa * 0.05 m^3 - 164.94 kPa * 0.2 m^3) / (1.3 - 1) = (50 - 32.99) / 0.3 = 56.70 kJ.
- Step 3 (Calculate Heat Transfer Q): Apply First Law, Q - W = Delta U =¿ Q = W + Delta U = 56.70 kJ + (-40 kJ) = 16.70 kJ.

Example 2: Polytropic Expansion II

- Analysis: The positive value of W (56.70 kJ) means the system does work on the surroundings, while the positive Q (16.70 kJ) indicates net heat must be supplied during expansion to satisfy the energy balance.

P-V Diagram: Polytropic Expansion



Example 3: Steam Turbine (Open System) I

- Problem: A steady flow steam turbine receives steam at 1.5 kg/s. Inlet: P1=2 MPa, T1=350 C, h1=3137 kJ/kg, C1=50 m/s, z1=6 m. Outlet: P2=0.1 MPa, quality x2=0.9 (h2=2450 kJ/kg), C2=120 m/s, z2=2 m. Heat loss is 15 kW. Find power output.
- Step 1 (Inlet Energy): Specific kinetic energy KE1 = C1^2/2000 = 50^2/2000 = 1.25 kJ/kg. Specific potential energy PE1 = g*z1/1000 = 9.81 * 6 / 1000 = 0.059 kJ/kg.
- Step 2 (Outlet Energy): Specific kinetic energy KE2 = C2^2/2000 = 120^2/2000 = 7.20 kJ/kg. Specific potential energy PE2 = g*z2/1000 = 9.81 * 2 / 1000 = 0.020 kJ/kg.
- Step 3 (Heat Loss per unit mass): Specific heat transfer q = Q_dot / m_dot = -15 kW / 1.5 kg/s = -10 kJ/kg (negative because heat leaves the control volume).

Example 3: Steam Turbine (Open System) II

- Step 4 (Set up SFEE): h1 + KE1 + PE1 + q = h2 + KE2 + PE2 + w =¿ 3137 + 1.25 + 0.059 - 10 = 2450 + 7.20 + 0.020 + w.
- Step 5 (Solve for Shaft Work and Power): 3128.31 kJ/kg = 2457.22 kJ/kg + w =¿ w = 671.09 kJ/kg. Power output W_dot = m_dot * w = 1.5 kg/s * 671.09 kJ/kg = 1006.64 kW.

Example 4: Adiabatic Air Nozzle I

- Problem: Air enters an insulated nozzle at 300 kPa, 470 K with a velocity of 30 m/s. It exits the nozzle at 100 kPa and 350 K. Calculate the exit velocity of the air.
- Assumptions: Steady flow, adiabatic operation (q = 0), no shaft work (w = 0), and potential energy changes are negligible (Delta z = 0). Specific heat for air C_p = 1.005 kJ/kg-K.
- Step 1 (Simplify SFEE): The general energy equation simplifies under these assumptions to: h1 + C1^2/2 = h2 + C2^2/2.
- Step 2 (Express enthalpy change): For an ideal gas, h1 - h2 = C_p * (T1 - T2) = 1.005 kJ/kg-K * (470 - 350) K = 120.6 kJ/kg.
- Step 3 (Convert to standard SI units): Velocity calculations require enthalpy in J/kg. h1 - h2 = 120.6 kJ/kg * 1000 J/kJ = 120,600 J/kg.

Example 4: Adiabatic Air Nozzle II

- Step 4 (Solve for exit velocity C2): C2^2 = 2 * (h1 - h2) + C1^2 = 2 * 120,600 J/kg + (30 m/s)^2 = 241,200 + 900 = 242,100 m^2/s^2. Thus, C2 = sqrt(242,100) = 492.04 m/s.

Summary & Problem-Solving Steps I

- System Verification: Clearly identify if the system is closed (constant mass) or open (mass flows across boundary) to choose the correct version of the 1st Law.
- Process Boundary Conditions: Recognize process types: Isobaric (P=C), Isothermal (T=C), Adiabatic (Q=0), Isochoric (V=C), or Polytropic (PV^n=C) to calculate work correctly.
- Property Table usage vs Ideal Gas laws: Use steam/refrigerant tables for multi-phase substances, and ideal gas state equations (Pv=RT) for air and gases.
- Check Units: Always match specific kinetic energy (divide C^2 by 2000) and specific potential energy (divide g*z by 1000) to keep units in kJ/kg.
- Coordinate and Diagram Verification: Check if work and heat values logically match diagrammatic areas (area under curve on P-V diagram equals work done).

Limitations of the first law of thermodynamics. I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 3: Second Law of Thermodynamics
- Lecture 20: Limitations of the First Law of Thermodynamics
- Instructor: Mechanical Engineering Professor
- Overview: Understanding why the conservation of energy principle alone is insufficient to predict the feasibility or direction of thermodynamic processes.

Lecture Agenda I

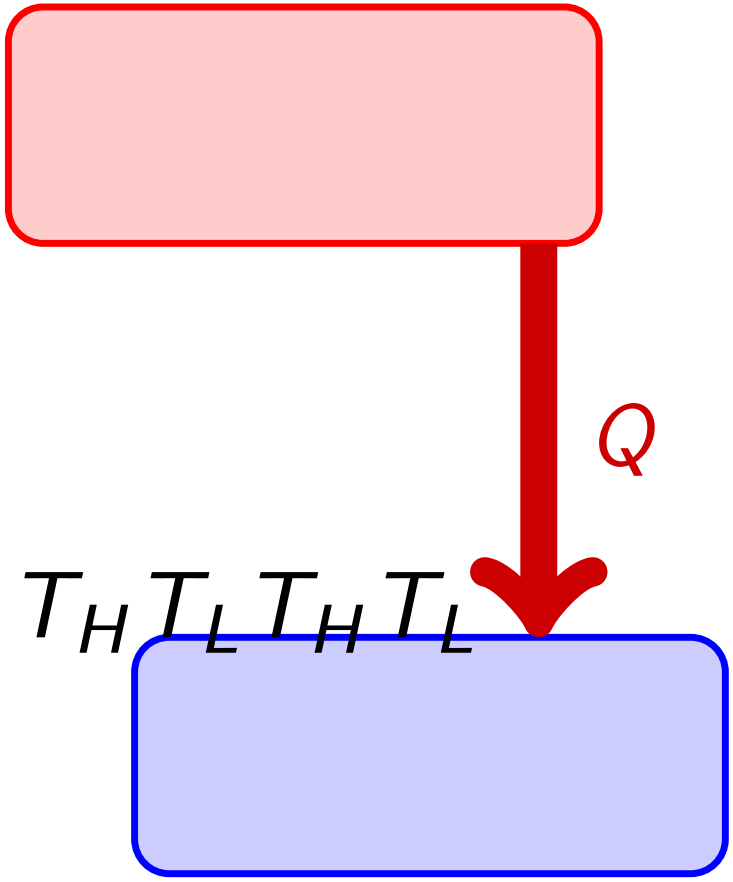
- Recap of the First Law of Thermodynamics and its conservation principle
- The concept of process directionality and spontaneity in nature
- The quality of energy: distinguishing between heat and work
- Physical examples demonstrating the limitations of the First Law
- Introduction to the need for the Second Law of Thermodynamics

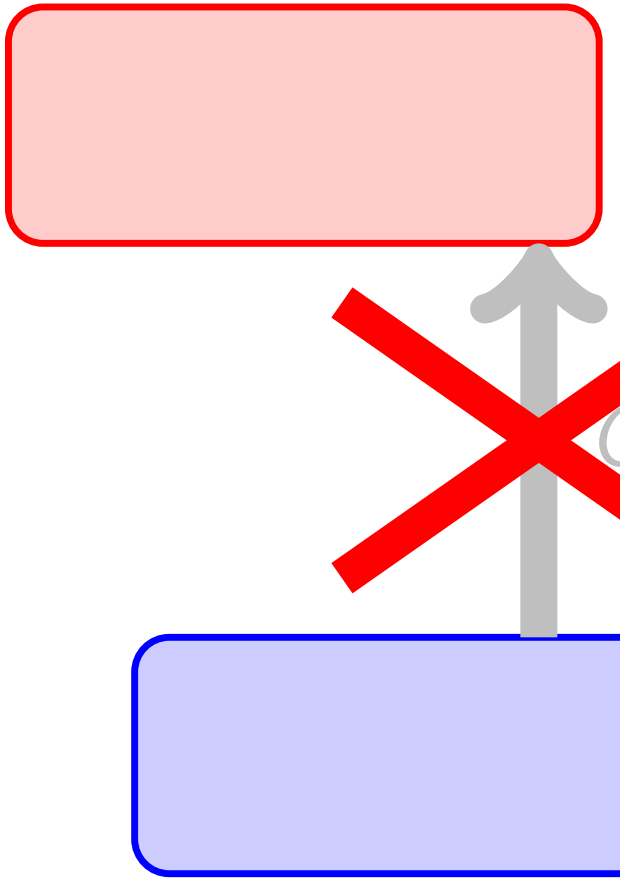
Recap: The First Law of Thermodynamics I

- The First Law is a statement of the conservation of energy: energy can change forms but cannot be created or destroyed.
- For a closed system undergoing a thermodynamic cycle, the net heat transfer equals the net work transfer: $\oint \delta Q = \oint \delta W$.

- For a non-cyclic process in a closed system, the energy balance is given by: Δ U = Q - W.
- Major Limitation: The First Law only requires that energy balances out. It does not place any restriction on the direction of energy transfer.
- According to the First Law, any process is theoretically possible as long as the quantity of energy remains conserved.

Spontaneous Direction of Heat Flow





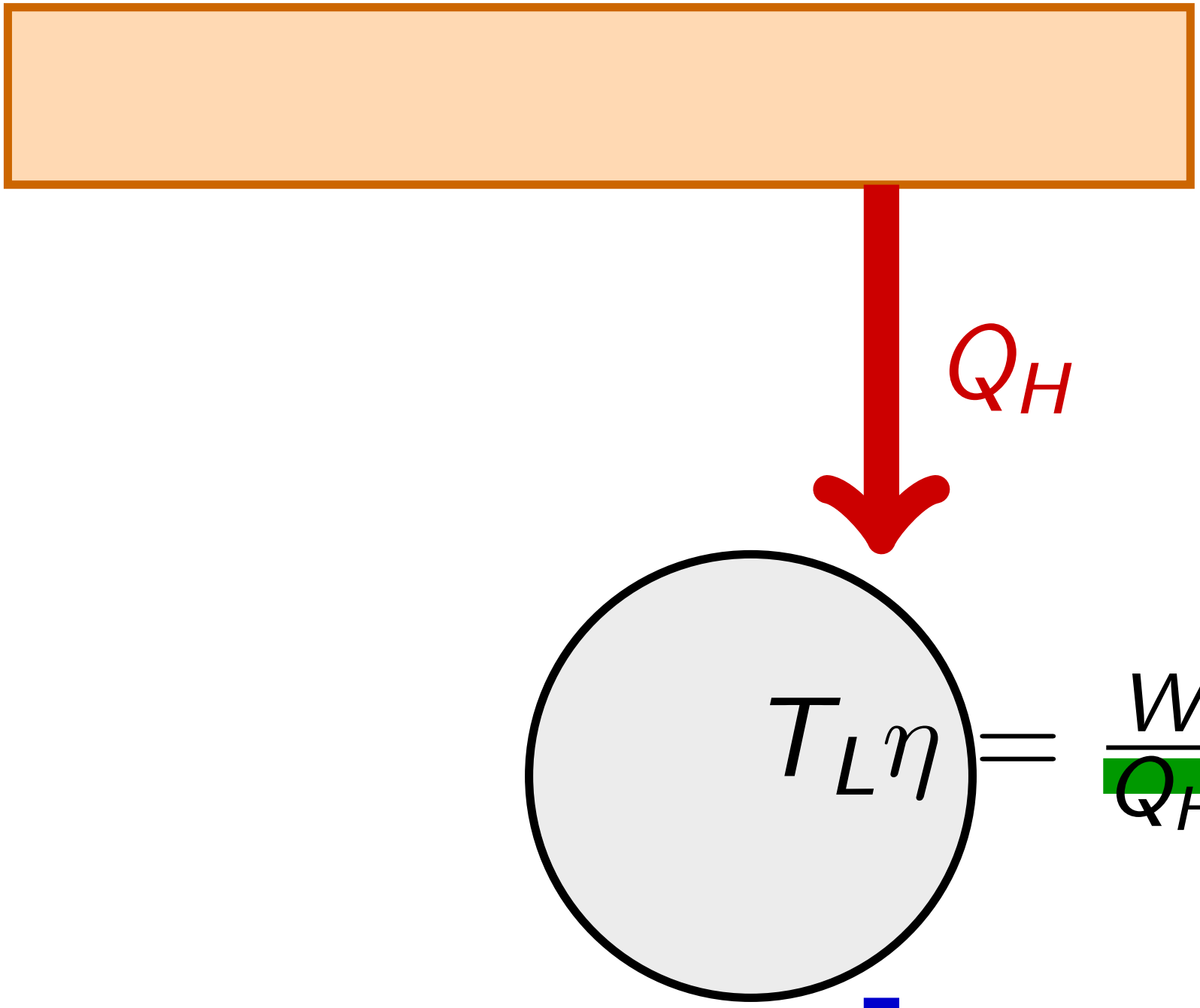
Limitation 1: The Direction of a Process I

- A hot cup of coffee left in a cooler room spontaneously cools down. The reverse process (the coffee extracting heat from the room to boil itself) never happens, even though it would conserve energy.
- A pressurized gas expands spontaneously into a vacuum. The gas will never spontaneously compress back into its original container volume.
- A wheel turning in a fluid does work, raising the fluid's temperature. Supplying heat to the fluid will not spontaneously make the wheel rotate in reverse.
- These examples show that natural physical processes have a preferred direction that the First Law cannot predict.

Limitation 2: The Quality of Energy I

- The First Law treats heat and work as completely equivalent forms of energy (1 J of heat is equal to 1 J of work).
- In practice, Work is high-grade energy: it can be converted entirely and easily into other forms of energy (like heat or potential energy).
- Heat is low-grade energy: it cannot be fully converted into work on a continuous basis, even under ideal frictionless conditions.
- Any cyclic heat engine converting heat to work must reject a portion of heat to a low-temperature reservoir, limiting efficiency.
- The First Law fails to recognize the degradation of energy quality during real thermodynamic processes.

Thermodynamic Limit of Heat Conversion



Detailed Examples of First Law Inadequacy I

- Paddle-wheel work input: Work input (W) is easily converted 100% to heat, increasing internal energy. However, adding heat back will not spontaneously turn the paddle wheel to yield work.
- Resistor heating: Passing electrical current (I) through a resistor (R) generates heat (I^2R) at 100% efficiency. Heating the resistor will never generate an electric current in the wire.
- Free expansion: Gas expanding into an evacuated chamber results in zero temperature change ($Q = 0, W = 0, \Delta U = 0$). Yet, the gas never spontaneously compresses back into the original chamber.

The Path to the Second Law I

- Because the First Law does not restrict process direction, a new law is required to define thermodynamic feasibility.
- The Second Law of Thermodynamics establishes the criteria for whether a process can occur.
- It introduces the concept of Entropy (S) to mathematically determine the direction of spontaneous processes.
- It introduces Exergy (available energy) to measure the quality and maximum work potential of an energy source.
- It leads to the Kelvin-Planck and Clausius statements, setting limits on heat engine and heat pump performance.

Summary of First Law Limitations I

- Energy conservation is a necessary but not a sufficient condition for a process to take place.
- Real processes proceed spontaneously in a direction that increases the total entropy of the universe ($dS_{\text{total}} \geq 0$).
- Energy has quality as well as quantity; work is a higher-grade energy than heat.
- Thermodynamic cycles converting heat to work must reject heat to a low-temperature sink, making 100% efficiency impossible.
- Next Lecture: Formulation of the Second Law of Thermodynamics and the Kelvin-Planck statement.

Concept of heat source, heat sink, heat engine, heat pump, refrigerator, reversible process and irreversible process. I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 3: Second Law of Thermodynamics
- Lecture 21: Thermal Reservoirs, Cyclic Devices, and Reversibility
- Department of Mechanical Engineering

Agenda I

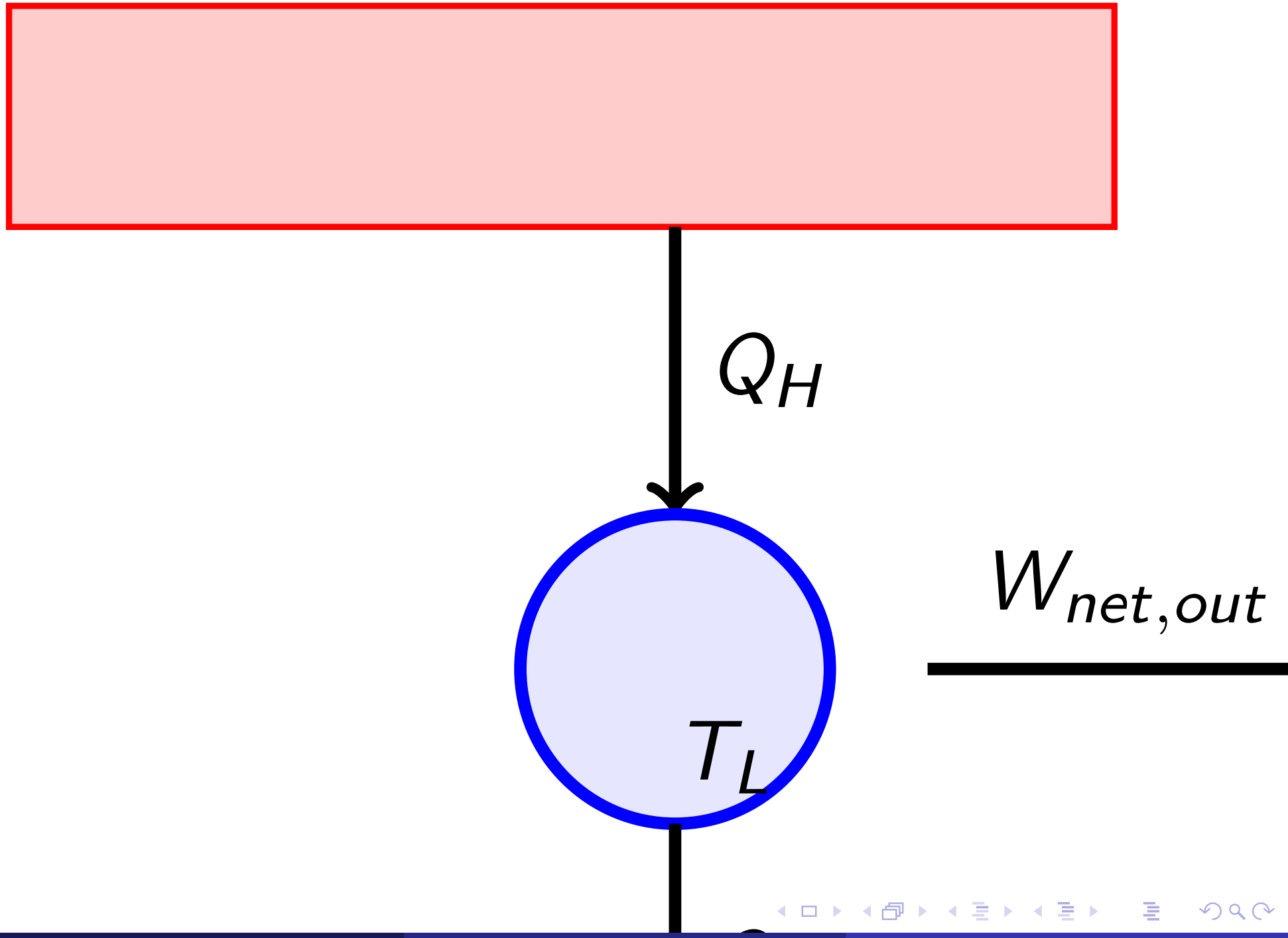
- Thermal Energy Reservoirs: Heat Sources and Heat Sinks
- Heat Engines and Thermal Efficiency (η_th)
- Refrigerators, Heat Pumps, and Coefficient of Performance (COP)
- Thermodynamic Reversibility: Concepts and Idealizations
- Irreversible Processes and Common Sources of Irreversibility
- Internal vs. External Reversibility

Thermal Energy Reservoirs: Heat Source and Heat Sink I

- A thermal energy reservoir is a hypothetical body with a relatively large thermal energy capacity (mass $\times specificheat)thatcanabsorborsupplyfiniteamountsofheatwithoutu$

- Heat Source: A thermal reservoir that supplies heat to a thermodynamic system (e.g., solar collectors, industrial furnaces, geothermal wells).
- Heat Sink: A thermal reservoir that absorbs rejected heat from a system (e.g., atmospheric air, oceans, large rivers).
- Mathematical limit: As heat transfer Q approaches infinity, the temperature change Δ T of the reservoir approaches zero.
- These reservoirs serve as idealized thermal boundaries for energy conversion systems, facilitating the application of the Second Law.

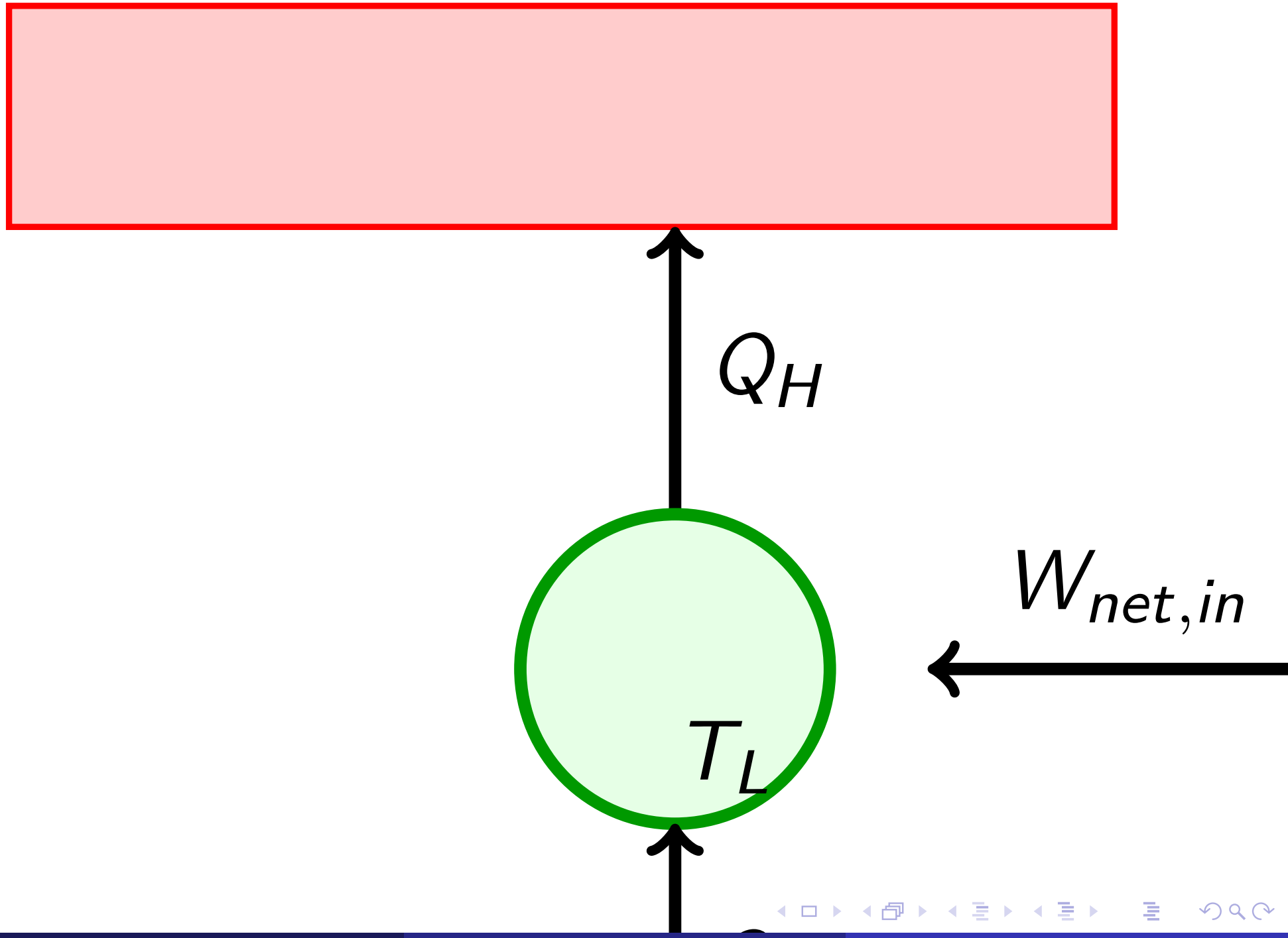
Schematic Representation of a Heat Engine



Heat Engines and Thermal Efficiency I

- Heat engines are cyclic thermodynamic devices designed to convert heat into useful work output.
- Operating characteristics: 1. They receive heat Q_H from a high-temperature source at T_H; 2. They convert a part of this heat to work W_net,out; 3. They reject the remaining heat Q_L to a low-temperature sink at T_L; 4. They operate on a cycle.
- First Law Conservation of Energy for a closed power cycle: W_net,out = Q_H - Q_L.
- Thermal Efficiency (η_th) is defined as: η_th = Net Work Output$\frac{}{\text{Total Heat Input}} = \frac{W_net,out}{Q_H} = 1 - \frac{Q_L}{Q_H}$.
- By the Kelvin-Planck statement of the Second Law, no heat engine can convert all input heat into work; thus, η_th ¡ 100%.

Schematic of Refrigerators and Heat Pumps



Refrigerators, Heat Pumps, and Coefficient of Performance (COP) I

- Refrigerators and heat pumps are cyclic devices that transfer thermal energy from a low-temperature medium to a high-temperature medium, requiring net work input W_net,in.
- Refrigerator objective: Maintain a cold space at T_L by removing heat Q_L.
- COP of Refrigerator: COP_R = Desired Output$\frac{}{\text{Required Input}} = \frac{Q_L}{W_net,in} = \frac{Q_L}{Q_H - Q_L}$.
- Heat Pump objective: Maintain a warm space at T_H by supplying heat Q_H.

Refrigerators, Heat Pumps, and Coefficient of Performance (COP) II

- COP of Heat Pump: $\text{COP\_HP} = \frac{\text{Desired Output}}{\text{Required Input}} = \frac{Q\_H}{W\_net,in} = \frac{Q\_H}{Q\_H - Q\_L}$.
- For identical operating conditions: $\text{COP\_HP} = \text{COP\_R} + 1$. Note that COP values can be, and typically are, greater than unity.

Concept of Reversible Processes I

- A reversible process is defined as a thermodynamic process that can be reversed without leaving any trace (permanent changes) on the system and its surroundings.
- Characteristics: 1. Both system and surroundings return to their exact initial states; 2. The process proceeds through a continuous series of equilibrium states.
- Reversible processes represent idealized limits. Actual physical processes can approach them but never achieve perfect reversibility.
- A quasi-equilibrium process, which occurs infinitesimally slowly, is the thermodynamic model used to analyze reversible processes.
- Engineers use reversible processes as standard benchmarks to evaluate the efficiency of real processes and minimize losses.

Irreversible Processes and Sources of Irreversibility I

- An irreversible process is any process that does not meet the strict definition of reversibility. Once it occurs, restoring the system requires changes in the surroundings.
- Friction: Mechanical work is dissipated into thermal energy, which cannot be converted back to work spontaneously.
- Unrestrained Expansion: The expansion of a gas into a vacuum or lower-pressure region without producing work; restoring the state requires external compressor work.
- Heat Transfer through a Finite Temperature Difference: Heat flows spontaneously down a finite temperature gradient. Reversing this flow requires a refrigeration cycle and work input.
- Other sources: Chemical reactions, electrical resistance, mixing of different substances, and inelastic deformation.

Internal and External Reversibility I

- Internal Reversibility: No irreversibilities occur within the boundaries of the system during the process (e.g., quasi-equilibrium expansion or compression without internal friction).
- External Reversibility: No irreversibilities occur outside the system boundaries (i.e., in the surrounding environment) during the process.
- Example: Heat transfer between a thermal reservoir and a system is externally reversible if the system boundary temperature matches the reservoir temperature ($dT \rightarrow 0$).

- A process is defined as totally reversible (or simply reversible) if it is both internally and externally reversible.

Internal and External Reversibility II

- Minimizing internal and external irreversibilities is critical for maximizing work output in engines and minimizing work input in refrigeration.

Statement of the second law of thermodynamics: I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 3: Second Law of Thermodynamics
- Lecture 22: Statement of the Second Law of Thermodynamics
- Instructor: Professor of Mechanical Engineering

Agenda I

- Limitations of the First Law of Thermodynamics
- Thermal Energy Reservoirs (Source and Sink)
- Heat Engines and Thermal Efficiency
- The Kelvin-Planck Statement of the Second Law
- Refrigerators, Heat Pumps, and Coefficient of Performance (COP)
- The Clausius Statement of the Second Law
- Equivalence of the Kelvin-Planck and Clausius Statements
- Perpetual Motion Machines (PMM1 and PMM2)

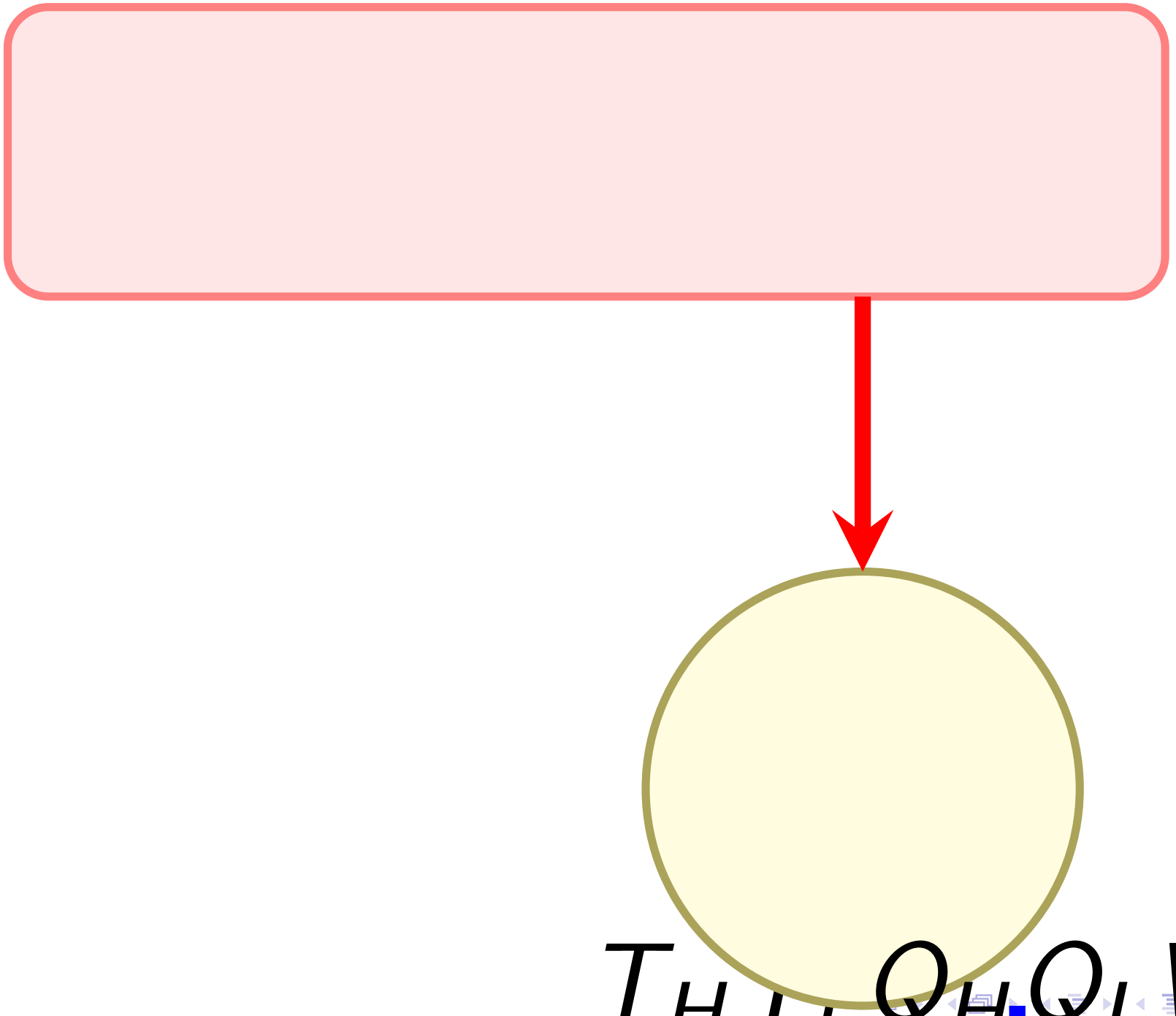
Limitations of the First Law I

- The First Law of Thermodynamics asserts energy conservation: energy can neither be created nor destroyed, only transformed.
- However, the First Law does not restrict the direction of a process; it permits heat to flow spontaneously from a cold body to a hot body.
- Real-world processes have a natural direction: a hot cup of coffee cools down in a room, but a cold cup never spontaneously heats up by drawing energy from the room.
- A process cannot occur unless it satisfies both the First and Second Laws of Thermodynamics.
- The Second Law identifies the direction of processes and defines the quality of energy (e.g., heat is a low-grade energy compared to work).

Thermal Energy Reservoirs I

- A Thermal Energy Reservoir (TER) is a hypothetical body with a large thermal energy capacity (mass x specific heat) that can absorb or supply finite amounts of heat without undergoing any change in temperature.
- Source: A reservoir that supplies thermal energy (e.g., solar radiation, geothermal wells, combustion furnaces).
- Sink: A reservoir that absorbs thermal energy (e.g., oceans, rivers, the atmospheric air).
- In practice, large bodies of water or the atmosphere behave as TERs due to their massive thermal inertia.
- TERs are crucial for analyzing heat engines, refrigerators, and heat pumps.

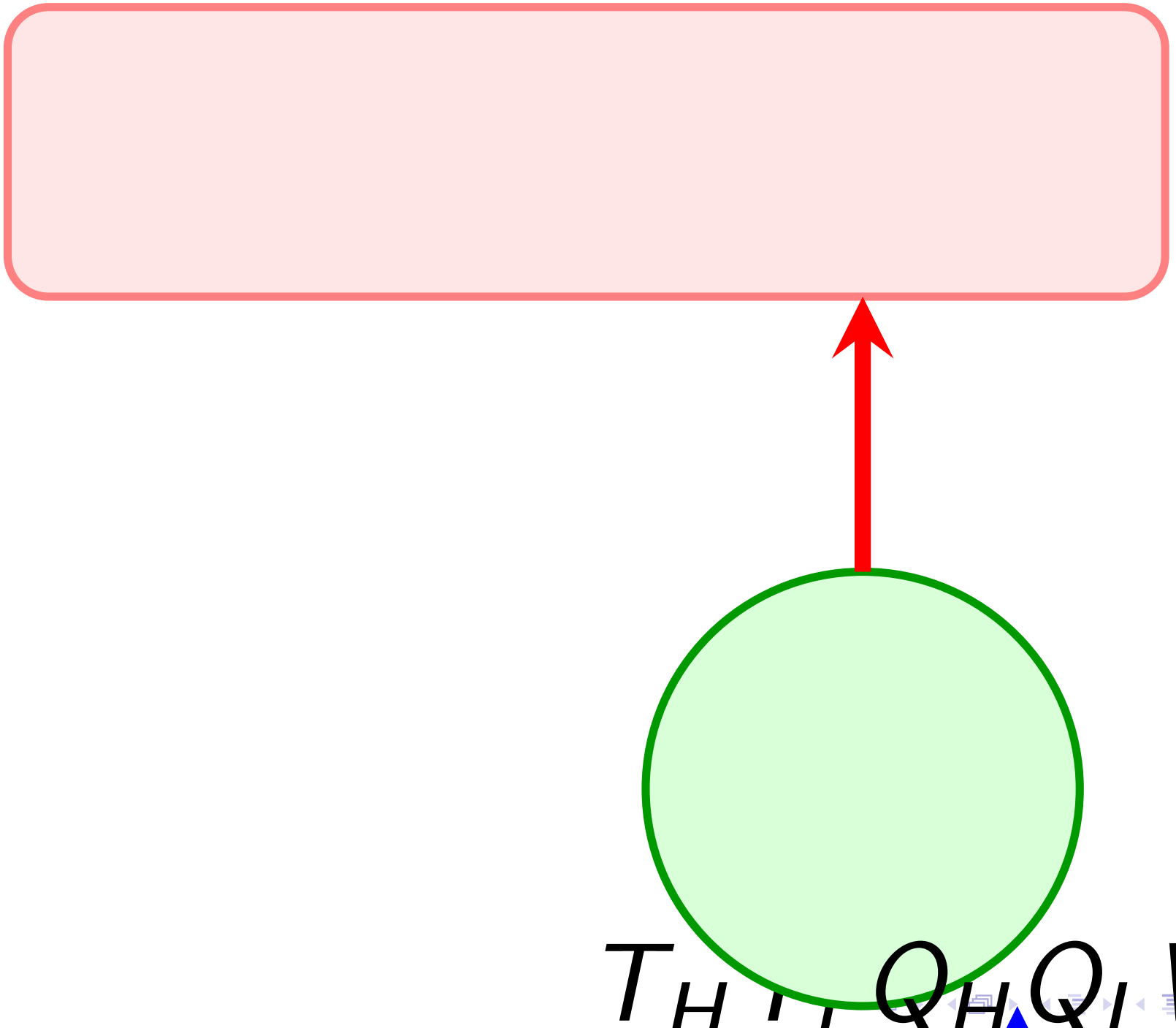
Schematic of a Heat Engine



The Kelvin-Planck Statement I

- Statement: It is impossible for any device that operates on a thermodynamic cycle to receive heat from a single reservoir and produce a net amount of work.
- This means that no heat engine can have a thermal efficiency of 100 percent.
- To continuously convert heat to work, a heat engine must exchange heat with a low-temperature sink as well as a high-temperature source.
- The rejected heat (Q_L) is a necessary 'waste' required by the Second Law, not just a practical design defect.
- Thermal efficiency is defined as: $\eta_{th} = W_{net}/Q_H = 1 - Q_L/Q_H$, which is strictly less than 1.

Schematic of a Refrigerator / Heat Pump



The Clausius Statement I

- Statement: It is impossible to construct a device that operates in a cycle and produces no other effect than the transfer of heat from a lower-temperature body to a higher-temperature body.
- In simple terms, heat cannot spontaneously flow from a cooler body to a warmer body without work input.
- A refrigerator or heat pump requires a net work input (W_{in}) from an external source (like an electric motor) to accomplish this transfer.
- The Coefficient of Performance (COP) measures the effectiveness of these devices.
- $\text{COP}_R = Q_L/W_{in}$ for refrigerators; $\text{COP}_{HP} = Q_H/W_{in}$ for heat pumps. Because $Q_H = Q_L + W_{in}$, we have $\text{COP}_{HP} = \text{COP}_R + 1$.

Equivalence of the Two Statements I

- The Kelvin-Planck and Clausius statements of the Second Law are equivalent; violating one statement automatically violates the other.
- Proof 1 (Violating Clausius leads to violating Kelvin-Planck): Consider a device that transfers heat Q_L from cold to hot reservoir without work input (violates Clausius). Pair it with a heat engine taking Q_H from hot, rejecting Q_L to cold, and producing work $W = Q_H - Q_L$.
- The combined system draws net heat $Q_H - Q_L$ from the hot reservoir and converts it entirely into work, with zero net heat exchange with the cold reservoir. This violates the Kelvin-Planck statement.

Equivalence of the Two Statements II

- Proof 2 (Violating Kelvin-Planck leads to violating Clausius): Consider a heat engine that takes heat Q_H from a single reservoir and converts it entirely into work $W = Q_H$ (violates Kelvin-Planck). Let this work run a refrigerator transferring Q_L from cold to hot reservoir.
- The combined system transfers heat Q_L from the cold reservoir to the hot reservoir with no net external work input. This violates the Clausius statement.

Perpetual Motion Machines I

- A Perpetual Motion Machine is any device that violates a thermodynamic law.
- Perpetual Motion Machine of the First Kind (PMM1): A device that produces work without any energy input, violating the First Law of Thermodynamics (energy conservation).
- Perpetual Motion Machine of the Second Kind (PMM2): A device that operates in a cycle and performs work while exchanging heat with a single thermal reservoir, violating the Second Law (Kelvin-Planck statement).
- PMM2 could potentially have an efficiency of 100%, and even though it complies with energy conservation (First Law), it is physically impossible.
- Many inventors have claimed to build PMM1s or PMM2s, but none have succeeded because they run counter to fundamental laws of nature.

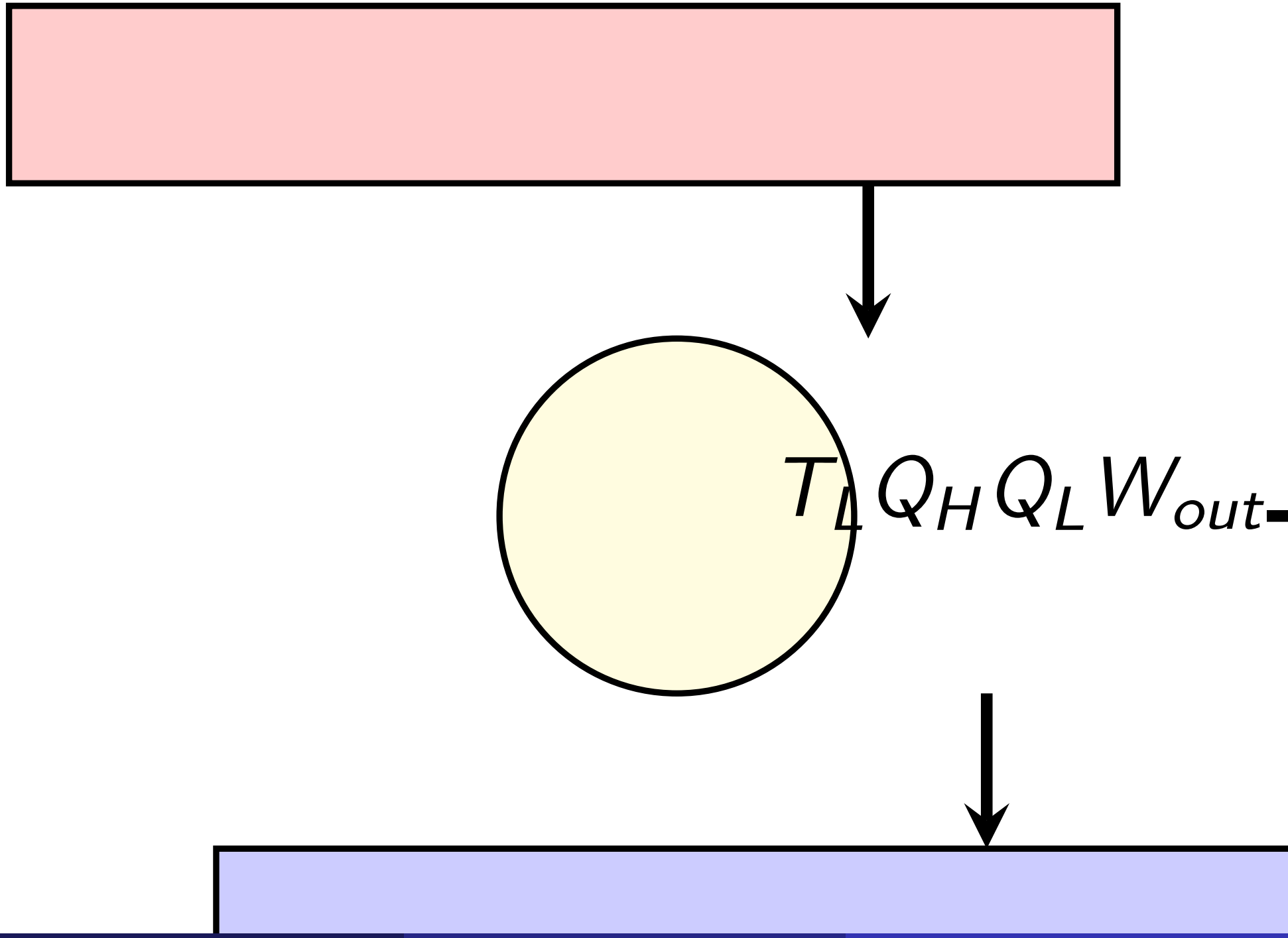
Applications of the second law of thermodynamics. I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 3: Second Law of Thermodynamics
- Lecture 23: Applications of the Second Law of Thermodynamics
- Instructor: Professor of Mechanical Engineering

Agenda I

- Thermodynamic performance limits of heat engines
- Refrigerators and heat pumps operating cycles and COP boundaries
- The Carnot principles and thermodynamic temperature scale derivation
- The Clausius inequality and its role in identifying cycle feasibility
- Entropy balances for control volumes under steady and transient states
- Isentropic efficiencies of common engineering devices (turbines, compressors, nozzles)

Schematic of a Heat Engine Cycle



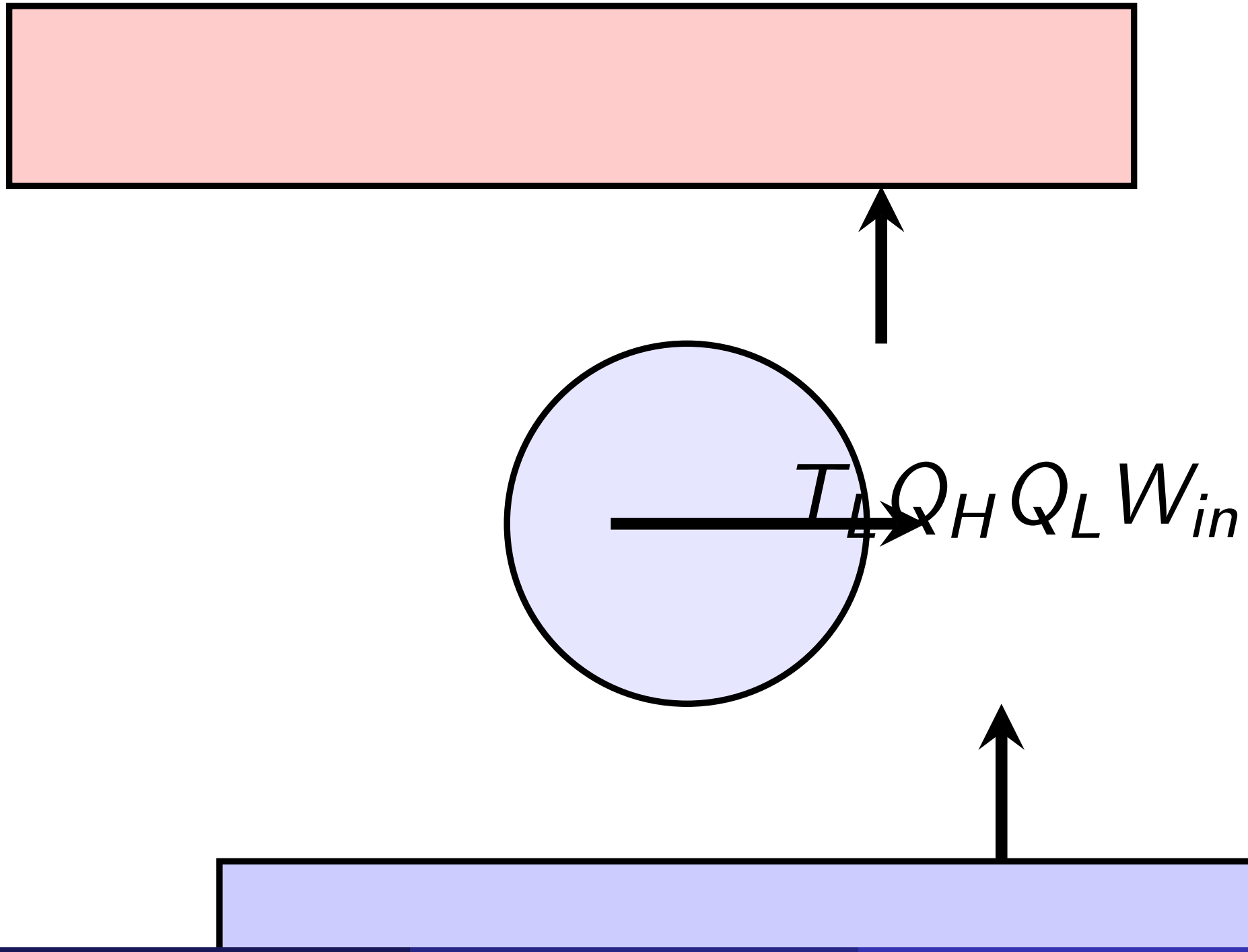
Thermodynamic Performance of Heat Engines I

- A heat engine is a cyclically operating system designed to convert thermal energy from a high-temperature source into mechanical work, rejecting waste heat to a low-temperature sink.
- First Law Formulation (Energy Conservation): The net work output equals the net heat transfer: $W_{net,out} = Q_H - Q_L$, where Q_H is heat absorbed and Q_L is heat rejected.
- Thermal Efficiency (η_{th}): Defined as the ratio of net work output to the heat input: $\eta_{th} = W_{net,out}/Q_H = 1 - Q_L/Q_H$.
- Second Law Limitation (Kelvin-Planck Statement): No heat engine can have a thermal efficiency of 100% ($\eta_{th} < 1$). Some heat Q_L must always be rejected to a colder reservoir.

Thermodynamic Performance of Heat Engines II

- Carnot Efficiency Limit: For a completely reversible engine operating between reservoirs at T_H and T_L, the efficiency is maximum: $\eta_{th,max} = 1 - T_L/T_H$, where temperatures are in Kelvin.

Schematic of Refrigerator / Heat Pump



Performance Metrics: Coefficient of Performance (COP) I

- Refrigerators and heat pumps are cyclic devices designed to transfer heat from a low-temperature region to a high-temperature region by consuming external work input.
- Coefficient of Performance for Refrigerator (COP_R): Defined as the desired cooling effect divided by the work input: $COP_R = Q_L/W_{in} = Q_L/(Q_H - Q_L)$.
- Coefficient of Performance for Heat Pump (COP_{HP}): Defined as the desired heating effect divided by the work input: $COP_{HP} = Q_H/W_{in} = Q_H/(Q_H - Q_L)$.
- Thermodynamic Relationship: The COP of a heat pump is always greater than that of a refrigerator operating between the same limits by exactly 1: $COP_{HP} = COP_R + 1$.

Performance Metrics: Coefficient of Performance (COP) II

- Carnot (Reversible) Limits: $COP_{R,max} = T_L/(T_H - T_L)$ and $COP_{HP,max} = T_H/(T_H - T_L)$, indicating that performance degrades as the temperature span $(T_H - T_L)$ increases.

The Carnot Principles I

- The Carnot principles establish absolute limits on the performance of all cyclic heat engines, refrigerators, and heat pumps.
- Principle 1: The efficiency of an irreversible heat engine is always less than the efficiency of a totally reversible engine operating between the same two reservoirs.
- Principle 2: The efficiencies of all totally reversible heat engines operating between the same two reservoirs are the same, regardless of the working fluid used.
- These principles are direct consequences of the Second Law of Thermodynamics (demonstrated by assuming a more efficient engine and showing it violates Kelvin-Planck).

The Carnot Principles II

- Thermodynamic Temperature Scale: The second principle implies that for reversible cycles, $Q_H/Q_L = T_H/T_L$, defining an absolute temperature scale independent of substance properties.

The Clausius Inequality I

- The Clausius Inequality is a quantitative mathematical formulation of the second law for any thermodynamic cycle: $\oint \delta Q/T \leq 0$.
- Reversible Cycles: The cyclic integral is exactly zero, $\oint (\delta Q/T)_{rev} = 0$, which allows the definition of entropy (S) as a state property ($dS = (\delta Q/T)_{rev}$).
- Irreversible Cycles: The cyclic integral is strictly less than zero, $\oint \delta Q/T < 0$, due to internal and external irreversibilities like friction or finite temperature differences.
- Entropy Generation (S_{gen}): For any process, the entropy change is $dS = \delta Q/T + dS_{gen}$, where $dS_{gen} \geq 0$ is the entropy generated by irreversibilities.
- Significance: The inequality provides a criterion for the feasibility of cycles. If $\oint \delta Q/T > 0$, the proposed cycle is physically impossible.

Entropy Balance and Second Law for Control Volumes I

- Engineering systems (turbines, compressors, heat exchangers) are analyzed as open systems (control volumes) where mass flows across the boundaries.
- Entropy Balance Equation: The rate of change of entropy within the control volume equals the rate of entropy transfer via heat and mass, plus the rate of entropy generation.
- Mathematical Form:
  $dS_{cv}/dt = \sum(\dot{Q}_j/T_j) + \sum \dot{m}_{in} s_{in} - \sum \dot{m}_{out} s_{out} + \dot{S}_{gen}$.
- Steady-Flow Devices: For steady-state ($dS_{cv}/dt = 0$) with one inlet and one outlet: $s_{out} - s_{in} = \sum(q_j/T_j) + s_{gen}$.
- Adiabatic Steady-Flow Devices: Since $q_j = 0$, the entropy change must satisfy $s_{out} \geq s_{in}$, with equality holding only for internally reversible (isentropic) processes.

Isentropic Efficiencies of Steady-Flow Devices I

- Isentropic efficiency compares the actual performance of a device to its performance under idealized, internally reversible (isentropic) conditions.
- Turbines (Work-Producing): Measures actual work output relative to the maximum possible isentropic work output: $\eta_T = W_a/W_s \approx (h_1 - h_{2a})/(h_1 - h_{2s})$.
- Compressors & Pumps (Work-Consuming): Measures minimum isentropic work required relative to the actual work input: $\eta_C = W_s/W_a \approx (h_{2s} - h_1)/(h_{2a} - h_1)$.
- Nozzles (Kinetic Energy-Producing): Ratio of the actual kinetic energy at the exit to the exit kinetic energy under isentropic expansion: $\eta_N = V_{2a}^2/V_{2s}^2 \approx (h_1 - h_{2a})/(h_1 - h_{2s})$.
- Engineering Goal: Minimize entropy generation (s_{gen}) in these devices through design optimizations (e.g., blade aerodynamics, low friction, minimized heat loss).

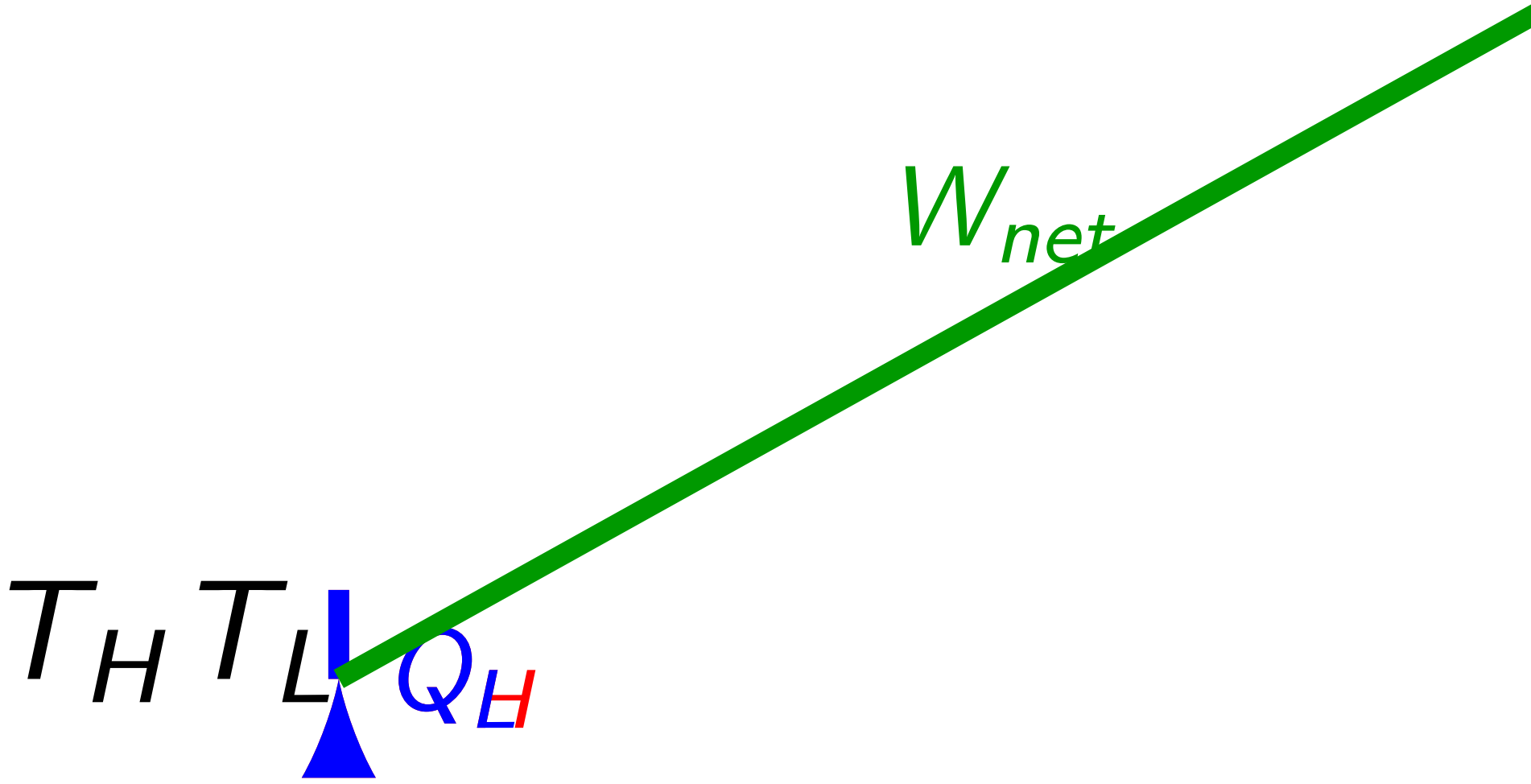
Simple numerical on thermal efficiency and Coefficient of Performance (C.O.P.). I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 3: Second Law of Thermodynamics
- Topics: Quantitative analysis of Heat Engines, Refrigerators, and Heat Pumps
- Key concepts: Thermal Efficiency (η_th) and Coefficient of Performance (C.O.P.)

Lecture Agenda I

- Core Thermodynamic Formulations: Efficiency vs. C.O.P.
- Visualizing Power and Refrigeration Cycles (Schematic Diagrams)
- Step-by-Step Analysis of Heat Engine Thermal Efficiency
- Solving Coefficient of Performance (C.O.P.) for Refrigerator Systems
- Solving Coefficient of Performance (C.O.P.) for Heat Pump Systems
- Exploring the Mathematical Relationship: $\text{COP\_HP} = \text{COP\_Ref} + 1$
- Evaluating Practical Cycles against Carnot Thermodynamic Limits

Heat Engine Schematic Diagram



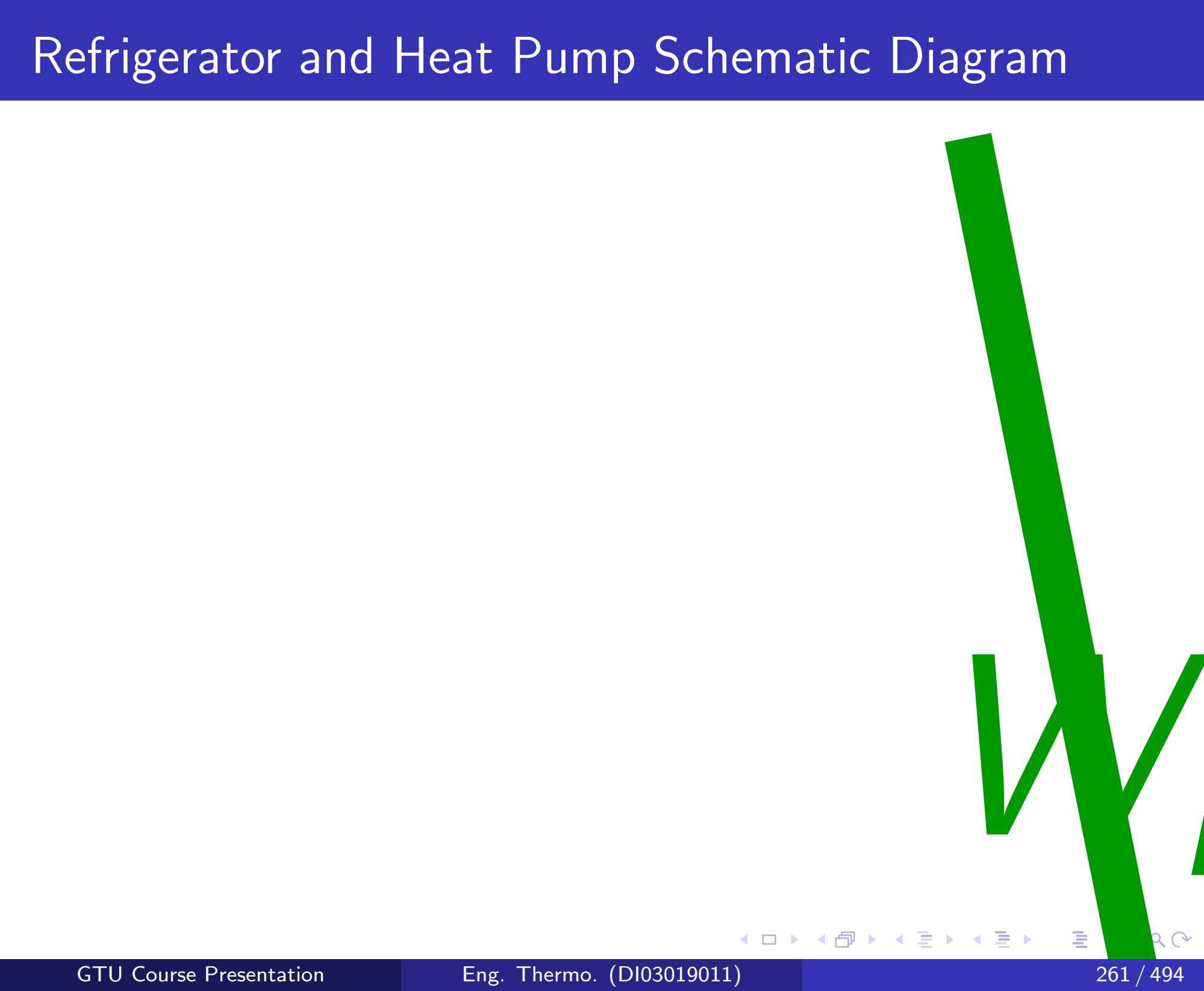
Understanding Thermal Efficiency (η_th) I

- A heat engine is a device operating in a thermodynamic cycle that converts heat from a high-temperature source into net work output.
- Thermal efficiency (η_th) is a measure of performance: η_th = (Net Work Output) / (Total Heat Input) = W_net / Q_H
- Applying the First Law of Thermodynamics to the cycle: W_net = Q_H - Q_L, where Q_H is heat input and Q_L is heat rejected.
- Combining equations: η_th = (Q_H - Q_L) / Q_H = 1 - (Q_L / Q_H). Efficiency is always less than 1 (100%) due to the Second Law.
- For a reversible Carnot engine, efficiency depends only on temperatures: η_th, Carnot = 1 - (T_L / T_H) [Note: T must be in Kelvin].

Numerical Problem 1: Heat Engine Analysis I

- Problem: A heat engine receives 900 kJ of heat per cycle from a source at 800 °C and rejects 540 kJ to a sink at 25 °C. Find: (a) Net work output, (b) Thermal efficiency, (c) Carnot efficiency, and (d) Feasibility.
- Part (a): Net work output W_net = Q_H - Q_L = 900 kJ - 540 kJ = 360 kJ per cycle.
- Part (b): Actual thermal efficiency η_th = W_net / Q_H = 360 / 900 = 0.40 or 40.0%.
- Part (c): Carnot efficiency η_th, Carnot = 1 - T_L/T_H. T_H = 800 + 273.15 = 1073.15 K, T_L = 25 + 273.15 = 298.15 K. η_th, Carnot = 1 - (298.15 / 1073.15) = 72.2%.
- Part (d): Feasibility check: Since actual efficiency (40.0%) is less than the reversible Carnot limit (72.2%), this engine cycle is thermodynamically possible and irreversible.

Refrigerator and Heat Pump Schematic Diagram



Coefficient of Performance (C.O.P.) Formulations I

- Refrigerators and Heat Pumps transfer heat from a low-temperature space to a high-temperature space, requiring a net work input.
- Performance is measured by C.O.P. = (Desired Effect) / (Required Work Input). Values of C.O.P. are typically greater than 1.
- For a Refrigerator: The objective is heat extraction from a cold space (Q_L). COP_Ref = Q_L / W_in = Q_L / (Q_H - Q_L).
- For a Heat Pump: The objective is heat delivery to a hot space (Q_H). COP_HP = Q_H / W_in = Q_H / (Q_H - Q_L).
- Inter-relation: COP_HP = Q_H / W_in = (Q_L + W_in) / W_in = COP_Ref + 1. Thus, COP_HP is always greater than COP_Ref by exactly 1.

Numerical Problem 2: Refrigerator COP Analysis I

- Problem: A kitchen refrigerator cools food at a rate of 3 kW while consuming 1.2 kW of electric power. Determine: (a) The actual C.O.P. of the refrigerator, and (b) The rate of heat rejection to the kitchen air.
- Part (a): Cooling load (desired cooling effect) $Q\_L = 3$ kW. Power input (work required) $W\_in = 1.2$ kW.
- $COP\_Ref = Q\_L \,/\, W\_in = 3.0$ kW $/$ 1.2 kW $= 2.50$. This means for every 1.0 kW of work input, 2.5 kW of heat is removed from the cold space.
- Part (b): Applying conservation of energy to the refrigerator cycle: $Q\_H = Q\_L + W\_in$.
- Calculate heat rejection rate: $Q\_H = 3.0$ kW $+ 1.2$ kW $= 4.2$ kW. This is the rate at which heat is discarded to the ambient kitchen air.

Numerical Problem 3: Heat Pump COP Analysis I

- Problem: A heat pump maintains a house at 22 °C by extracting heat from outdoor air at -5 °C. The house loses heat through walls/windows at a rate of 60,000 kJ/h, and the heat pump compressor consumes 5 kW of power. Find: (a) COP_HP, (b) Rate of heat absorption from outdoors, and (c) Carnot COP_HP.
- Part (a): Desired heating rate Q_H = 60,000 kJ/h = 60,000 / 3,600 s = 16.67 kW. Power input W_in = 5.0 kW.
- Calculate actual COP: COP_HP = Q_H / W_in = 16.67 kW / 5.0 kW = 3.33.
- Part (b): Heat absorption rate from outdoors Q_L = Q_H - W_in = 16.67 kW - 5.0 kW = 11.67 kW (or 42,000 kJ/h).
- Part (c): Carnot COP_HP = T_H / (T_H - T_L). T_H = 22 + 273.15 = 295.15 K, T_L = -5 + 273.15 = 268.15 K. COP_HP, Carnot = 295.15 / (295.15 - 268.15) = 295.15 / 27 = 10.93.

Thermodynamic Limit Summary & Formula Matrix I

- Efficiency (η_th) applies to power cycles. C.O.P. applies to refrigeration (COP_Ref) and heat pump (COP_HP) cycles.
- Formulas for actual cycles: η_th = W_net/Q_H = 1 - Q_L/Q_H; COP_Ref = Q_L/W_in; COP_HP = Q_H/W_in.
- Formulas for Carnot (reversible) cycles: η_th, Carnot = 1 - T_L/T_H; COP_Ref,Carnot = T_L/(T_H-T_L); COP_HP,Carnot = T_H/(T_H-T_L).
- COP Relation: COP_HP = COP_Ref + 1 for systems operating between the same two temperature reservoirs.
- Thermodynamic constraints: Actual performance is always lower than Carnot performance (η_th ¡ η_Carnot, COP ¡ COP_Carnot) due to entropy generation.

Concept of Entropy and its T-ds equation. (Without Derivations) I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 3: Second Law of Thermodynamics and Entropy
- Lecture 25: Concept of Entropy, T-ds Equations, and the Third Law of Thermodynamics

Lecture Agenda I

- Thermodynamic Definition of Entropy & Clausius Inequality
- Physical Significance & Molecular Disorder (Boltzmann Relation)
- Graphical Representation on Temperature-Entropy (T-s) Diagrams
- The First T-ds Equation: Origin and Formulation
- The Second T-ds Equation: Origin and Formulation
- Validity and Application of T-ds Equations to Real and Ideal Processes
- Constant Volume vs. Constant Pressure Slopes on T-s Diagrams
- Statement of the Third Law of Thermodynamics
- Absolute Entropy and Behavior at Absolute Zero (0 K)

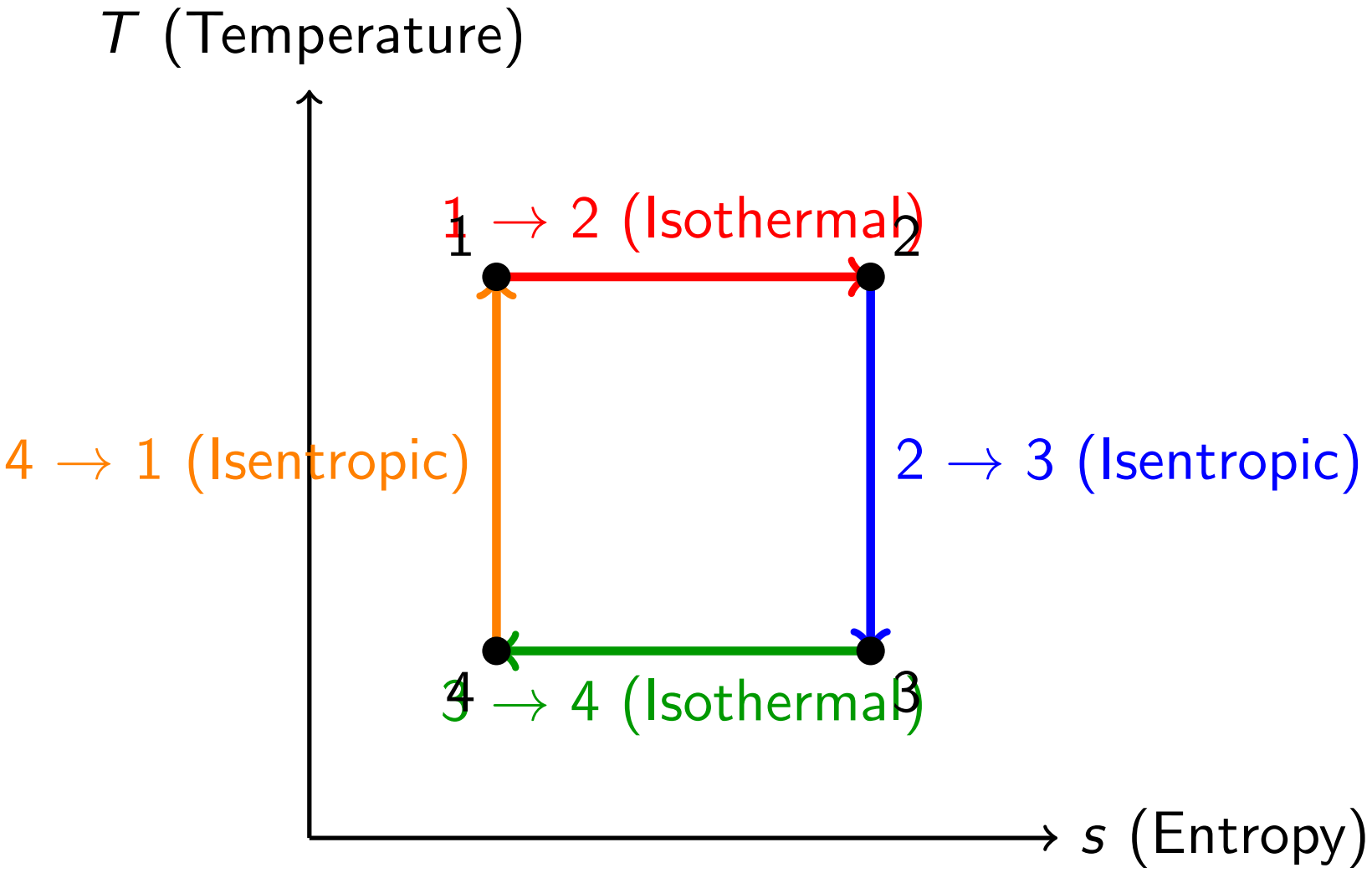
Thermodynamic Definition of Entropy I

- Clausius Inequality establishes that the cyclic integral of dQ/T is less than or equal to zero for all thermodynamic cycles.
- For an internally reversible process, the cyclic integral of dQ/T is exactly zero, proving that dQ_rev/T represents a change in a state property.
- This state property is defined as Entropy (S), where the differential change is $dS = (dQ/T)_{rev}$.
- For real, irreversible processes, the entropy change is $dS = dQ/T + dS_{gen}$, where dS_{gen} represents the entropy generated by irreversibilities ($dS_{gen} > 0$).
- Entropy is an extensive thermodynamic property, measured in units of kJ/K, while specific entropy (s) is intensive with units of kJ/kg-K.

Physical Interpretation of Entropy I

- Molecular Disorder: Boltzmann's relation S = k * ln(W) links thermodynamic entropy to statistical probability, where W is the number of microstates.
- Energy Degradation: Entropy measures the degradation of energy quality; work is fully convertible to heat, but heat cannot be fully converted to work.
- When heat is transferred across a finite temperature difference, the overall entropy of the universe increases, rendering some energy unavailable for work.
- Principle of Increase of Entropy: The entropy of any isolated system must always increase (dS_isolated ¿ 0) or remain constant (dS_isolated = 0) in the limit of a reversible process.

Diagram: Reversible Carnot Cycle on T-s Plane



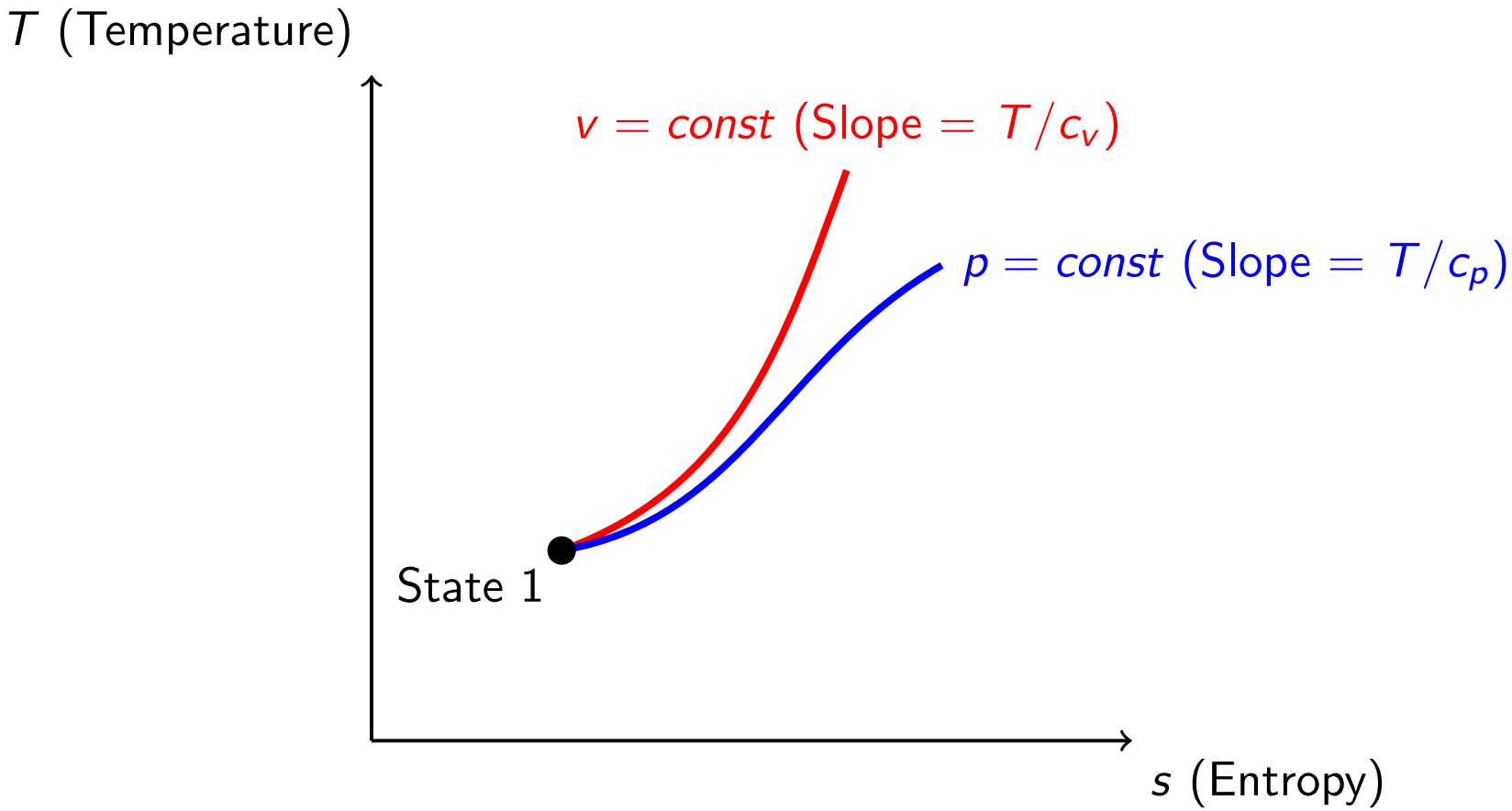
The First T-ds Equation I

- The first Gibbsian T-ds equation is formulated as: T*ds = du + p*dv.
- It is derived by combining the differential First Law of Thermodynamics for a closed system (dq = du + dw) with the Second Law relations.
- For an internally reversible process, the heat transfer is dq = T*ds and the boundary work is dw = p*dv, yielding the fundamental equation.
- Universal Validity: Even though the relation is derived assuming a reversible process, it consists entirely of state properties.
- Consequently, the first T-ds equation is valid for any process (reversible or irreversible) between two equilibrium states in a simple compressible system.

The Second T-ds Equation I

- The second Gibbsian T-ds equation is formulated as: T*ds = dh - v*dp.
- It is obtained by substituting the definition of enthalpy, h = u + p*v, which in differential form is dh = du + p*dv + v*dp.
- Solving for (du + p*dv) yields (dh - v*dp); substituting this back into the first T-ds equation gives the second T-ds relation.
- This relation is particularly useful for analyzing open systems, flow processes, and control volumes where enthalpy is the primary energy quantity.
- Similar to the first relation, it is composed exclusively of properties and holds true for all thermodynamic paths between states.

Diagram: Slopes of Constant Volume and Pressure on T-s Plane



Statement of the Third Law of Thermodynamics I

- Nernst-Simon Statement: The entropy of a pure crystalline substance at absolute zero temperature (0 Kelvin) is exactly zero.
- Mathematical Statement: lim (as T approaches 0) of $S = 0$ for a perfect, defect-free crystal.
- Physical Basis: At absolute zero, all translation, rotation, and vibration of molecules cease, leaving the system in its single lowest-energy quantum state ($W = 1$).
- Significance of the Law: It establishes an absolute physical datum for entropy, allowing scientists to define absolute entropy values instead of just entropy differences.
- Exceptions: Substances that do not form a perfect crystal at absolute zero (e.g. amorphous solids, glasses, or mixtures) exhibit residual entropy (S ¿ 0).

Summary of Entropy & Property Relations I

- Entropy change (dS = dQ/T_rev) is a path-independent state function that measures system disorder.
- The first and second T-ds relations (T*ds = du + p*dv and T*ds = dh - v*dp) relate thermal parameters to mechanical parameters.
- T-ds equations apply universally to both reversible and irreversible paths because they are composed entirely of state properties.
- On T-s coordinates, constant volume lines have a steeper slope than constant pressure lines because specific heat at constant pressure (c_p) is always greater than at constant volume (c_v).
- The Third Law provides the absolute reference zero for entropy calculations by designating S = 0 at T = 0 K for perfect crystals.

Concept of Ideal gas. I

- Course: Engineering Thermodynamics
- Unit 4: Ideal Gases and Thermodynamic Processes
- Instructor: Department of Mechanical Engineering
- Focus: Theoretical formulation, assumptions, and state equations of ideal gases

Agenda I

- Definition of an Ideal Gas and physical significance
- Microscopic kinetic theory assumptions
- The Ideal Gas Equation of State ($Pv = RT$)
- Avogadro's Hypothesis and the Universal Gas Constant
- Visualizing behavior on P-v coordinates
- Deviation from Ideal Gas behavior: Compressibility Factor (Z)
- Graphical representation of Z-factor charts
- Joule's Law: Dependency of internal energy and enthalpy on temperature

What is an Ideal Gas? I

- An ideal gas is a theoretical gas composed of many randomly moving point particles whose only interactions are perfectly elastic collisions.
- It serves as a simplified model to approximate the behavior of real gases at specific thermodynamic states.
- Real gases behave like an ideal gas under conditions of low pressure and high temperature.
- Under these conditions, the intermolecular forces become negligible and the volume occupied by the gas molecules themselves is trivial compared to the container volume.
- It provides a fundamental reference state for evaluating the thermodynamic performance of real systems.

Microscopic Assumptions (Kinetic Theory) I

- 1. Negligible Molecular Volume: The actual volume occupied by the gas molecules themselves is negligible compared to the total volume of the container.
- 2. No Intermolecular Forces: There are no forces of attraction or repulsion between the gas molecules, or between the molecules and the container walls, except during collisions.
- 3. Continuous Random Motion: Molecules are in a constant, rapid, random straight-line motion, colliding frequently with each other and the walls.
- 4. Perfectly Elastic Collisions: All collisions (molecule-molecule and molecule-wall) are perfectly elastic, meaning total kinetic energy is conserved.
- 5. Newtonian Mechanics: The motion of the molecules obeys Newton's laws of motion, and gravity has a negligible effect on their path.

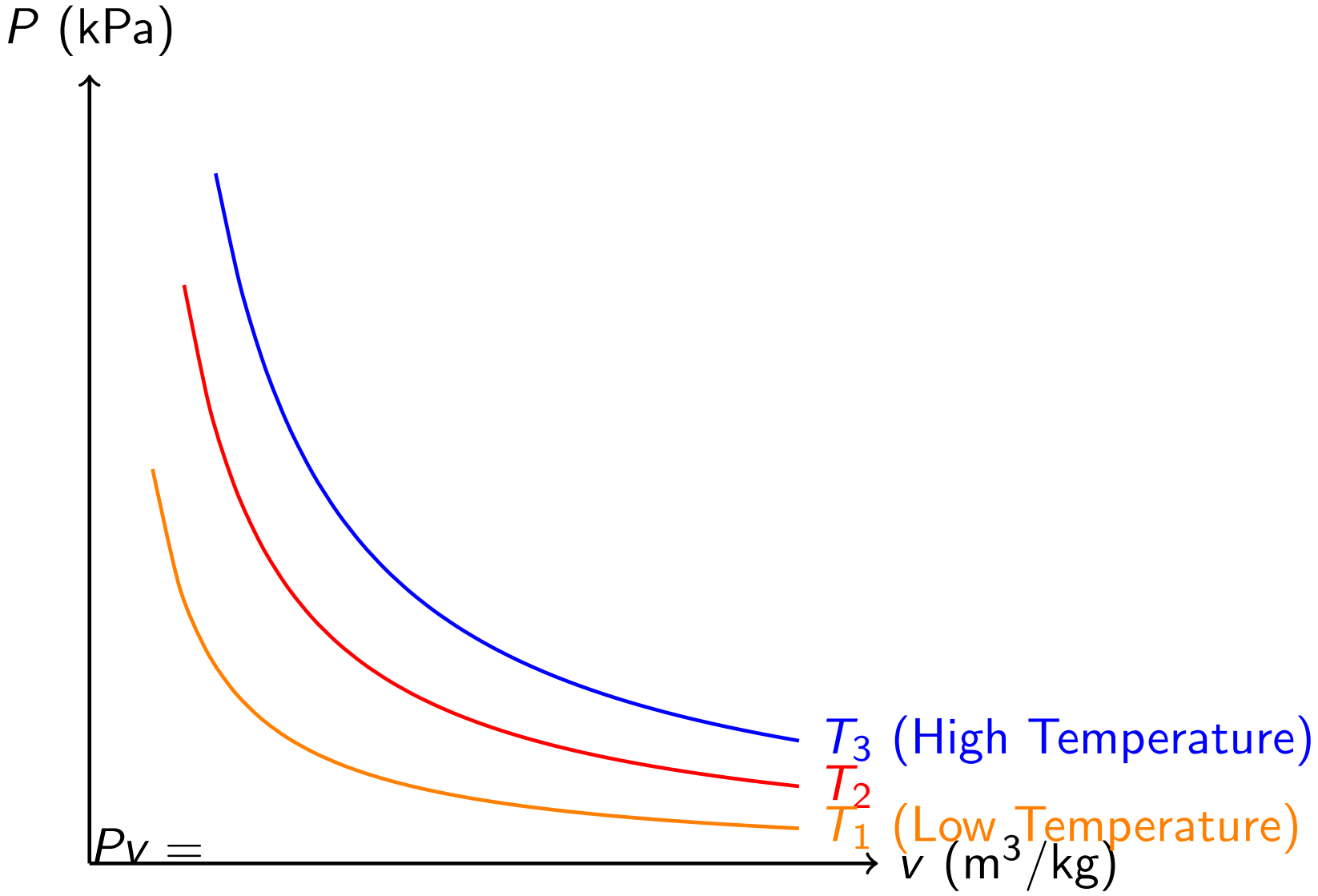
Ideal Gas Equation of State I

- The state of an ideal gas is governed by the equation of state: $PV = mRT$ or $Pv = RT$, where $v = V/m$ is the specific volume.
- Here, P is the absolute pressure (Pa), V is the volume (m^3), m is the mass (kg), T is the absolute temperature (K), and R is the specific gas constant (J/kg*K).
- The specific gas constant R depends on the molecular weight of the gas: $R = Ru / M$, where Ru is the Universal Gas Constant (8.31447 kJ/kmol*K) and M is the molar mass (kg/kmol).
- Alternative molar form: $PV = nRuT$, where n is the number of moles ($n = m/M$).

Ideal Gas Equation of State II

- Historical Gas Laws combined: Boyle's Law (V proportional to 1/P at constant T), Charles's Law (V proportional to T at constant P), and Gay-Lussac's Law (P proportional to T at constant V).

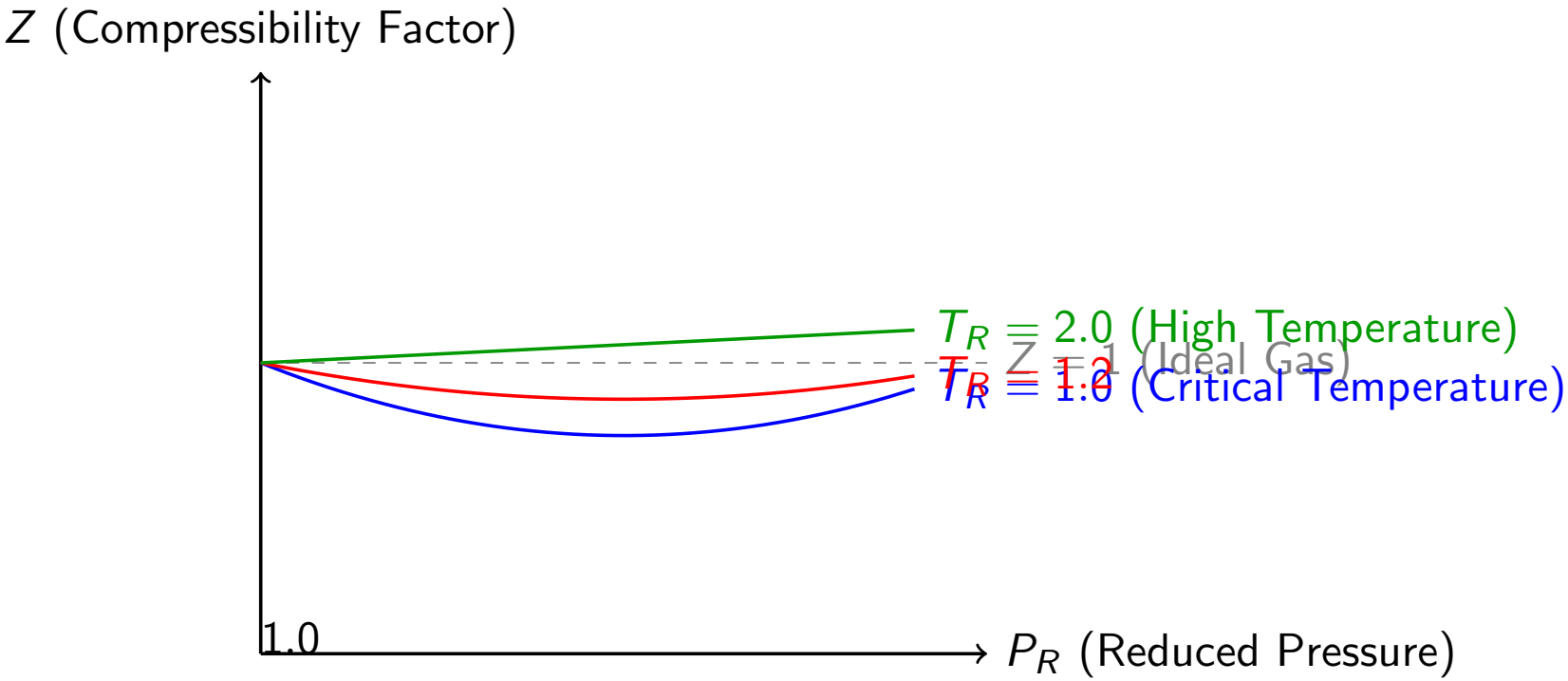
Pressure-Volume (P-v) Diagram of an Ideal Gas



Real Gases and the Compressibility Factor (Z) I

- Real gases deviate from ideal behavior because real molecules have finite volume and exert intermolecular forces on one another.
- To quantify this deviation, we define the Compressibility Factor (Z): $Z = Pv/(RT) = PV/(nRuT)$.
- For an ideal gas, $Z = 1$ under all temperatures and pressures.
- For real gases, Z can be greater than, equal to, or less than 1, depending on the state.
- If Z ¡ 1, intermolecular attraction forces dominate, making the gas more compressible than an ideal gas.
- If Z ¿ 1, molecular volume (repulsive forces) dominates, making the gas less compressible than an ideal gas.
- As pressure approaches zero (P -¿ 0), all real gases behave as ideal gases, and Z -¿ 1.

Compressibility Factor (Z) vs. Reduced Pressure



Thermodynamic Properties of Ideal Gases I

- Joule's Law: The internal energy (u) of an ideal gas depends solely on its temperature, i.e., $u = u(T)$. It is independent of pressure or volume.
- Since $h = u + Pv$ and $Pv = RT$, the enthalpy (h) of an ideal gas also depends only on temperature: $h = u(T) + RT = h(T)$.
- Consequently, the specific heats of an ideal gas are functions of temperature only: $cv(T) = du/dT$ and $cp(T) = dh/dT$.
- Relation between specific heats: $cp(T) - cv(T) = R$, known as Mayer's Relation.
- Specific heat ratio: $k = cp / cv$. For a monoatomic ideal gas, k is approximately 1.67; for diatomic gases, k is approximately 1.40 at room temperature.

Summary and Key Takeaways I

- Ideal gas behavior is a mathematical construct based on negligible molecular volume and zero intermolecular forces.
- The state equation Pv = RT serves as a reasonably accurate model for real gases at low pressure (relative to critical pressure Pc) and high temperature (relative to critical temperature Tc).
- The compressibility factor Z = Pv/RT measures deviation from ideal gas behavior; as P -¿ 0, Z -¿ 1 for all gases.
- Both internal energy (u) and enthalpy (h) are purely functions of temperature (T) for an ideal gas.
- Understanding the ideal gas model is crucial as it forms the baseline for analyzing advanced thermodynamic cycles (Otto, Diesel, Brayton, etc.).

Concept of Ideal gas. I

- Course: Engineering Thermodynamics
- Unit 4: Ideal Gases and Thermodynamic Processes
- Lecture 27: Concept of Ideal Gas
- Instructor: Professor of Mechanical Engineering

Agenda I

- Introduction to the Ideal Gas Model
- Microscopic Assumptions of an Ideal Gas
- The Equation of State: $pV = mRT$ and $pV = nR_uT$
- Difference between Real Gas and Ideal Gas behavior
- Compressibility Factor (Z) and the Ideal Gas Limit
- Visualizing Ideal Gas Relations: p-v and T-v diagrams

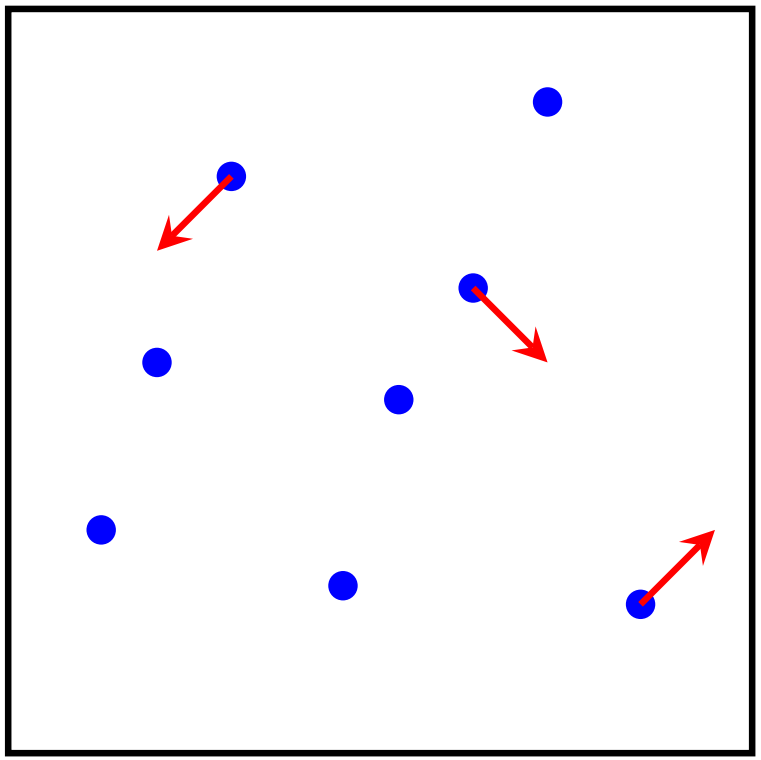
The Ideal Gas Model I

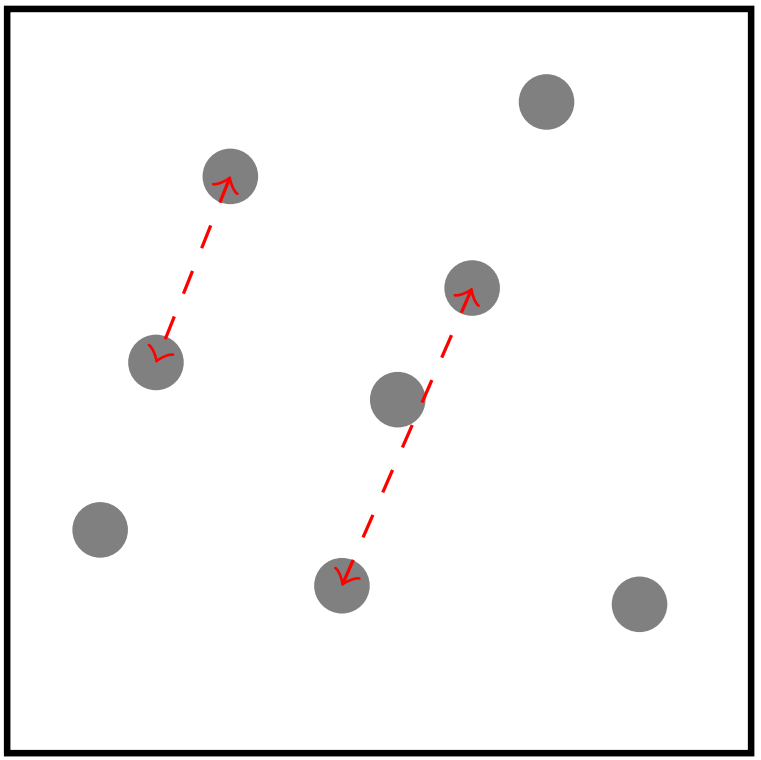
- An ideal gas is a theoretical gas composed of many randomly moving point particles whose only interactions are perfectly elastic collisions.
- It serves as a simplified model to approximate the behavior of real gases under various thermodynamic conditions.
- At low pressures and high temperatures, real gases behave very similarly to an ideal gas.
- Examples of real gases that behave ideally at standard conditions include Nitrogen (N2), Oxygen (O2), Hydrogen (H2), Helium (He), and Air.

Microscopic Assumptions I

- The gas consists of a very large number of molecules in constant, random, straight-line motion.
- The volume of the molecules themselves is negligibly small compared to the total volume occupied by the gas (treated as point masses).
- There are no intermolecular forces of attraction or repulsion between the molecules, except during collisions.
- All collisions between molecules and with the walls of the container are perfectly elastic (kinetic energy is conserved).
- The average kinetic energy of the molecules is directly proportional to the absolute temperature of the gas.

Ideal vs Real Gas Molecules





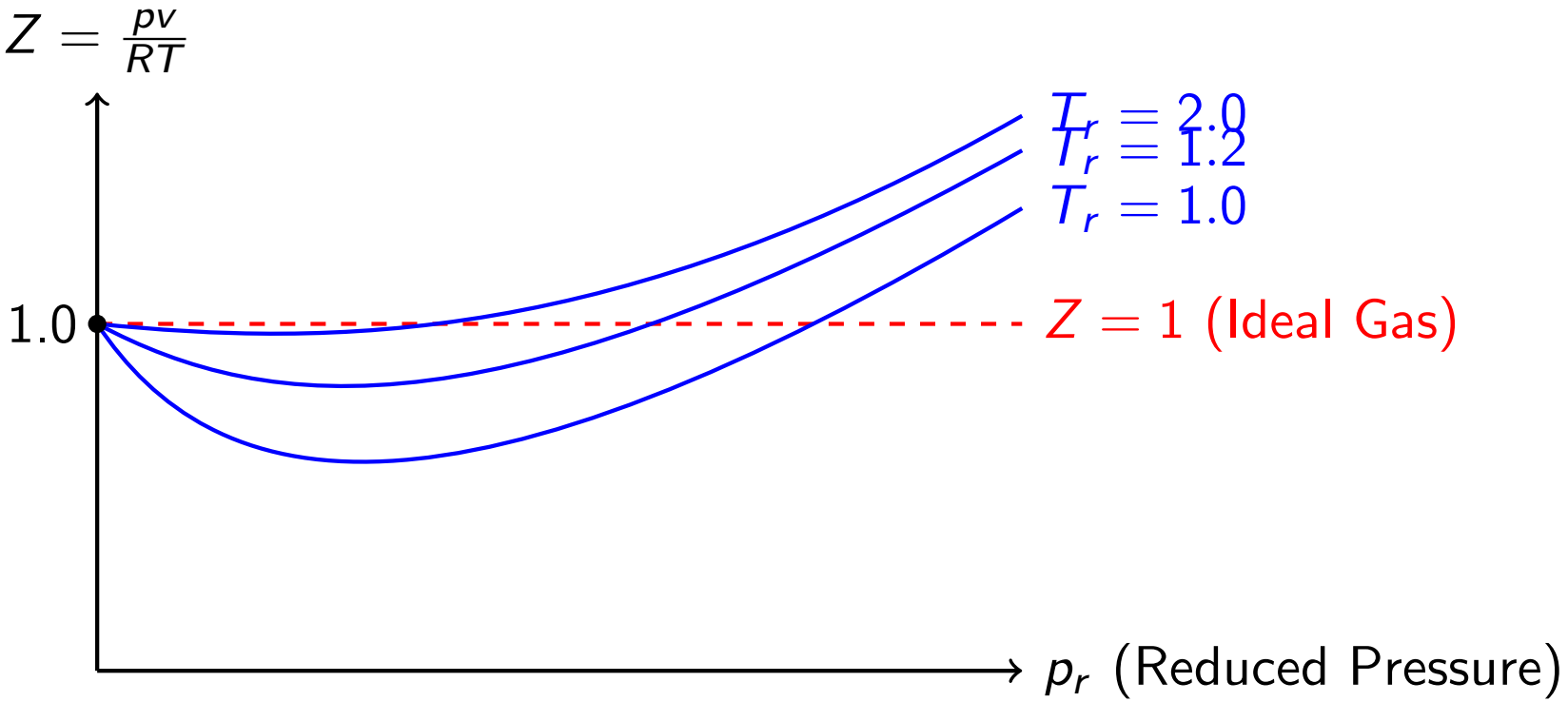
The Ideal Gas Equation of State I

- The relationship between pressure (p), absolute temperature (T), and specific volume (v) or total volume (V) of a gas.
- Equation form: p * V = m * R * T
- Alternatively: p * V = n * Ru * T, where n is the number of moles and Ru is the Universal Gas Constant (Ru = 8.31447 kJ/kmol*K).
- Individual Gas Constant: R = Ru / M, where M is the molar mass of the specific gas. For Air, M = 28.97 kg/kmol, so R = 0.287 kJ/kg*K.
- Specific volume formulation: p * v = R * T, where v = V / m is the specific volume (m^3/kg).

Relational Gas Laws I

- Boyle's Law: At constant temperature (T = const), pressure is inversely proportional to volume: p1 * V1 = p2 * V2.
- Charles's Law: At constant pressure (p = const), volume is directly proportional to absolute temperature: V1 / T1 = V2 / T2.
- Gay-Lussac's Law: At constant volume (V = const), pressure is directly proportional to absolute temperature: p1 / T1 = p2 / T2.
- Avogadro's Law: Equal volumes of all ideal gases at the same temperature and pressure contain the same number of molecules.

Compressibility Factor (Z) Diagram



Ideal Gas vs. Real Gas I

- Real gases deviate from ideal gas behavior because molecules have finite volume and exert intermolecular forces on each other.
- At high pressure, molecules are packed closely together, making their molecular volume non-negligible.
- At low temperature, molecules move slowly, allowing intermolecular attractive forces to pull them together, reducing the pressure exerted on the container walls.
- Ideal gas assumption is highly accurate when: Pressure (p) is much lower than the critical pressure (p_cr), and/or Temperature (T) is much higher than the critical temperature (T_cr).

Summary I

- An ideal gas is a thermodynamic model that assumes point-like molecules with no intermolecular forces and elastic collisions.
- The equation of state, $pV = mRT$, relates pressure, volume, and temperature, and is widely used for engineering calculations.
- The compressibility factor $Z = pv / (RT)$ quantifies deviation from ideal behavior; $Z = 1$ for an ideal gas.
- Air, Nitrogen, Oxygen, and other light gases can be treated as ideal gases at standard conditions for engineering design.

Boyle's law, Charle's law and Gay-Lussac law for ideal gases. I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 4: Ideal Gases and Thermodynamic Processes
- Lecture 28: Boyle's law, Charle's law and Gay-Lussac law
- Department of Mechanical Engineering

Agenda I

- The Ideal Gas Model and Kinetic Theory Assumptions
- Boyle's Law: Isothermal Volume-Pressure Relationship
- Charles's Law: Isobaric Temperature-Volume Relationship
- Gay-Lussac's Law: Isochoric Temperature-Pressure Relationship
- The Combined Gas Law and the Ideal Gas Equation of State
- Engineering Applications, Limits of Validity, and Compressibility Factor

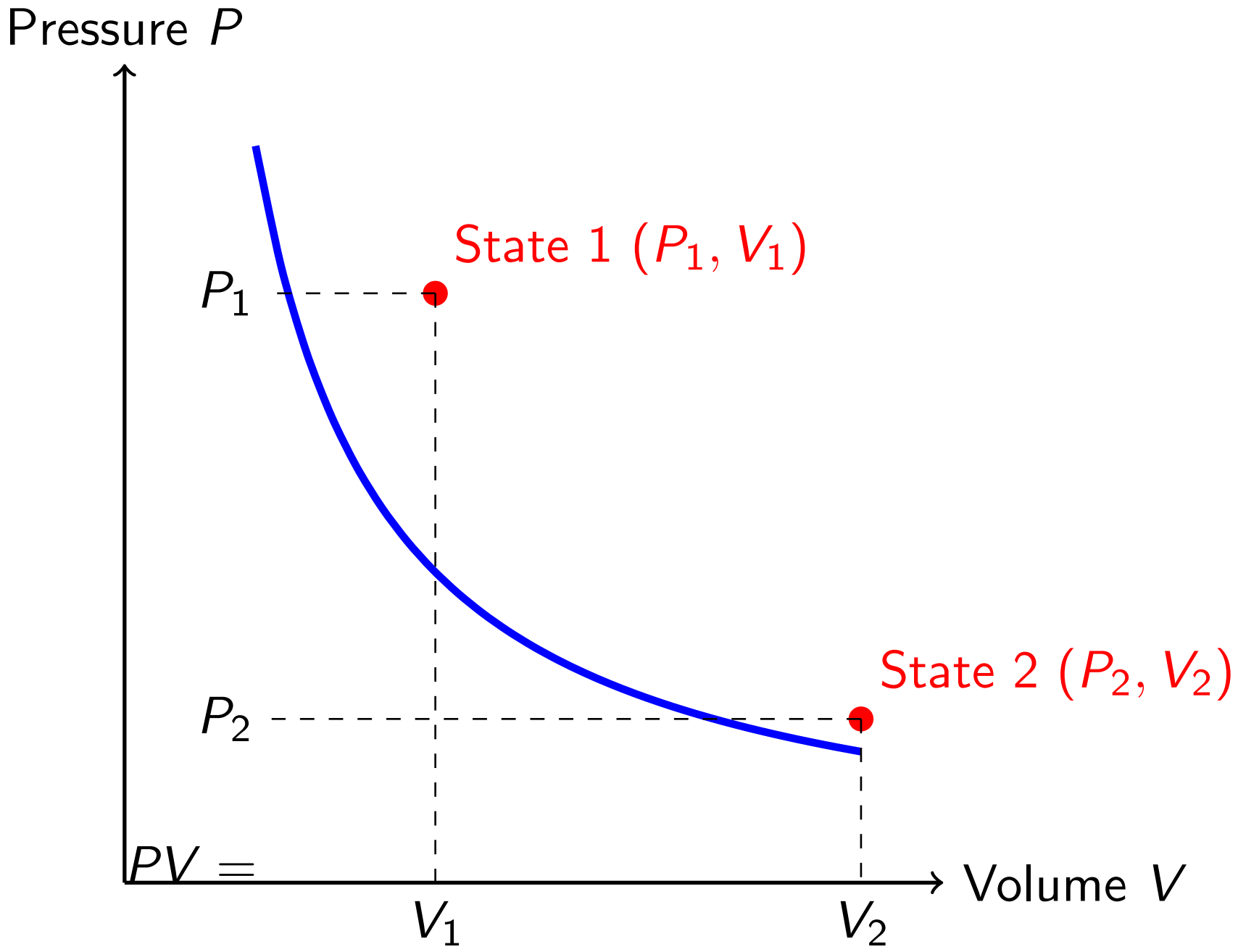
The Ideal Gas Model and Key Assumptions I

- An ideal gas is a simplified model representing real gases under low-pressure and high-temperature conditions.
- Assumption 1 (Volume): The volume occupied by the gas molecules themselves is negligible compared to the total volume of the container.
- Assumption 2 (Forces): Intermolecular forces (attractive or repulsive) between gas molecules are assumed to be zero.
- Assumption 3 (Motion): Molecules are in continuous, random motion, colliding with each other and the container walls.
- Assumption 4 (Collisions): All molecular collisions are perfectly elastic, meaning kinetic energy is conserved during impacts.
- Engineering Significance: Serves as a baseline reference for analyzing real gases and simplifies thermodynamic equations of state.

Boyle's Law: Isothermal Volume-Pressure Relation I

- Formulated by Robert Boyle in 1662, describing isothermal processes where temperature is held constant ($T = \text{const}$).
- Statement: For a fixed mass of an ideal gas at constant temperature, absolute pressure is inversely proportional to volume.
- Mathematical formulation: $P \propto 1/V \implies PV = C$ (constant), which implies $P_1 V_1 = P_2 V_2$.
- Thermodynamic Work: During an isothermal expansion/compression, boundary work is $W_{1-2} = \int P dV = P_1 V_1 \ln(V_2/V_1)$.
- Practical engineering applications: Operation of low-speed air compressors, pneumatic actuators, and pressure vessel sizing.

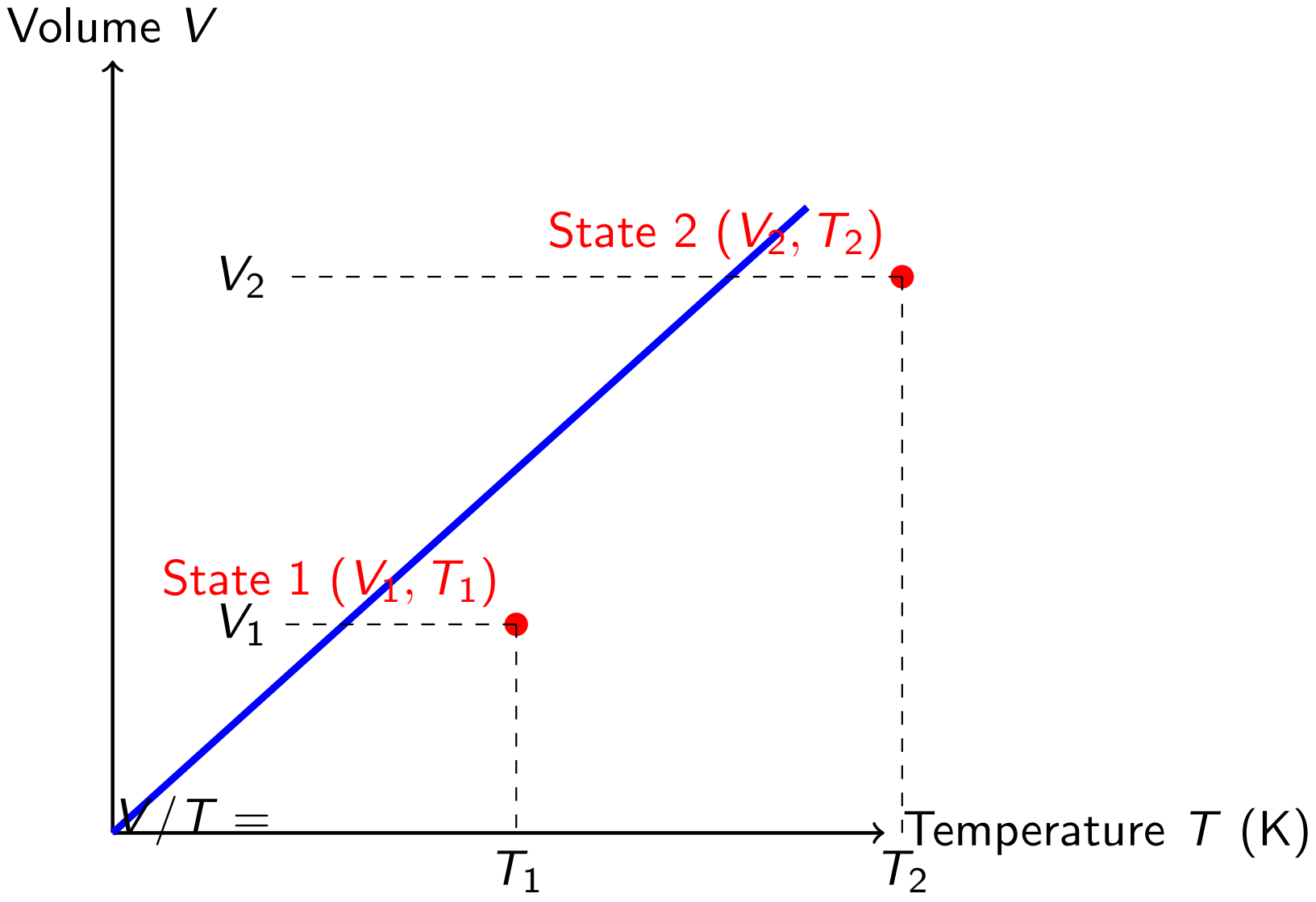
Boyle's Law: Isothermal Expansion Diagram



Charles's Law: Isobaric Temperature-Volume Relation I

- Formulated by Jacques Charles in 1787, describing isobaric processes where pressure is held constant ($P = \text{const}$).
- Statement: For a fixed mass of an ideal gas at constant pressure, volume is directly proportional to absolute temperature.
- Mathematical formulation: $V \propto T \implies V/T = C$ (constant), which implies $V_1/T_1 = V_2/T_2$.
- Absolute Temperature: The temperature T must be expressed in Kelvin ($T(K) = T(^\circ\text{C}) + 273.15$).
- Thermodynamic Work: Work done during an isobaric process is simply evaluated as $W_{1-2} = P(V_2 - V_1)$.
- Practical engineering applications: Hot air balloons, isobaric heat exchangers (condensers/evaporators), and gas thermometer calibration.

Charles's Law: Isobaric Heating Diagram



Gay-Lussac's Law: Isochoric Temperature-Pressure Relation I

- Formulated by Joseph Louis Gay-Lussac in 1802, describing isochoric processes where volume is held constant ($V = \text{const}$).
- Statement: For a fixed mass of an ideal gas at constant volume, absolute pressure is directly proportional to absolute temperature.
- Mathematical formulation: $P \propto T \implies P/T = C$ (constant), which implies $P_1/T_1 = P_2/T_2$.
- Boundary Work: Since volume is constant, the boundary work is zero ($W_{1-2} = \int PdV = 0$).
- Heat Transfer: Heat added goes entirely into raising internal energy: $Q_{1-2} = \Delta U = mc_v(T_2 - T_1)$.
- Practical engineering applications: Pressure safety valves on boilers, gas cylinders exposed to heat, and constant-volume combustion chambers.

The Combined Gas Law and the Ideal Gas Equation I

- Combined Gas Law: For a fixed mass of gas, combining the three laws yields $P_1 V_1 / T_1 = P_2 V_2 / T_2$.
- Equation of State: Generalizing for any quantity, we get $PV = nR_u T$, where $R_u = 8.3144\ \text{kJ/kmol} \cdot \text{K}$ is the universal gas constant.
- Mass-Based Form: $PV = mRT$, where m is mass (kg) and $R = R_u / M$ is the characteristic gas constant for a specific gas.
- Example values of R: For air, $R_{\text{air}} = 0.287\ \text{kJ/kg} \cdot \text{K}$; for Helium, $R_{\text{He}} = 2.0769\ \text{kJ/kg} \cdot \text{K}$.
- Limits of Validity: The ideal gas equation of state is highly accurate when pressure is low relative to critical pressure ($P \ll P_{cr}$) and temperature is high relative to critical temperature ($T \gg T_{cr}$).

The Combined Gas Law and the Ideal Gas Equation II

- Compressibility Factor: Deviation from ideal behavior is measured by $Z = PV/(RT)$, where $Z = 1$ indicates perfect ideal gas behavior.

Summary of Gas Laws & Key Equations I

- Equation of State: $PV = mRT$ and $PV = nR_uT$
- Boyle's Law (Isothermal, $T = \text{const}$): $P_1V_1 = P_2V_2$
- Charles's Law (Isobaric, $P = \text{const}$): $V_1/T_1 = V_2/T_2$
- Gay-Lussac's Law (Isochoric, $V = \text{const}$): $P_1/T_1 = P_2/T_2$
- Critical Reminder: Pressure P and Temperature T must always be in absolute terms (T in Kelvin, P in Pascal absolute).

Boyle's law, Charle's law and Gay-Lussac law for ideal gases. I

- Course: Engineering Thermodynamics
- Unit 4: Ideal Gases and Thermodynamic Processes
- Lecture 29: Boyle's law, Charle's law and Gay-Lussac law for ideal gases.
- Instructor: Mechanical Engineering Professor

Agenda I

- Introduction to Ideal Gas Behaviour
- Boyle's Law (Constant Temperature Process)
- Charles's Law (Constant Pressure Process)
- Gay-Lussac's Law (Constant Volume Process)
- Combined Ideal Gas Relation
- Thermodynamic Diagrams (P-V, T-V, P-T)
- Engineering Applications and Real Gas Deviations

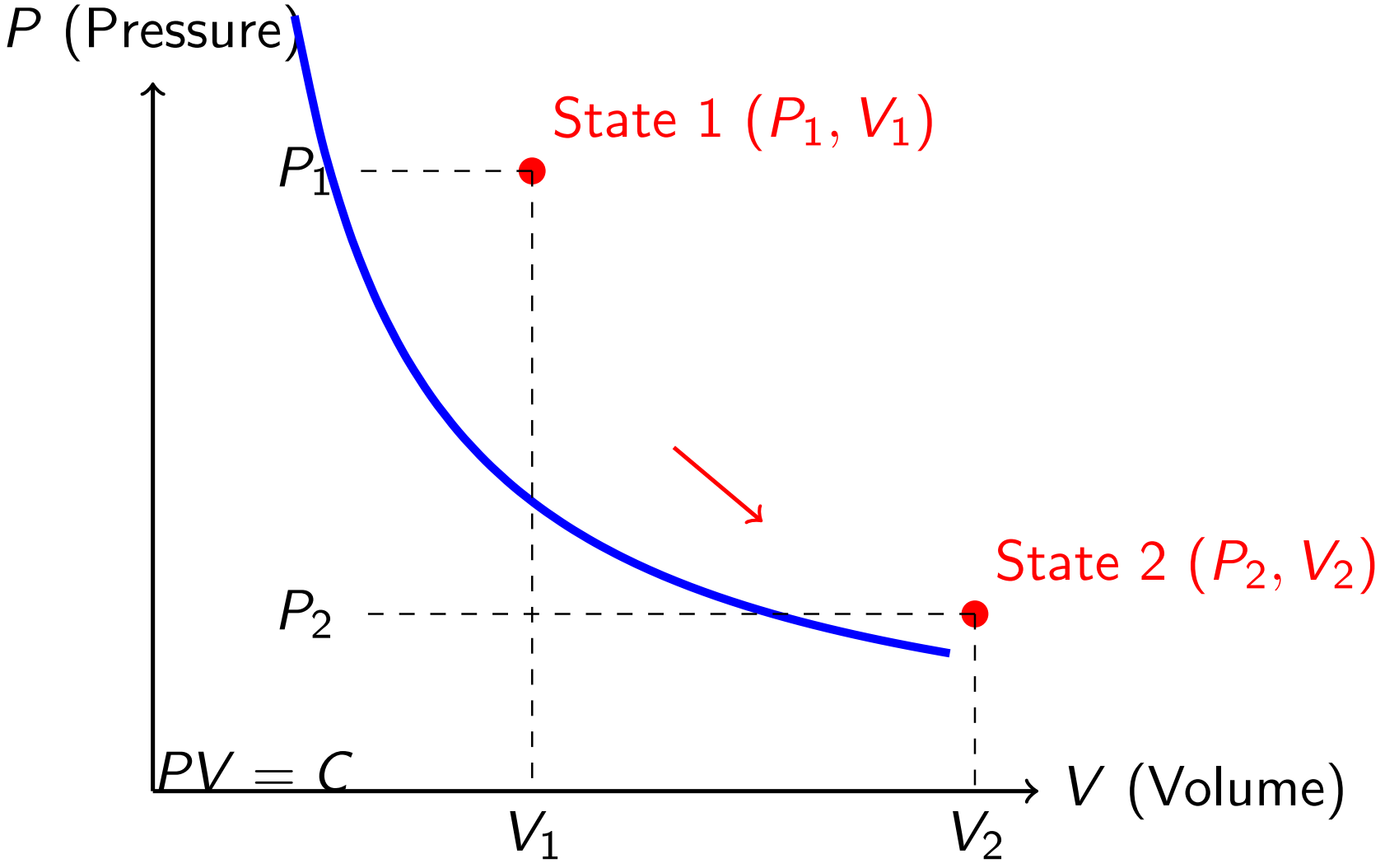
Introduction to the Ideal Gas State I

- An ideal gas is a hypothetical gas whose molecules exhibit negligible intermolecular forces and undergo perfectly elastic collisions.
- Equation of State: P * v = R * T, where P is absolute pressure (Pa), v is specific volume (m^3/kg), T is absolute temperature (K), and R is the specific gas constant (J/kg.K).
- At low pressures and high temperatures relative to their critical points, real gases behave closely to ideal gas predictions.
- The state of a simple compressible system is completely specified by two independent thermodynamic properties.

Boyle's Law (Isothermal Behaviour) I

- Formulated by Robert Boyle in 1662, this law governs constant-temperature (isothermal) processes: T = constant.
- Statement: For a fixed mass of an ideal gas, absolute pressure is inversely proportional to its volume: P * V = constant (or P1 * V1 = P2 * V2).
- Thermodynamic implication: The path of an isothermal process on a P-V diagram is a hyperbola.
- Mechanical application: Compression strokes in reciprocating compressors operating at slow speeds with high heat rejection rates.

Boyle's Law P-V Diagram



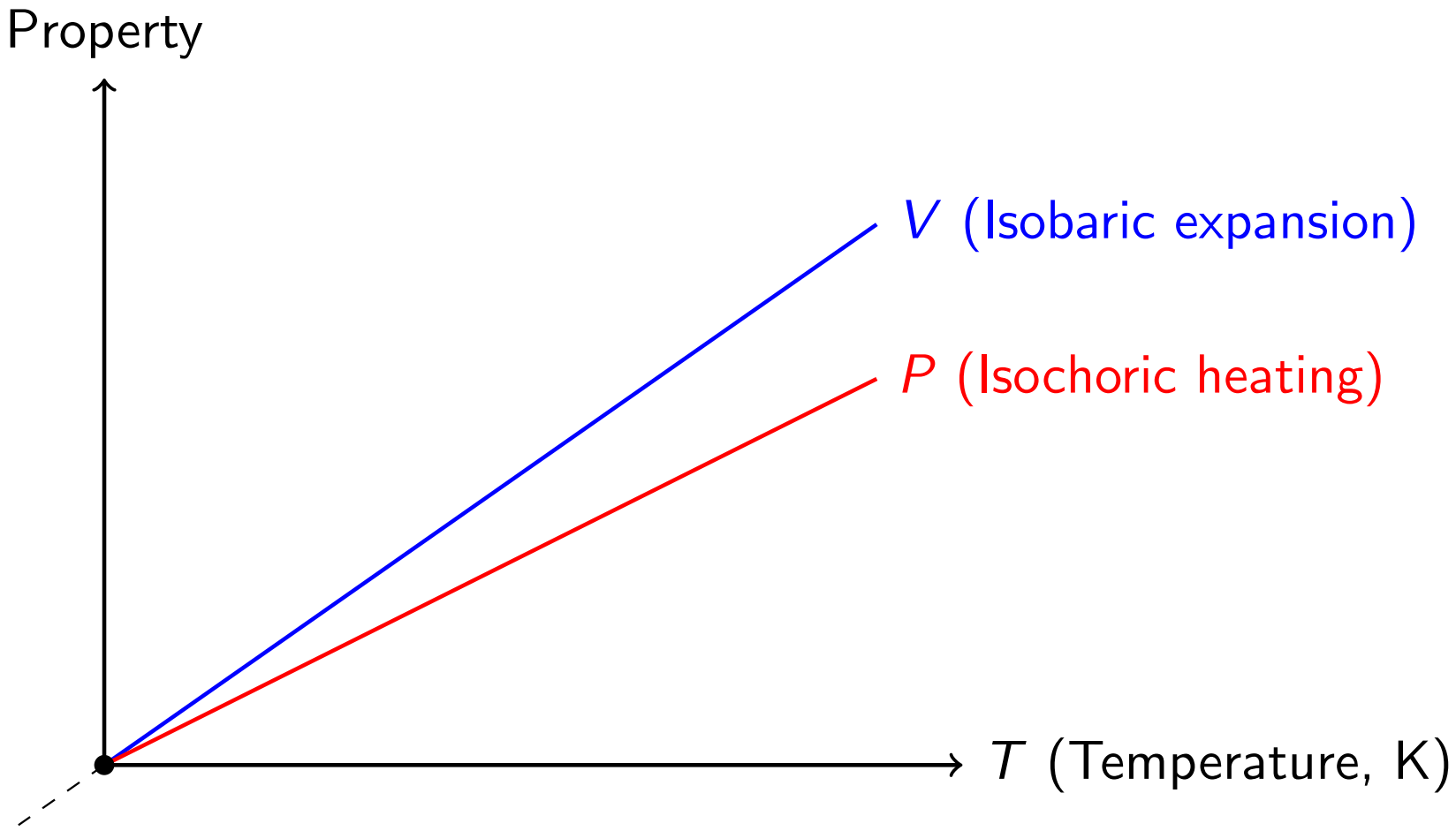
Charles's Law (Isobaric Behaviour) I

- Formulated by Jacques Charles in 1787, this law governs constant-pressure (isobaric) processes: P = constant.
- Statement: The volume of a fixed mass of an ideal gas is directly proportional to its absolute temperature: V / T = constant (or V1 / T1 = V2 / T2).
- Important: Temperature MUST be in Kelvin (absolute scale). The zero-volume intercept occurs at absolute zero (0 K or -273.15 °C).
- Thermodynamic implication: Work done during an isobaric expansion is simply W = P * (V2 - V1), where expansion is driven by heat transfer.

Gay-Lussac's Law (Isochoric Behaviour) I

- Formulated by Joseph Louis Gay-Lussac in 1802, this law governs constant-volume (isochoric or isometric) processes: V = constant.
- Statement: The absolute pressure of a fixed mass of an ideal gas is directly proportional to its absolute temperature: P / T = constant (or P1 / T1 = P2 / T2).
- Thermodynamic implication: Since there is no boundary work in a rigid container (W = 0), the heat transfer equals the change in internal energy: Q = dU = m * c_v * (T2 - T1).
- Practical engineering applications: Pressure safety valves on rigid storage tanks (boilers, gas cylinders) to prevent overpressurization during heating.

Isobaric vs Isochoric Processes



The Combined Ideal Gas Law I

- By combining Boyle's, Charles's, and Gay-Lussac's laws, we get the relation: (P1 * V1) / T1 = (P2 * V2) / T2 = constant.
- This allows the calculation of an unknown state property when a fixed mass of gas undergoes any change of state.
- Universal Gas Constant: R_u = 8.314 kJ/kmol.K. The specific gas constant R = R_u / M, where M is the molecular weight of the gas.
- Limitations of the Ideal Gas Laws: Deviations become significant at high pressures (P ¿ 10 atm) and low temperatures (near condensation).

Summary of Gas Laws & Engineering Equations I

- Boyle's Law: P1 * V1 = P2 * V2 at constant T. Hyperbolic path on P-V diagram.
- Charles's Law: V1 / T1 = V2 / T2 at constant P. Linear path on V-T diagram.
- Gay-Lussac's Law: P1 / T1 = P2 / T2 at constant V. Linear path on P-T diagram.
- Combined Law: P1 * V1 / T1 = P2 * V2 / T2. Ideal gas equation of state: P * V = m * R * T.
- Engineering applications include refrigeration, combustion engines, air compressors, and pneumatic machinery.

Characteristic gas equation and Universal gas constant, Specific heats of gas and their relationship. I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 4: Ideal Gases and Thermodynamic Processes
- Lecture 30: Characteristic gas equation and Universal gas constant, Specific heats of gas and their relationship.
- Instructor: Professor of Mechanical Engineering

Agenda I

- Introduction to Ideal Gas Laws and Equation of State
- Characteristic Gas Equation and the Characteristic Gas Constant (R)
- Universal Gas Constant (Ru) and its Relationship with R
- Specific Heats of Gases: Constant Volume (Cv) and Constant Pressure (Cp)
- Thermodynamic Derivation of Mayer's Relation (Cp - Cv = R)
- Ratio of Specific Heats (Gamma) and its Physical Significance

The Ideal Gas Equation of State I

- An ideal gas is a hypothetical gas whose molecules exhibit no intermolecular attractive forces and undergo perfectly elastic collisions.
- At low pressures and high temperatures, real gases behave closely to ideal gases.
- Boyle's Law states that at constant temperature, pressure is inversely proportional to volume: $\mathrm{p} \propto 1/V$.
- Charles's Law states that at constant pressure, volume is directly proportional to absolute temperature: $\mathrm{V} \propto T$.
- Combining these laws yields the equation of state: $\mathrm{pV} = \mathrm{mRT}$, where m is the mass of the gas and R is the characteristic gas constant.

Characteristic vs. Universal Gas Constants I

- The characteristic gas constant, R, is specific to each individual gas. It is calculated as R = Ru / M, where M is the molecular weight of the gas.
- The SI units of the characteristic gas constant R are J/(kg K) or kJ/(kg K). For example, for air, R $\approx 287 J/(kgK) or 0.287 kJ/(kgK)$.

- The universal gas constant, Ru, is a fundamental physical constant applicable to all ideal gases.
- Ru has a value of approximately 8.3144 J/(mol K) or 8.3144 kJ/(kmol K).
- Avogadro's law states that equal volumes of all gases at the same temperature and pressure contain the same number of molecules, which implies pV = n Ru T, where n is the number of moles (n = m/M).

$pV = mRT$; [dashed] (1.5,0) node[below] V_A – (1.5, 2.33) – (0, 2.33) node[left] p_A; [black] (1.5, 2.33) circle (2pt) node[above right] A;

Specific Heats of Gases: Cp and Cv I

- Specific heat is defined as the amount of heat energy required to raise the temperature of a unit mass of a substance by one degree.
- Unlike solids and liquids which have nearly constant volume and pressure, gases expand significantly when heated, requiring two distinct specific heats.
- Specific heat at constant volume (Cv) is the heat required to raise the temperature of a unit mass of gas by one degree while maintaining a constant volume: $Cv = (du/dT)_v$.
- Specific heat at constant pressure (Cp) is the heat required to raise the temperature of a unit mass of gas by one degree while maintaining a constant pressure: $Cp = (dh/dT)_p$.
- Because a gas heated at constant pressure expands and performs boundary work, Cp is always greater than Cv (Cp ¿ Cv).

$Q_p = \Delta U + W$
$Q_v = \Delta U$
Hence, $C_p > C_v$;

Derivation of Mayer's Relation (Cp - Cv = R) I

- Enthalpy (h) is defined as: $h = u + pv$.
- For an ideal gas, internal energy (u) depends only on temperature: $u = u(T)$ and $du = Cv\ dT$.
- From the characteristic gas equation, the flow work term pv can be replaced by RT: $pv = RT$.
- Substituting this into the enthalpy equation yields: $h = u + RT$.
- Differentiating both sides with respect to temperature T gives: $dh/dT = du/dT + d(RT)/dT$.
- Since $dh = Cp\ dT$ and $du = Cv\ dT$, we substitute these into the differentiated equation: $Cp = Cv + R$.
- This yields Mayer's Relation: $Cp - Cv = R$, demonstrating that the difference between specific heats is the characteristic gas constant.

Ratio of Specific Heats (Gamma) I

- The ratio of specific heats is defined as: γ = Cp / Cv. It is also known as the adiabatic index or isentropic expansion factor.
- Since Cp is always greater than Cv, the value of γ is always greater than 1.
- For a monatomic gas (e.g., Helium, Argon), molecular degrees of freedom f = 3 (translation only). This leads to Cv = (3/2)R, Cp = (5/2)R, and $\gamma \approx 1.67$.

- For a diatomic gas (e.g., Nitrogen, Oxygen, Air) at room temperature, degrees of freedom f = 5 (3 translation + 2 rotation). This leads to Cv = (5/2)R, Cp = (7/2)R, and $\gamma \approx 1.40$.

Ratio of Specific Heats (Gamma) II

- For polyatomic gases (e.g., CO2, H2O), vibrational modes contribute, increasing f, which reduces γ closer to 1.30 or lower.
- The value of γ is critical in calculating temperature and pressure changes during adiabatic/isentropic processes.

Summary and Key Engineering Equations I

- The characteristic gas equation: pV = mRT (mass basis) or pV = nRuT (mole basis).
- Gas constant relations: R = Ru / M, and the difference relation: Cp - Cv = R.
- Using the ratio of specific heats γ = Cp / Cv, we can express Cp and Cv in terms of R and γ: Cv = R / (γ - 1) and Cp = γ R / (γ - 1).
- These relationships are fundamental for analyzing thermodynamic cycles, including the Otto, Diesel, and Brayton cycles.
- In gas dynamics and compressible flow, γ determines the speed of sound: a = $\sqrt{\gamma}$ R T.

Thermodynamic Processes, its representation on P-V (Pressure- Volume) and T-S (Temperature-Entropy) diagram: I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 4: Ideal Gases and Thermodynamic Processes
- Lecture 31: Thermodynamic Processes and Graphical Representations
- Presenter: Mechanical Engineering Department

Agenda I

- Introduction to Thermodynamic Processes & the Polytropic Index (n)
- Analysis of Isobaric (P=C) and Isochoric (V=C) Processes
- Analysis of Isothermal (T=C) and Isentropic (S=C) Processes
- Comparison on the Pressure-Volume (P-V) Plane
- Comparison on the Temperature-Entropy (T-S) Plane
- Summary of Work, Heat, and Internal Energy Formulations

The Polytropic Process & Index Classification I

- Many thermodynamic processes can be modeled by the polytropic relation: P * V^n = Constant, where 'n' is the polytropic index.
- The index 'n' dictates the type of process and determines the path on thermodynamic diagrams.
- Isobaric Process: n = 0 (Pressure is constant, P = C)
- Isothermal Process: n = 1 (Temperature is constant for an ideal gas, PV = C)
- Polytropic Process: 1 ¡ n ¡ gamma (General expansion/compression with heat transfer)
- Isentropic/Adiabatic Process: n = gamma (Reversible & adiabatic, no heat transfer, S = C)
- Isochoric Process: n = infinity (Volume is constant, V = C)

Constant Pressure (Isobaric) Process (n = 0) I

- Governing Equation: P = Constant, which implies V1/T1 = V2/T2 (Charles's Law for ideal gases).
- Boundary Work: W = P * (V2 - V1) = m * R * (T2 - T1). This is represented as the area under the horizontal line on a P-V diagram.
- Heat Transfer: Q = m * Cp * (T2 - T1) = H2 - H1, where Cp is the specific heat capacity at constant pressure and H is enthalpy.
- Change in Internal Energy: dU = m * Cv * (T2 - T1).
- Entropy Change: dS = m * Cp * ln(T2/T1) - m * R * ln(P2/P1) = m * Cp * ln(T2/T1) (since P2 = P1).
- Representation: A horizontal line on P-V diagram, and a logarithmic curve with slope T/Cp on T-S diagram.

Constant Volume (Isochoric) Process (n = infinity) I

- Governing Equation: V = Constant, which implies P1/T1 = P2/T2 (Gay-Lussac's Law for ideal gases).
- Boundary Work: W = integral(P dV) = 0 (since dV = 0). No boundary work is performed during a constant volume process.
- Heat Transfer: Q = dU = m * Cv * (T2 - T1), where Cv is the specific heat capacity at constant volume.
- Change in Enthalpy: dH = m * Cp * (T2 - T1).
- Entropy Change: dS = m * Cv * ln(T2/T1) + m * R * ln(V2/V1) = m * Cv * ln(T2/T1) (since V2 = V1).
- Representation: A vertical line on P-V diagram, and a logarithmic curve with slope T/Cv on T-S diagram (steeper than isobaric).

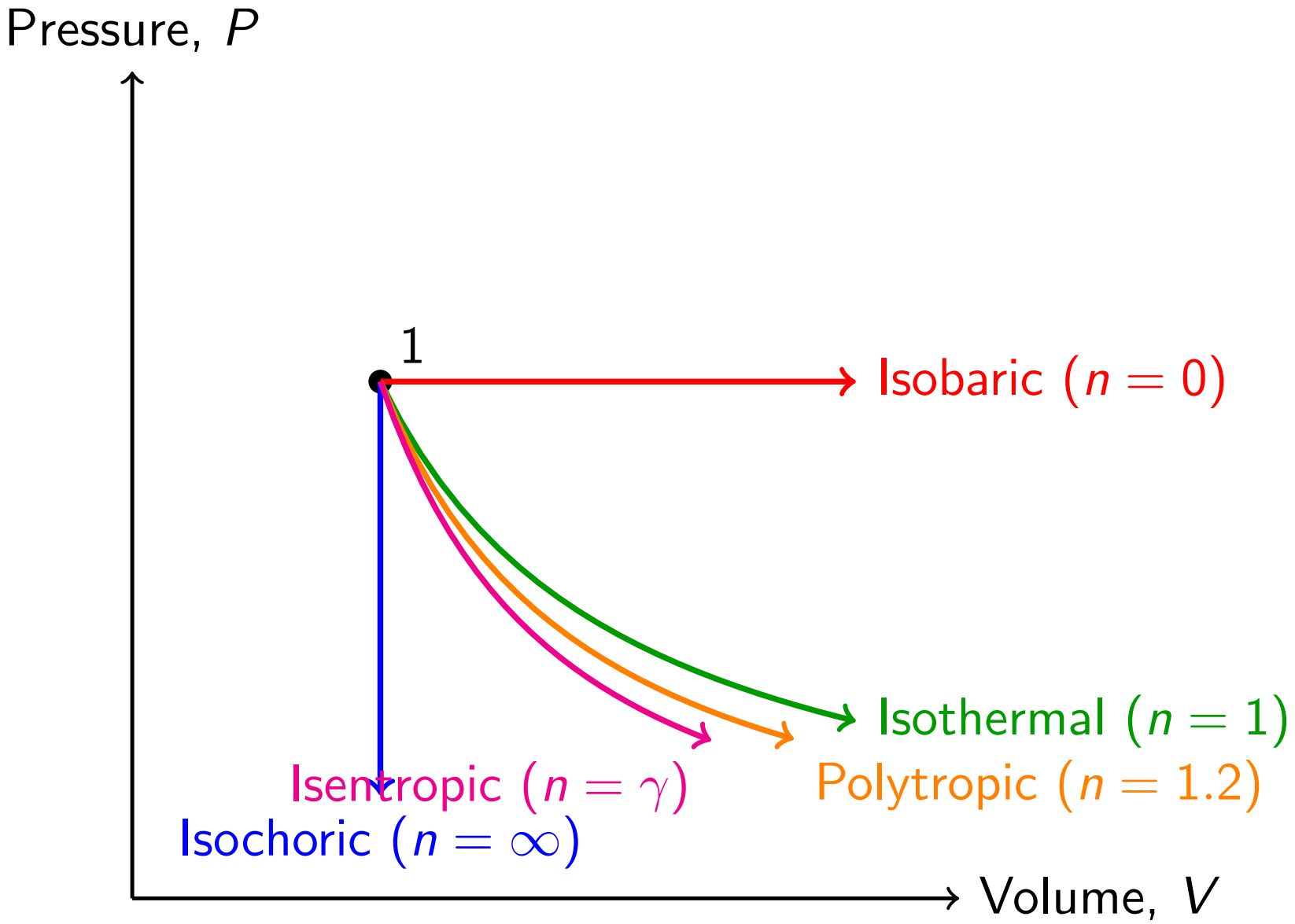
Constant Temperature (Isothermal) Process (n = 1) I

- Governing Equation: T = Constant, which implies P1 * V1 = P2 * V2 (Boyle's Law for ideal gases).
- Change in Internal Energy & Enthalpy: dU = 0 and dH = 0 (since internal energy and enthalpy of an ideal gas depend solely on temperature).
- Boundary Work: W = P1 * V1 * ln(V2/V1) = m * R * T * ln(V2/V1) = P1 * V1 * ln(P1/P2).
- Heat Transfer: Q = W = m * R * T * ln(V2/V1) (from the First Law of Thermodynamics, since dU = 0).
- Entropy Change: dS = Q/T = m * R * ln(V2/V1) = -m * R * ln(P2/P1).
- Representation: A hyperbola on the P-V diagram, and a horizontal line on the T-S diagram.

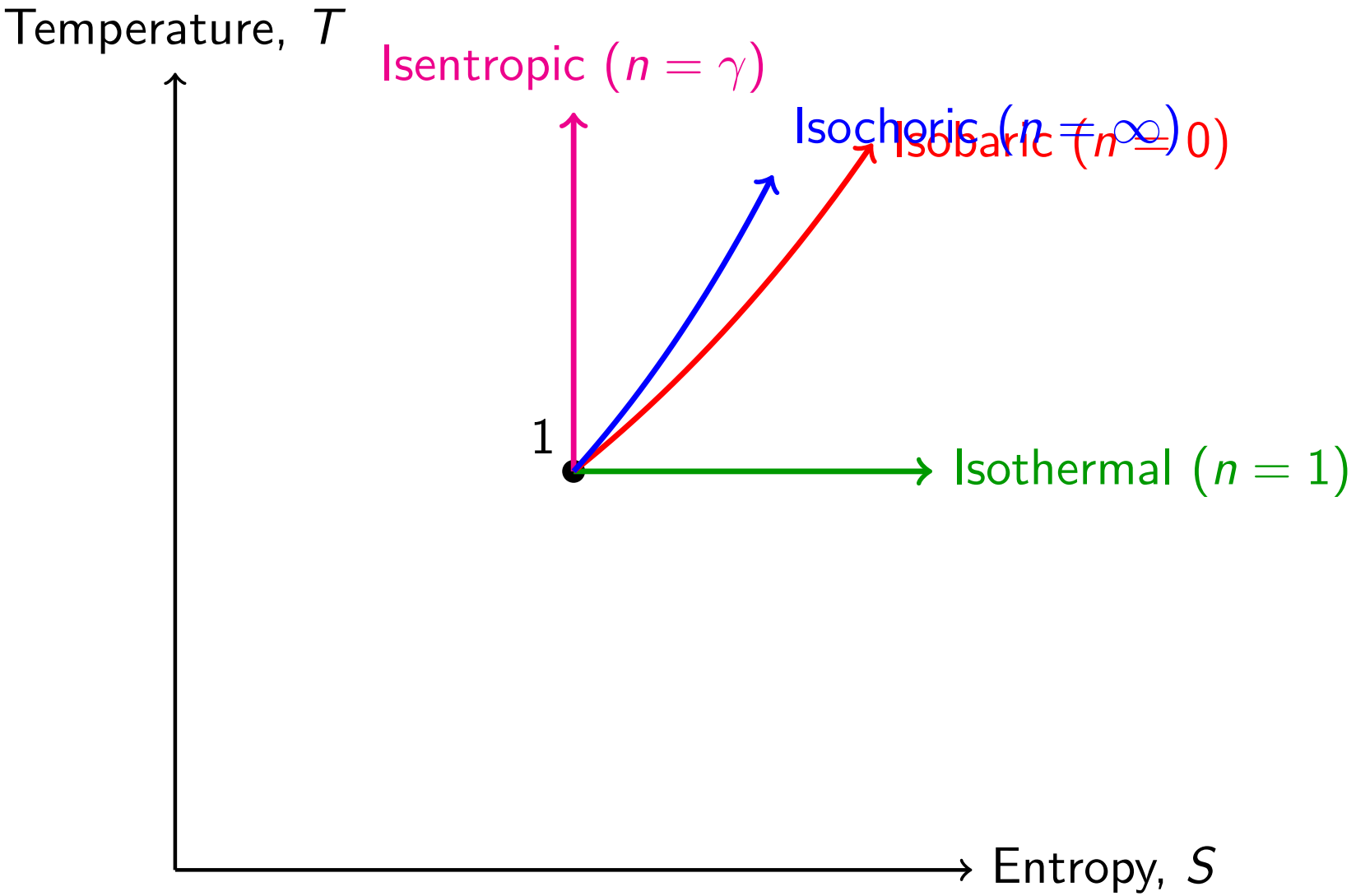
Reversible Adiabatic (Isentropic) Process (n = gamma) I

- Governing Equation: P * V^gamma = Constant, where gamma = Cp/Cv (adiabatic index). Also T * V^(gamma-1) = C and T^gamma * P^(1-gamma) = C.
- Adiabatic Condition: Q = 0 (no heat exchange with the surroundings).
- Isentropic Condition: dS = 0 (since it is reversible and adiabatic, the entropy remains constant, S = C).
- Change in Internal Energy: dU = m * Cv * (T2 - T1).
- Boundary Work: W = (P1*V1 - P2*V2) / (gamma - 1) = m * R * (T1 - T2) / (gamma - 1) = -dU.
- Representation: A curve on the P-V diagram that is steeper than the isothermal line, and a vertical line on the T-S diagram.

Graphical Comparison on the P-V Diagram



Graphical Comparison on the T-S Diagram



Thermodynamic Process Summary Table I

- Process: Isobaric — Index: n=0 — Work (W): P*(V2-V1) — Heat (Q): m*Cp*(T2-T1) — dU: m*Cv*(T2-T1) — dS: m*Cp*ln(T2/T1)
- Process: Isothermal — Index: n=1 — Work (W): P1*V1*ln(V2/V1) — Heat (Q): W — dU: 0 — dS: m*R*ln(V2/V1)
- Process: Polytropic — Index: n — Work (W): (P1*V1 - P2*V2)/(n-1) — Heat (Q): W * (gamma-n)/(gamma-1) — dU: m*Cv*(T2-T1) — dS: m*Cv*(gamma-n)/(gamma-1)*ln(T2/T1)
- Process: Isentropic — Index: n=gamma — Work (W): (P1*V1 - P2*V2)/(gamma-1) — Heat (Q): 0 — dU: m*Cv*(T2-T1) — dS: 0

Thermodynamic Process Summary Table II

- Process: Isochoric — Index: n=infinity — Work (W): 0 — Heat (Q): m*Cv*(T2-T1) — dU: m*Cv*(T2-T1) — dS: m*Cv*ln(T2/T1)
- Understanding these relations and their graphic slopes is fundamental to analyzing thermal systems, cycles (e.g., Carnot, Otto, Diesel), and power-generation machinery.

Thermodynamic Processes, its representation on P-V (Pressure- Volume) and T-S (Temperature-Entropy) diagram: I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 4: Ideal Gases and Thermodynamic Processes
- Lecture 32: Thermodynamic Processes and Diagrammatic Representations

Agenda I

- Core Concepts of Thermodynamic Processes and Path Functions
- Isochoric (Constant Volume) Process and Diagram Analysis
- Isobaric (Constant Pressure) Process and Diagram Analysis
- Isothermal (Constant Temperature) Process and Hyperbolic Paths
- Isentropic (Reversible Adiabatic) Process and Entropic Limits
- Polytropic Process ($PV^n = C$) as a Generalized Model
- Visualizing State Changes on Pressure-Volume (P-V) Diagrams
- Visualizing State Changes on Temperature-Entropy (T-S) Diagrams
- Comparative Summary of Work, Heat, and Property Changes

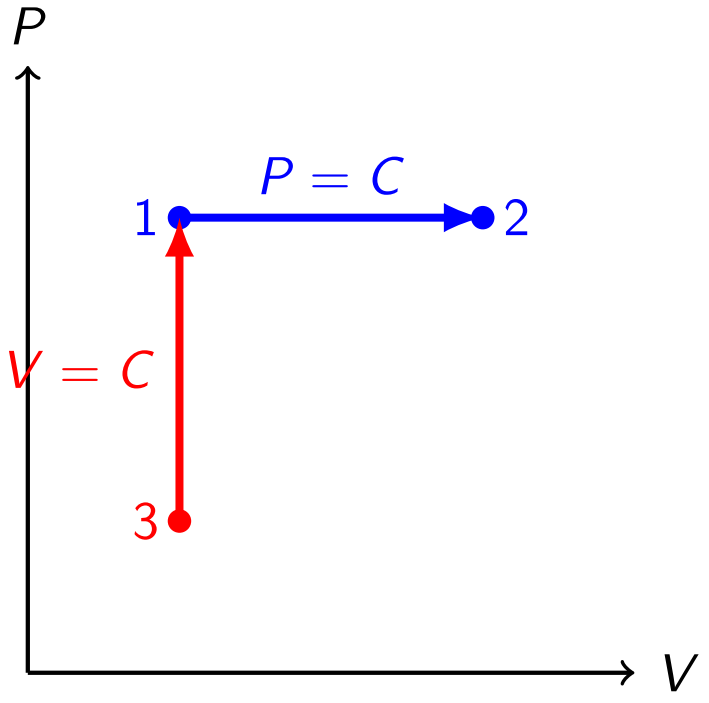
Isochoric (Constant Volume) Process I

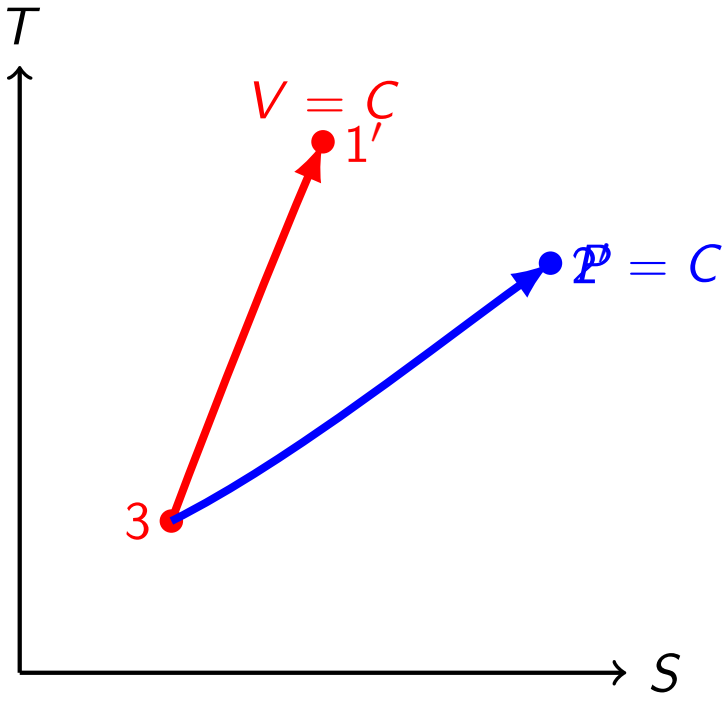
- Definition: A process in which the volume of the system remains constant ($V = C$, $dV = 0$). Also known as isometric or isovolumetric process.
- Governing Equation: Derived from the Ideal Gas Law ($PV = mRT$), indicating that absolute pressure is directly proportional to temperature: $P_1/T_1 = P_2/T_2$.
- Boundary Work: Calculated as $W_{1-2} = \int PdV$. Since $dV = 0$, the boundary work done by or on the system is zero ($W_{1-2} = 0$).
- First Law Application: From $Q - W = \Delta U$, since $W = 0$, all heat added goes into changing internal energy: $Q_{1-2} = \Delta U = mC_v(T_2 - T_1)$.
- Entropy Change: Calculated using $dS = mC_v dT/T$. Integrating yields $\Delta S = mC_v \ln(T_2/T_1)$.

Isobaric (Constant Pressure) Process I

- Definition: A thermodynamic process in which the system pressure remains constant ($P = C$, $dP = 0$).
- Governing Equation: From the Ideal Gas Law ($PV = mRT$), absolute volume is directly proportional to temperature: $V_1/T_1 = V_2/T_2$ (Charles's Law).
- Boundary Work: Computed directly as $W_{1-2} = \int_1^2 PdV = P(V_2 - V_1) = mR(T_2 - T_1)$ for an ideal gas.
- First Law Application: Heat transfer is $Q_{1-2} = \Delta U + W_{1-2} = (U_2 - U_1) + P(V_2 - V_1) = H_2 - H_1 = mC_p(T_2 - T_1)$, expressing heat in terms of enthalpy.
- Entropy Change: Evaluated using $dS = mC_p dT/T$. Integrating gives $\Delta S = mC_p \ln(T_2/T_1)$.

Isochoric & Isobaric Processes on P-V and T-S Diagrams





Isothermal (Constant Temperature) Process I

- Definition: A process in which the temperature of the system remains constant ($T = C$, $dT = 0$). For an ideal gas, internal energy remains constant ($\Delta U = 0$).
- Governing Equation: From the Ideal Gas Law, $PV = mRT = C$. Therefore, pressure is inversely proportional to volume: $P_1 V_1 = P_2 V_2$ (Boyle's Law).
- Boundary Work: Obtained by integrating $W_{1-2} = \int_1^2 P dV = \int_1^2 \frac{C}{V} dV = P_1 V_1 \ln\left(\frac{V_2}{V_1}\right) = mRT \ln\left(\frac{P_1}{P_2}\right)$.
- First Law Application: Since $\Delta U = 0$ for an ideal gas, the first law ($Q_{1-2} - W_{1-2} = \Delta U$) simplifies to $Q_{1-2} = W_{1-2}$.
- Entropy Change: Because T is constant, $\Delta S = Q_{rev}/T = mR \ln(V_2/V_1)$. This is represented by a horizontal line on a T-S diagram.

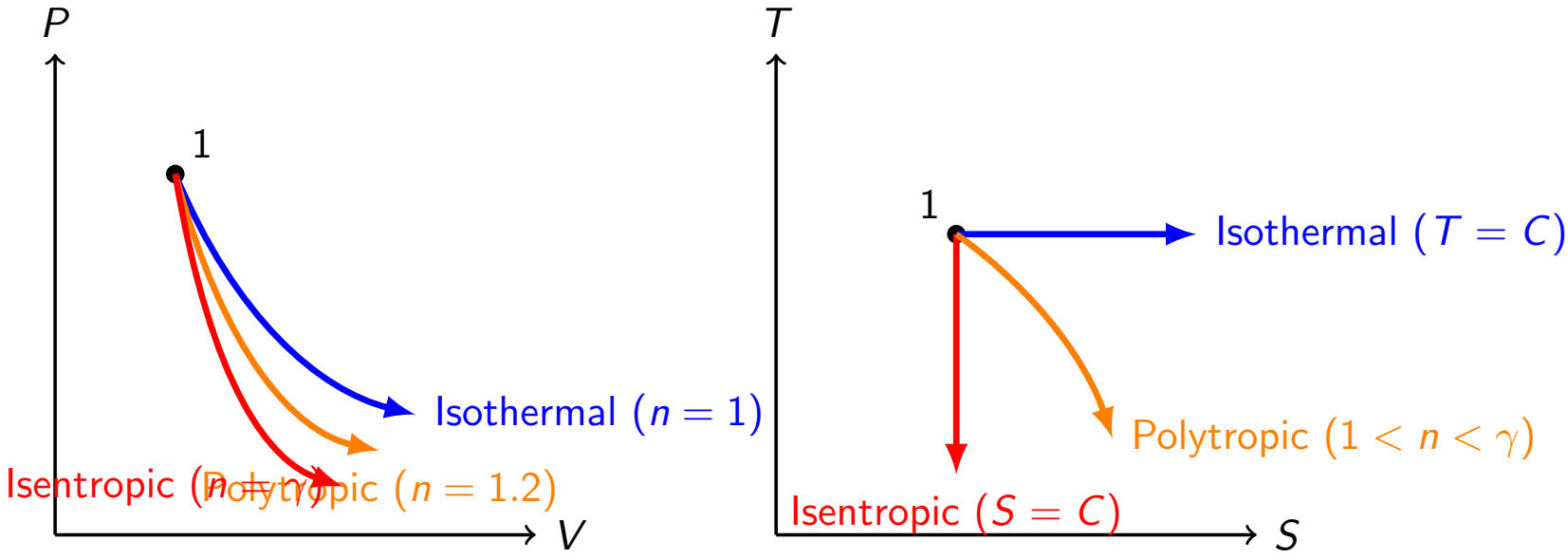
Isentropic (Reversible Adiabatic) Process I

- Definition: A thermodynamic process that is both adiabatic ($Q = 0$) and reversible, resulting in no change in entropy ($S = C$, $dS = 0$).
- Governing Equation: Governed by the relation $PV^{\gamma} = C$, where $\gamma = C_p/C_v$ is the adiabatic index or ratio of specific heats ($\gamma \approx 1.4$ for air).
- Boundary Work: Derived from the process equation: $W_{1-2} = \frac{P_1V_1 - P_2V_2}{\gamma - 1} = \frac{mR(T_1 - T_2)}{\gamma - 1}$.
- First Law Application: With $Q_{1-2} = 0$, the first law reduces to $W_{1-2} = -\Delta U = mC_v(T_1 - T_2)$, meaning expansion work reduces internal energy.
- Entropy Change: By definition, $\Delta S = 0$, which is represented as a vertical line on a Temperature-Entropy (T-S) diagram.

Polytropic Process I

- Definition: A generalized process that represents real-world expansion and compression processes, defined by the equation $PV^n = C$, where n is the polytropic index.
- Physical Scope of n: The index n can range from $-\infty$ to $+\infty$: $n = 0$ is isobaric, $n = 1$ is isothermal, $n = \gamma$ is isentropic, and $n = \infty$ is isochoric.
- Boundary Work: For $n \neq 1$, the work done is given by $W_{1-2} = \frac{P_1V_1 - P_2V_2}{n-1} = \frac{mR(T_1 - T_2)}{n-1}$.
- Heat Transfer: Incorporating work and internal energy: $Q_{1-2} = W_{1-2}\left(\frac{\gamma - n}{\gamma - 1}\right) = mC_n(T_2 - T_1)$, where $C_n = C_v\left(\frac{n-\gamma}{n-1}\right)$.
- P-V and T-S Slopes: The slope on a P-V diagram lies between the isothermal ($n = 1$) and isentropic ($n = \gamma$) curves when $1 < n < \gamma$.

Isothermal, Isentropic & Polytropic Processes



Summary and Comparison of Thermodynamic Processes I

- Isochoric ($V = C$, $n = \infty$): Work $W = 0$, Heat $Q = \Delta U = mC_v\Delta T$. Represented as a vertical line on P-V and a steep logarithmic curve on T-S.
- Isobaric ($P = C$, $n = 0$): Work $W = P\Delta V$, Heat $Q = \Delta H = mC_p\Delta T$. Represented as a horizontal line on P-V and a shallower logarithmic curve on T-S.
- Isothermal ($T = C$, $n = 1$): Work $W = P_1V_1\ln(V_2/V_1)$, Heat $Q = W$. Represented as a hyperbola on P-V and a horizontal line on T-S.
- Isentropic ($S = C$, $n = \gamma$): Work $W = \frac{P_1V_1 - P_2V_2}{\gamma - 1}$, Heat $Q = 0$. Represented as a curve steeper than isothermal on P-V, and a vertical line on T-S.

Summary and Comparison of Thermodynamic Processes II

- Polytropic ($PV^n = C$): Work $W = \frac{P_1V_1 - P_2V_2}{n-1}$, Heat $Q = W\left(\frac{\gamma - n}{\gamma - 1}\right)$. The general process matching real-world paths depending on index n.

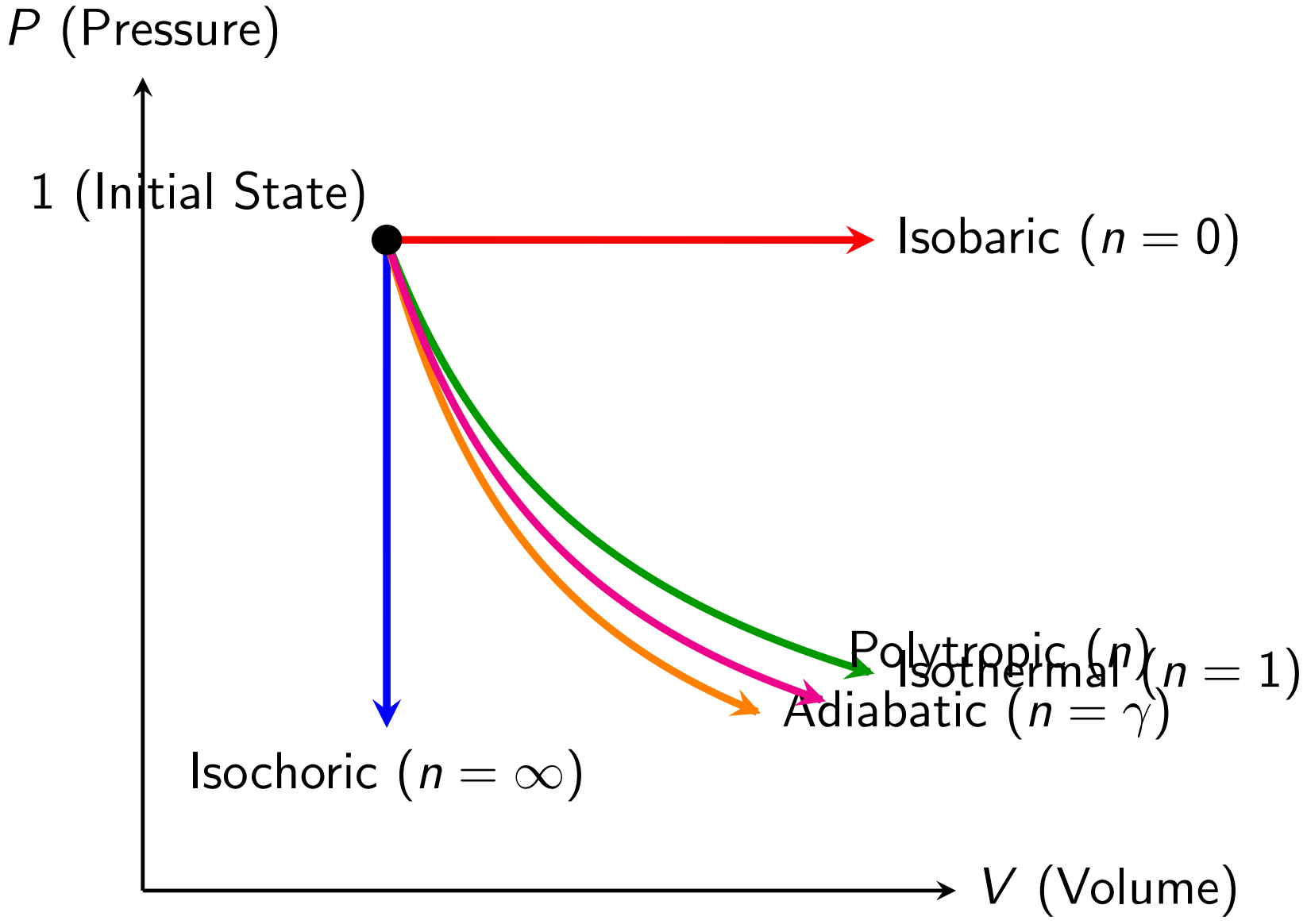
Equations of P-V-T relationship, work transfer, heat transfer and internal energy of the above processes. (Without derivations) I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 4: Ideal Gases and Thermodynamic Processes
- Lecture 33: Non-Flow Process Relations and Energy Interactions
- Department of Mechanical Engineering

Agenda I

- Overview of thermodynamic processes on a P-V diagram
- Isochoric (Constant Volume) process relations and energy transfer
- Isobaric (Constant Pressure) process relations and energy transfer
- Isothermal (Constant Temperature) process relations and energy transfer
- Adiabatic (Isentropic) process relations and energy transfer
- Polytropic (General) process relations and energy transfer
- Comparative summary of equations for P-V-T, W, Q, and delta U

Graphical Representation of Non-Flow Processes



Isochoric Process (Constant Volume, $V = C$) I

- P-V-T Relation: $\frac{P_1}{T_1} = \frac{P_2}{T_2}$ or $\frac{P}{T} = C$ (Gay-Lussac's Law).
- Displacement Work (W): $W_{1-2} = \int_1^2 PdV = 0$ since volume change $dV = 0$.
- Internal Energy Change (ΔU): $\Delta U = mc_v(T_2 - T_1)$ (valid for all ideal gas processes).
- Heat Transfer (Q): From the First Law ($Q = \Delta U + W$), $Q_{1-2} = \Delta U = mc_v(T_2 - T_1)$.
- Enthalpy Change (ΔH): $\Delta H = mc_p(T_2 - T_1)$.

Isobaric Process (Constant Pressure, $P = C$) I

- P-V-T Relation: $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ or $\frac{V}{T} = C$ (Charles's Law).
- Displacement Work (W):
  $W_{1-2} = \int_1^2 PdV = P(V_2 - V_1) = mR(T_2 - T_1)$.
- Internal Energy Change (ΔU): $\Delta U = mc_v(T_2 - T_1)$.
- Heat Transfer (Q): From the First Law,
  $Q_{1-2} = \Delta U + W_{1-2} = mc_v(T_2 - T_1) + mR(T_2 - T_1) = mc_p(T_2 - T_1) = \Delta H$.

Isothermal Process (Constant Temperature, $T = C$) I

- P-V-T Relation: $P_1 V_1 = P_2 V_2$ or $PV = C$ (Boyle's Law).
- Internal Energy Change (ΔU): Since temperature is constant, $\Delta U = mc_v \Delta T = 0$.
- Enthalpy Change (ΔH): Similarly, $\Delta H = mc_p \Delta T = 0$.
- Displacement Work (W): $W_{1-2} = \int_1^2 PdV = P_1 V_1 \ln\left(\frac{V_2}{V_1}\right) = mRT \ln\left(\frac{V_2}{V_1}\right) = P_1 V_1 \ln\left(\frac{P_1}{P_2}\right)$.
- Heat Transfer (Q): From the First Law, $Q_{1-2} = W_{1-2} = mRT \ln\left(\frac{V_2}{V_1}\right)$.

Adiabatic / Isentropic Process ($PV^\gamma = C,\ Q = 0$) I

- Governing Law: $PV^\gamma = C$, where $\gamma = c_p/c_v$ is the ratio of specific heats.
- P-V-T Relations: $P_1 V_1^\gamma = P_2 V_2^\gamma$, $\frac{T_2}{T_1} = \left(\frac{V_1}{V_2}\right)^{\gamma-1}$, and $\frac{T_2}{T_1} = \left(\frac{P_2}{P_1}\right)^{\frac{\gamma-1}{\gamma}}$.
- Heat Transfer (Q): By definition, $Q_{1-2} = 0$.
- Internal Energy Change (ΔU): $\Delta U = mc_v(T_2 - T_1)$.
- Displacement Work (W): $W_{1-2} = \frac{P_1V_1 - P_2V_2}{\gamma-1} = \frac{mR(T_1-T_2)}{\gamma-1} = -\Delta U$.

Polytropic Process ($PV^n = C$) I

- Governing Law: $PV^n = C$, where n is the polytropic index (typically $1 < n < \gamma$ for general expansion).
- P-V-T Relations: $P_1 V_1^n = P_2 V_2^n$, $\frac{T_2}{T_1} = \left(\frac{V_1}{V_2}\right)^{n-1}$, and $\frac{T_2}{T_1} = \left(\frac{P_2}{P_1}\right)^{\frac{n-1}{n}}$.
- Displacement Work (W): $W_{1-2} = \frac{P_1 V_1 - P_2 V_2}{n-1} = \frac{mR(T_1 - T_2)}{n-1}$ (for $n \neq 1$).
- Internal Energy Change (ΔU): $\Delta U = mc_v(T_2 - T_1)$.
- Heat Transfer (Q): $Q_{1-2} = W_{1-2} \times \left(\frac{\gamma - n}{\gamma - 1}\right)$.

System
(Ideal Gas); [-¿, red, ultra thick] (-2,2) – (0,2) node[midway, above] $Q > 0$ (In); [-¿, red, ultra thick] (0,1) – (-2,1) node[midway, below] $Q < 0$ (Out); [-¿, blue, ultra thick] (4,2) – (6,2) node[midway, above] $W > 0$ (Out); [-¿, blue, ultra thick] (6,1) – (4,1) node[midway, below] $W < 0$ (In);
ode at (2,-0.5) [anchor=north, align=center] **First Law of Thermodynamics**: $Q = \Delta U + W$;

Summary of Thermodynamic Process Equations I

- Isochoric ($V = C$): $P/T = C$, $W = 0$, $Q = \Delta U = mc_v\Delta T$
- Isobaric ($P = C$): $V/T = C$, $W = P\Delta V$, $\Delta U = mc_v\Delta T$, $Q = mc_p\Delta T = \Delta H$
- Isothermal ($T = C$): $PV = C$, $\Delta U = 0$, $Q = W = P_1V_1\ln(V_2/V_1)$
- Adiabatic ($PV^\gamma = C$): $Q = 0$, $\Delta U = mc_v\Delta T$, $W = -\Delta U = (P_1V_1 - P_2V_2)/(\gamma - 1)$
- Polytropic ($PV^n = C$): $W = (P_1V_1 - P_2V_2)/(n - 1)$, $Q = W(\gamma - n)/(\gamma - 1)$, $\Delta U = mc_v\Delta T$

Equations of P-V-T relationship, work transfer, heat transfer and internal energy of the above processes. (Without derivations) I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 4: Ideal Gases and Thermodynamic Processes
- Topic: Mathematical modeling of fundamental non-flow processes for ideal gases
- Focus: Direct equations for P-V-T relations, displacement work (W), heat transfer (Q), and internal energy change (dU) without derivation steps.

Agenda I

- Governing assumptions and fundamental thermodynamic equations
- Constant Volume (Isochoric) Process equations
- Constant Pressure (Isobaric) Process equations
- P-V Diagram visualization for Constant Volume and Constant Pressure processes
- Constant Temperature (Isothermal) Process equations
- Reversible Adiabatic (Isentropic) Process equations
- Polytropic Process equations
- Comparative P-V Diagram of all expansion/compression processes

Governing Assumptions & Fundamental Relations I

- Ideal Gas Equation of State: $PV = mRT$, where P is absolute pressure, V is volume, m is mass, R is the specific gas constant, and T is absolute temperature.
- First Law of Thermodynamics (Closed System): $Q - W = \Delta U$, where Q is heat transfer, W is boundary work, and ΔU is change in internal energy.
- Internal Energy Change (Joule's Law): $\Delta U = mC_v(T_2 - T_1)$ is valid for ANY process of an ideal gas.
- Enthalpy Change: $\Delta H = mC_p(T_2 - T_1)$ is valid for ANY process of an ideal gas.
- Specific Heat Relationships: Mayer's Relation is $C_p - C_v = R$, and the ratio of specific heats is defined as $\gamma = C_p/C_v$.

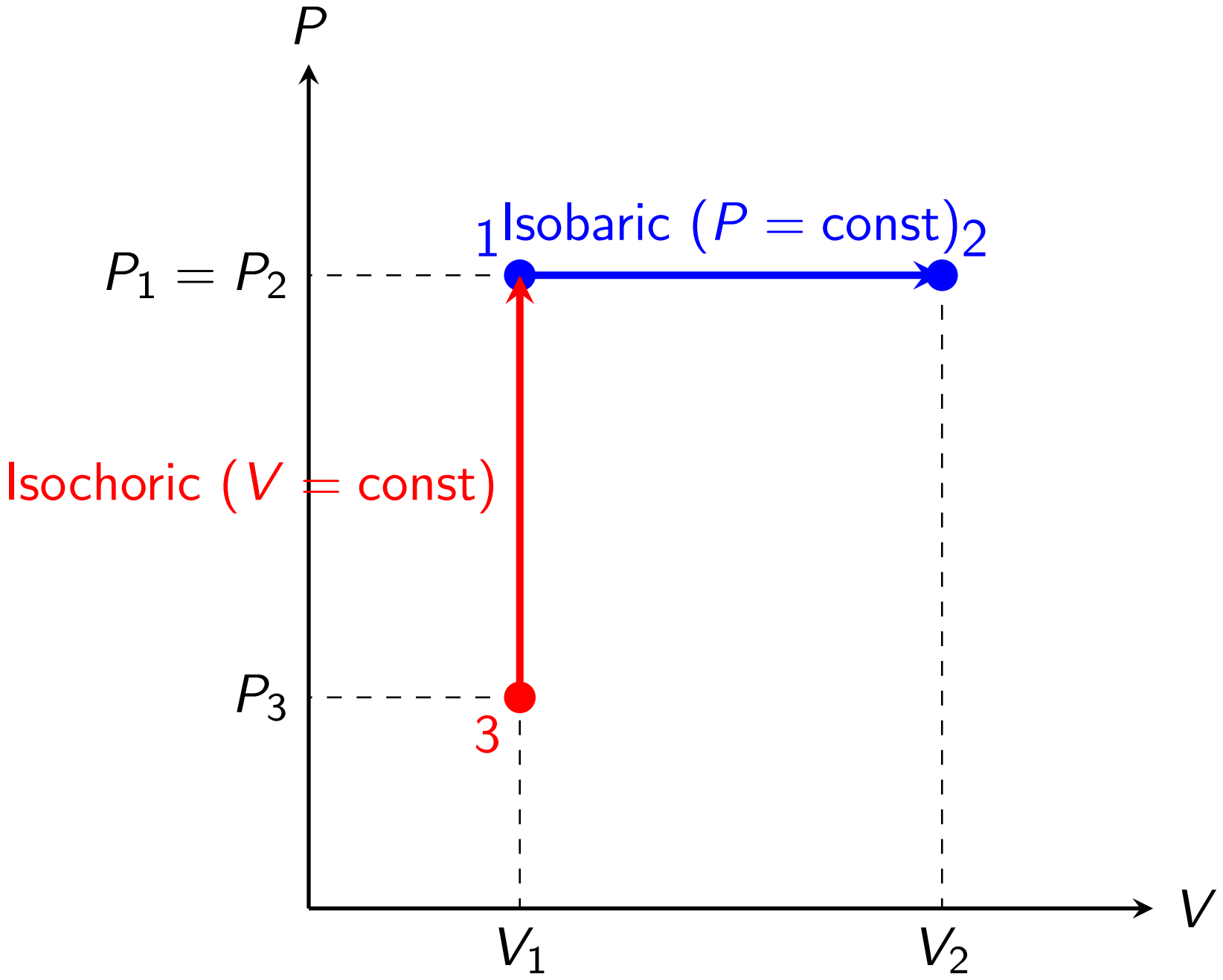
Constant Volume (Isochoric) Process I

- P-V-T Relationship: $\frac{P_1}{T_1} = \frac{P_2}{T_2}$ (since $V_1 = V_2$). Absolute pressure is directly proportional to absolute temperature.
- Work Transfer (W): Boundary work $W = \int_{V_1}^{V_2} PdV = 0$ since $dV = 0$. No displacement work is performed.
- Internal Energy Change (ΔU): $\Delta U = mC_v(T_2 - T_1) = \frac{V(P_2 - P_1)}{\gamma - 1}$.
- Heat Transfer (Q): From the First Law ($Q = W + \Delta U$), since $W = 0$, all heat transferred goes into changing the internal energy: $Q = \Delta U = mC_v(T_2 - T_1)$.

Constant Pressure (Isobaric) Process I

- P-V-T Relationship: $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ (since $P_1 = P_2$). Volume is directly proportional to absolute temperature.
- Work Transfer (W): Boundary work is given by $W = \int_{V_1}^{V_2} PdV = P(V_2 - V_1) = mR(T_2 - T_1)$.
- Internal Energy Change (ΔU): $\Delta U = mC_v(T_2 - T_1) = \frac{P(V_2 - V_1)}{\gamma - 1}$.
- Heat Transfer (Q): $Q = W + \Delta U = P(V_2 - V_1) + mC_v(T_2 - T_1) = mC_p(T_2 - T_1) = \Delta H$. Heat transfer equals change in enthalpy.

P-V Diagram: Isochoric & Isobaric Processes



Constant Temperature (Isothermal) Process I

- P-V-T Relationship: $P_1 V_1 = P_2 V_2$ (since $T_1 = T_2$). Pressure and volume are inversely proportional.
- Work Transfer (W): Boundary work is calculated as $W = \int_{V_1}^{V_2} P dV = P_1 V_1 \ln\left(\frac{V_2}{V_1}\right) = mRT \ln\left(\frac{V_2}{V_1}\right) = P_1 V_1 \ln\left(\frac{P_1}{P_2}\right)$.
- Internal Energy Change (ΔU): Since temperature is constant, $\Delta U = mC_v(T_2 - T_1) = 0$.
- Heat Transfer (Q): From the First Law ($Q = W + \Delta U$), because $\Delta U = 0$, all input heat is fully converted to boundary work: $Q = W = mRT \ln\left(\frac{V_2}{V_1}\right)$.

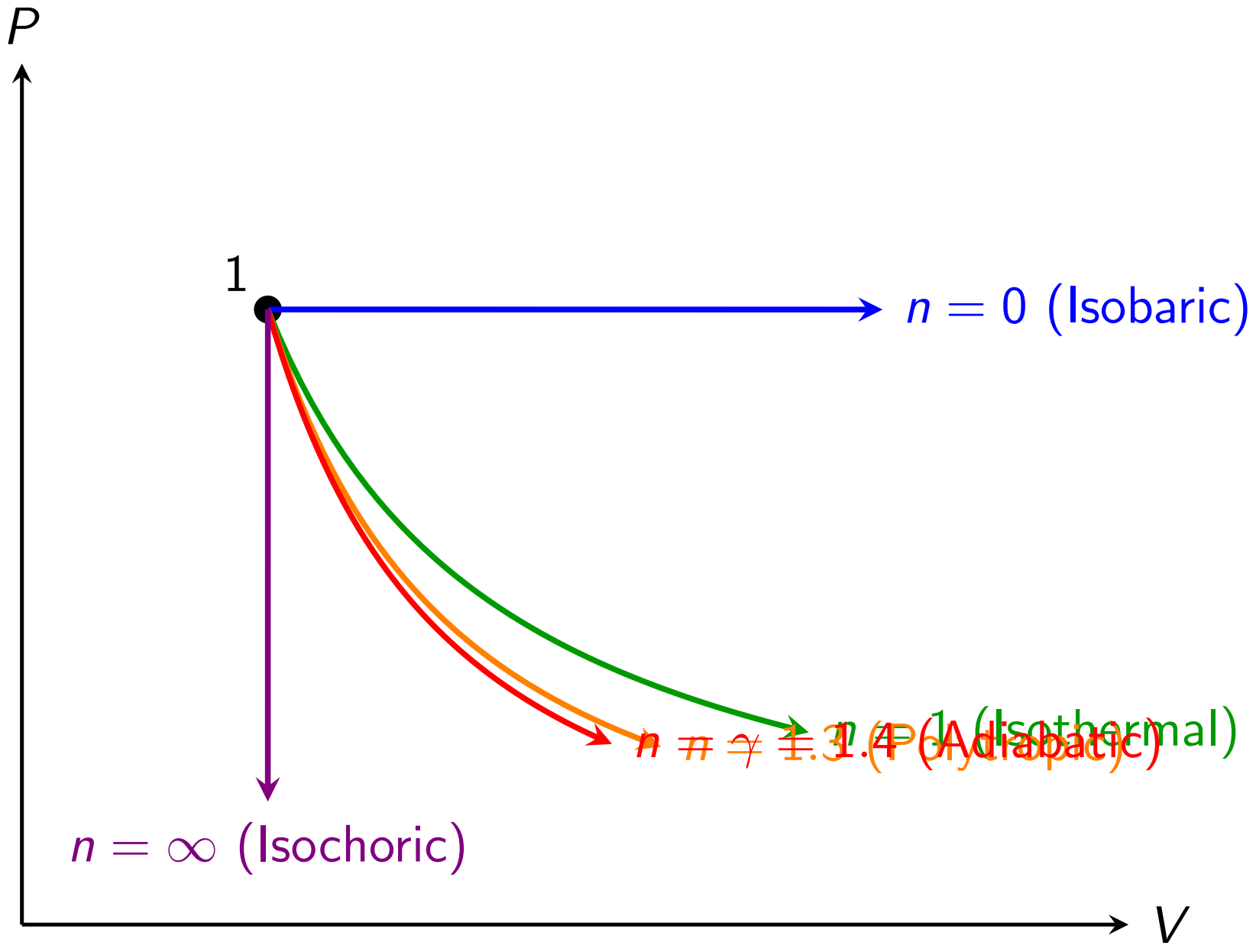
Reversible Adiabatic (Isentropic) Process I

- P-V-T Relationships: Governed by $PV^{\gamma} = \text{constant}$. The temperature-volume relation is $\frac{T_2}{T_1} = \left(\frac{V_1}{V_2}\right)^{\gamma-1}$, and the temperature-pressure relation is $\frac{T_2}{T_1} = \left(\frac{P_2}{P_1}\right)^{\frac{\gamma-1}{\gamma}}$.
- Work Transfer (W): Displacement work is given by $W = \frac{P_1V_1 - P_2V_2}{\gamma-1} = \frac{mR(T_1-T_2)}{\gamma-1}$.
- Internal Energy Change (ΔU): $\Delta U = mC_v(T_2 - T_1) = \frac{P_2V_2 - P_1V_1}{\gamma-1}$. Note that $\Delta U = -W$.
- Heat Transfer (Q): By definition of an adiabatic process, there is no heat exchange with the surroundings: $Q = 0$.

Polytropic Process I

- P-V-T Relationships: Governed by $PV^n = \text{constant}$, where n is the polytropic index (typically $1 < n < \gamma$). Relations:
  $\frac{T_2}{T_1} = \left(\frac{V_1}{V_2}\right)^{n-1}$ and $\frac{T_2}{T_1} = \left(\frac{P_2}{P_1}\right)^{\frac{n-1}{n}}$.
- Work Transfer (W): Boundary work is calculated as
  $W = \frac{P_1V_1 - P_2V_2}{n-1} = \frac{mR(T_1 - T_2)}{n-1}$.
- Internal Energy Change (ΔU):
  $\Delta U = mC_v(T_2 - T_1) = \frac{mR(T_2 - T_1)}{\gamma - 1}$.
- Heat Transfer (Q): From the First Law, heat is transferred along with work:
  $Q = W + \Delta U = \left(\frac{\gamma - n}{\gamma - 1}\right) W = \left(\frac{\gamma - n}{\gamma - 1}\right) \frac{P_1V_1 - P_2V_2}{n-1}$.

Comparison of Processes on P-V Coordinates



Simple Numerical Problems on Thermodynamic Processes I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 4: Ideal Gases and Thermodynamic Processes
- Lecture 35: Simple numerical based on the above.
- Instructor: Professor of Mechanical Engineering

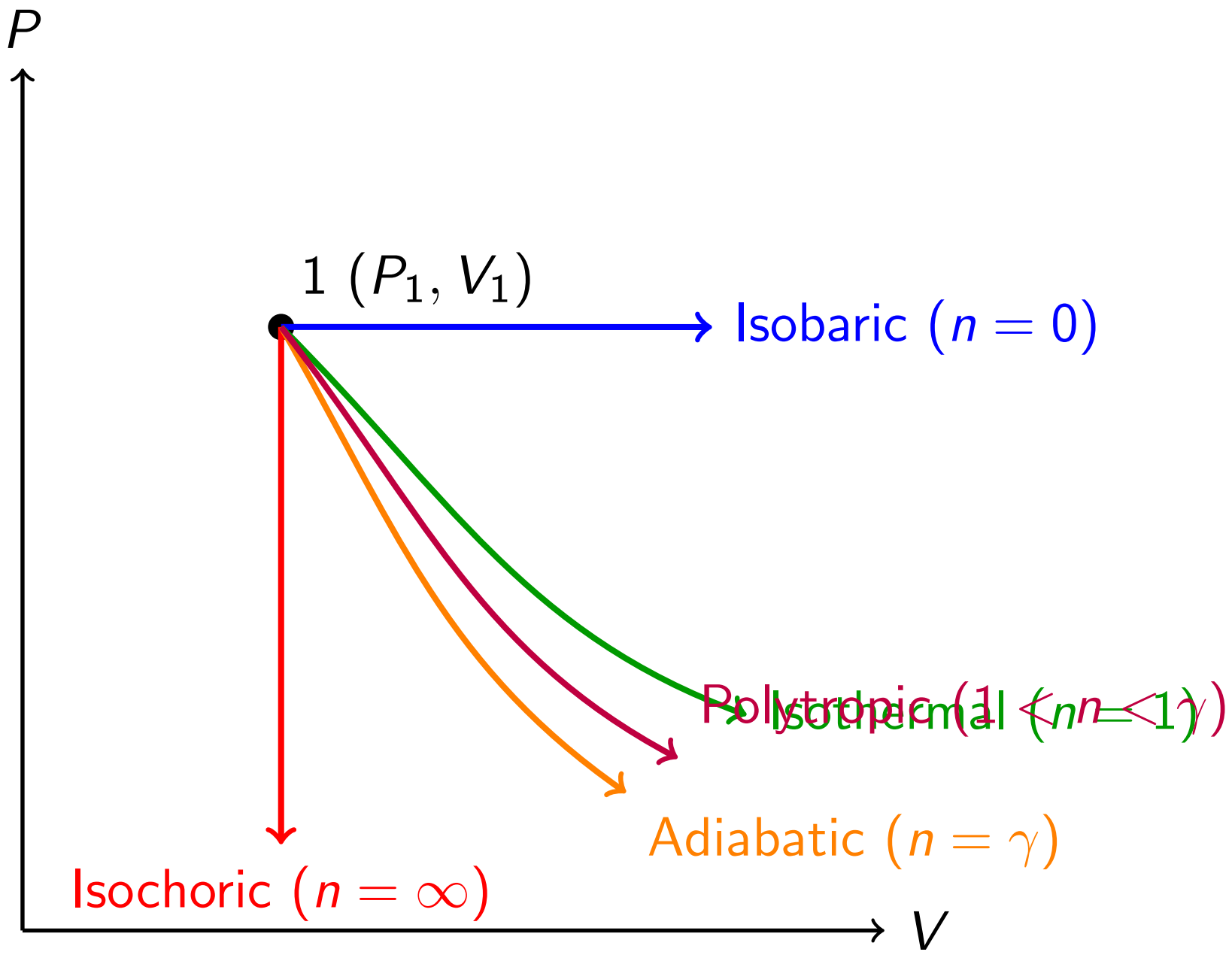
Lecture Agenda I

- Review of governing equations for ideal gas processes (Isobaric, Isochoric, Isothermal, Adiabatic, Polytropic)
- P-V representation of thermodynamic processes
- Numerical Problem 1: Work and heat transfer in a reversible adiabatic expansion
- T-s representation of thermodynamic processes
- Numerical Problem 2: Determination of polytropic index and heat interaction
- Key engineering insights and practical applications of non-flow processes

Summary of Key Thermodynamic Relations I

- Equation of State: P * V = m * R * T, where R is the characteristic gas constant.
- First Law of Thermodynamics for closed systems: Q = dU + W, where dU = m * Cv * dT.
- Isochoric (V=const): W = 0, Q = m * Cv * (T2 - T1).
- Isobaric (P=const): W = P * (V2 - V1), Q = m * Cp * (T2 - T1).
- Isothermal (T=const): W = Q = P1 * V1 * ln(V2 / V1).
- Adiabatic (P*V^gamma = const): Q = 0, W = (P1 * V1 - P2 * V2) / (gamma - 1).
- Polytropic (P*V^n = const): W = (P1 * V1 - P2 * V2) / (n - 1), Q = W * (gamma - n) / (gamma - 1).

Pressure-Volume (P-V) Diagram of Processes



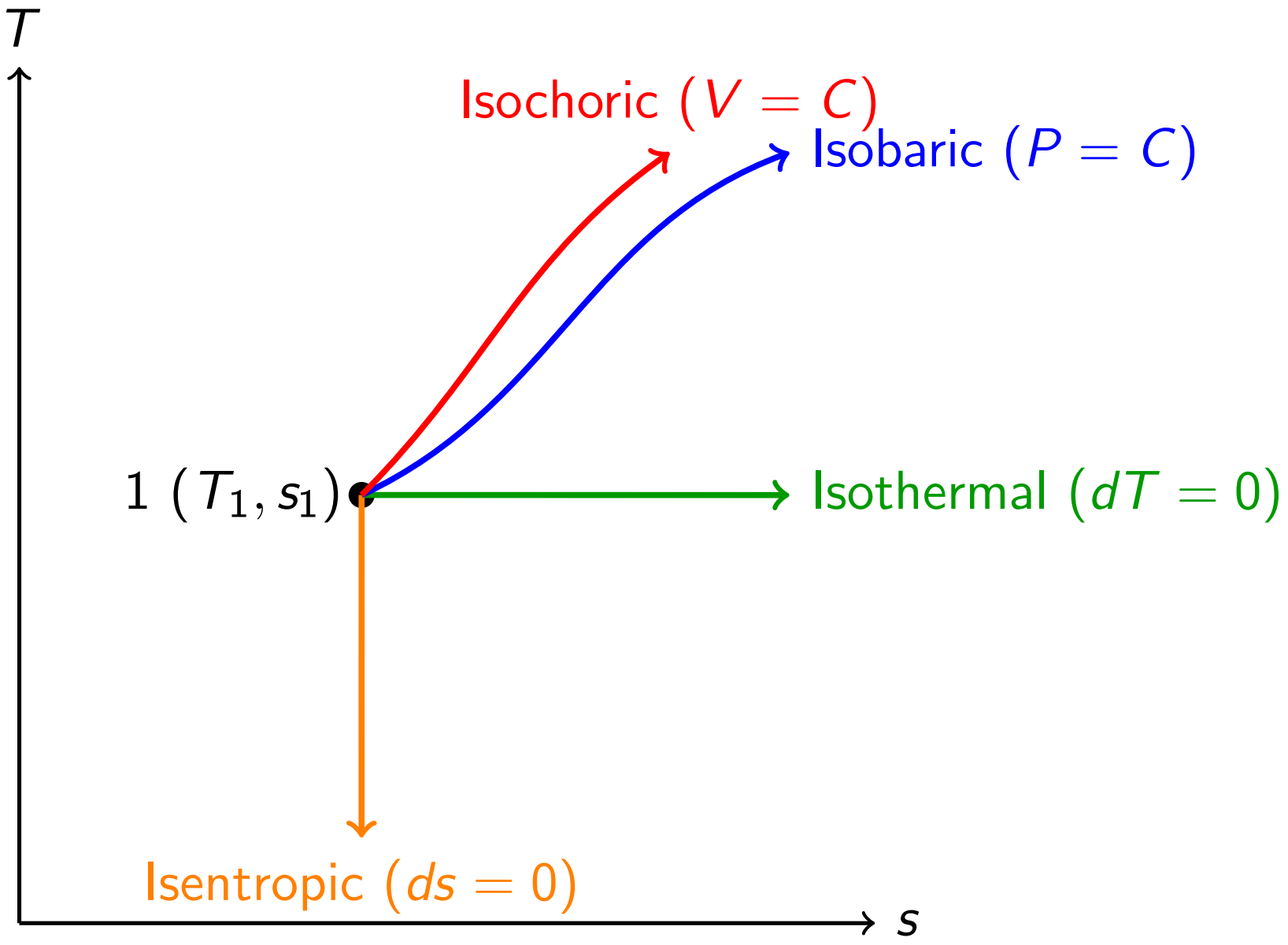
Numerical Problem 1: Reversible Adiabatic Expansion I

- Problem Statement: Air at a pressure of 10 bar and temperature of 600 K expands reversibly and adiabatically in a closed system to a pressure of 1 bar.
- Given properties for air: Characteristic gas constant R = 0.287 kJ/kg-K, Ratio of specific heats gamma = 1.4, Specific heat at constant volume Cv = 0.718 kJ/kg-K.
- Find the following values per unit mass of air (1 kg):
- 1. The final temperature of the air (T2) in Kelvin.
- 2. The work done by the air during expansion (W) in kJ/kg.
- 3. The change in internal energy (dU) in kJ/kg.
- 4. The heat transferred (Q) during the process in kJ/kg.

Solution to Numerical Problem 1 I

- Step 1: Calculate Final Temperature (T2). Using the adiabatic relation between temperature and pressure: T2 = T1 * (P2 / P1)^((gamma - 1)/gamma).
- Substitute values: T2 = 600 * (1 / 10)^((1.4 - 1)/1.4) = 600 * (0.1)^0.2857 = 600 * 0.5179 = 310.8 K.
- Step 2: Calculate Heat Transfer (Q). Since the process is reversible and adiabatic, Q = 0 kJ/kg.
- Step 3: Calculate Change in Internal Energy (dU). Using dU = Cv * (T2 - T1): dU = 0.718 * (310.8 - 600) = 0.718 * (-289.2) = -207.6 kJ/kg.
- Step 4: Calculate Work Done (W). Using the First Law: Q = dU + W =¿ W = Q - dU = 0 - (-207.6) = +207.6 kJ/kg.
- Alternative Work Calculation: W = R * (T1 - T2) / (gamma - 1) = 0.287 * (600 - 310.8) / 0.4 = 0.287 * 289.2 / 0.4 = 207.5 kJ/kg.

Temperature-Entropy (T-s) Diagram of Processes



Numerical Problem 2: Polytropic Process Analysis I

- Problem Statement: 0.5 kg of Nitrogen gas (N2) is compressed from P1 = 1 bar, T1 = 300 K to a final pressure P2 = 8 bar according to the law P*V^1.3 = constant.
- Given properties for Nitrogen: Molecular weight M = 28 kg/kmol (universal gas constant Ru = 8.314 kJ/kmol-K), Ratio of specific heats gamma = 1.4.
- Determine the following parameters:
- 1. The polytropic index (n) for the compression process.
- 2. The final temperature of the gas (T2) in Kelvin.
- 3. The work input required for compression (W) in kJ.
- 4. The heat transfer (Q) and its direction (heat rejected/absorbed) in kJ.

Solution to Numerical Problem 2 I

- Step 1: Identify polytropic index (n). The process law is P*V^1.3 = C, hence polytropic index n = 1.3.
- Step 2: Calculate characteristic gas constant (R). R = Ru / M = 8.314 / 28 = 0.2969 kJ/kg-K.
- Step 3: Calculate final temperature (T2). Using T2 = T1 * (P2 / P1)^((n - 1)/n): T2 = 300 * (8 / 1)^((1.3 - 1)/1.3) = 300 * (8)^0.2308 = 300 * 1.613 = 483.9 K.
- Step 4: Calculate work done (W). W = m * R * (T1 - T2) / (n - 1) = 0.5 * 0.2969 * (300 - 483.9) / (1.3 - 1) = 0.14845 * (-183.9) / 0.3 = -91.0 kJ. (Negative sign indicates work is done ON the gas).
- Step 5: Calculate heat transfer (Q). Q = W * (gamma - n) / (gamma - 1) = -91.0 * (1.4 - 1.3) / (1.4 - 1) = -91.0 * 0.1 / 0.4 = -22.75 kJ. (Negative sign indicates heat is rejected to the surroundings).

Summary and Key Engineering Takeaways I

- The polytropic index 'n' dictates the path and energy distribution of a non-flow process.
- When n = gamma, the process is adiabatic and no heat transfer occurs (Q = 0).
- When 1 ¡ n ¡ gamma, heat rejection occurs during compression, reducing the required work input compared to an adiabatic process.
- Isothermal compression (n = 1) represents the theoretical minimum work input required for a compression process.
- Always pay close attention to signs: negative work implies compression (work in), and negative heat implies heat rejection (heat out).
- Understanding these simple numerical problems forms the foundation for analyzing complex thermodynamic cycles (e.g., Gas Turbine, IC Engine).

Classifications of Thermodynamic cycles I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 5: Thermodynamic Cycles
- Lecture 36: Classifications of Thermodynamic Cycles
- Instructor: Professor of Mechanical Engineering

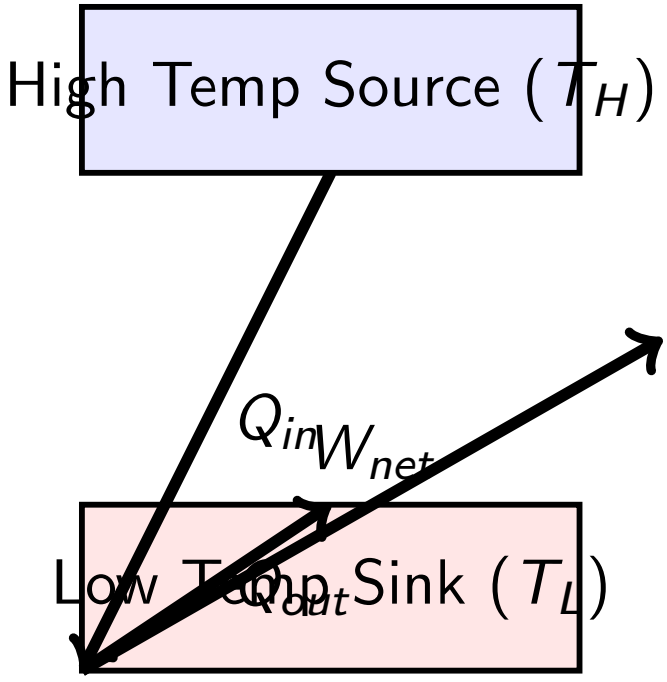
Agenda I

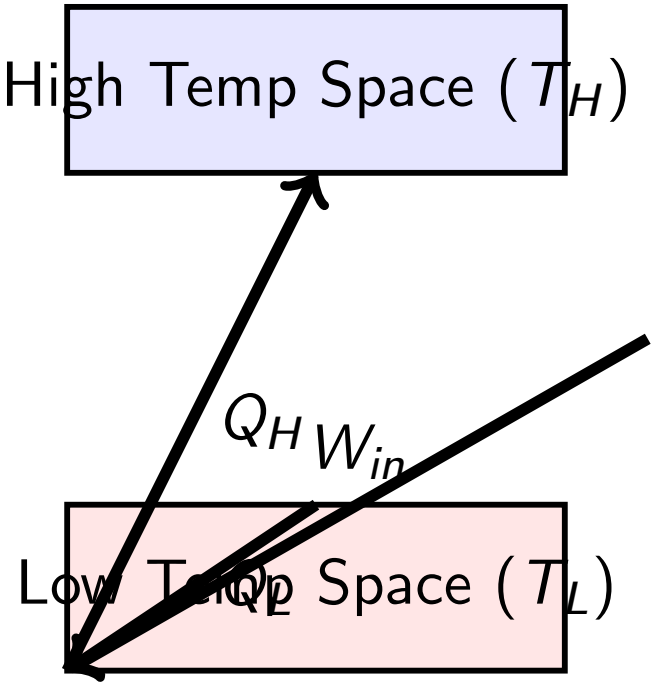
- Introduction to Thermodynamic Cycles and Processes
- Power Cycles vs. Refrigeration/Heat Pump Cycles
- Gas Power Cycles vs. Vapor Power Cycles
- Closed vs. Open Cycles (System Configurations)
- Reversible vs. Irreversible Cycles (Ideal vs. Actual)
- Summary and Key Cycle Classification Matrix

Power Cycles vs. Refrigeration Cycles I

- Thermodynamic cycles are broadly categorized by their primary energy conversion objective: Power Generation or Refrigeration/Heating.
- Power Cycles (Heat Engines): Net work output ($W_{net} > 0$) is produced from a net heat input ($Q_{in} > 0$) from a high-temperature source. Thermal efficiency is defined as $\eta_{th} = W_{net}/Q_{in}$.
- Refrigeration/Heat Pump Cycles (Reverse Heat Engines): Require a net work input ($W_{net} < 0$) to transfer heat from a low-temperature space (Q_L) to a high-temperature space (Q_H).
- Performance metric is the Coefficient of Performance (COP). For Refrigerators: $COP_R = Q_L/W_{net}$. For Heat Pumps: $COP_{HP} = Q_H/W_{net}$.

Energy Flow Diagrams: Heat Engine vs. Refrigerator

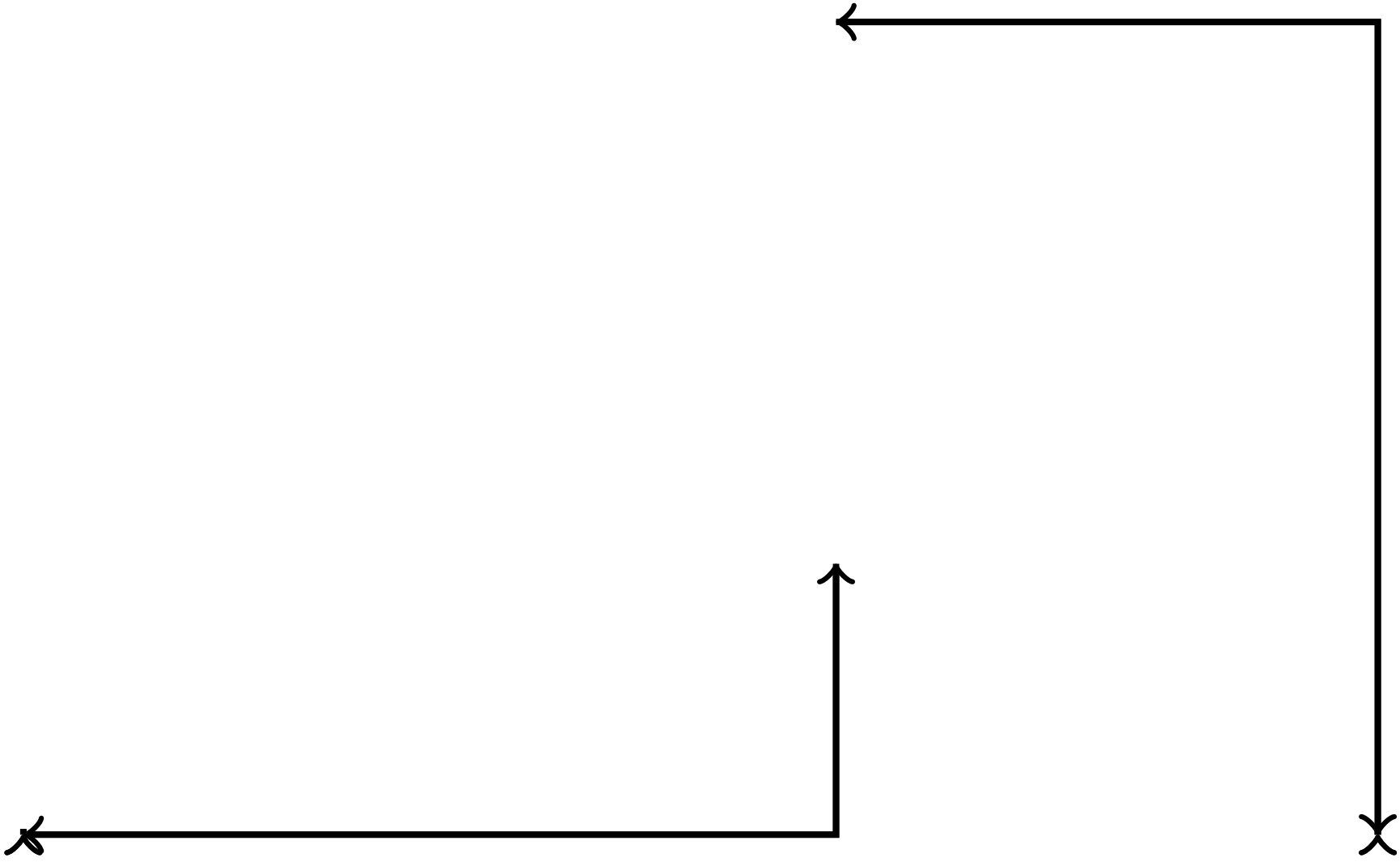




Gas Power Cycles vs. Vapor Power Cycles I

- Cycles are classified by the phase behavior of the working fluid throughout the process steps.
- Gas Power Cycles: The working fluid remains in the gas phase throughout the entire cycle (e.g., air in Otto, Diesel, Brayton, and Stirling cycles).
- Characteristics: Compression and expansion processes take place on gases; high peak temperatures; moderate pressures; requires larger compressor work.
- Vapor Power Cycles: The working fluid undergoes phase changes (liquid to vapor and vice versa) during the cycle (e.g., water/steam in Rankine cycles).
- Characteristics: Latent heat transfer occurs during vaporization/condensation; compression is performed on liquids (pumps), requiring very low power input.

Schematics: Gas (Brayton) vs. Vapor (Rankine) Cycles



Open vs. Closed Thermodynamic Cycles I

- Cycles are classified by whether the physical working fluid is recirculated or discharged.
- Closed Cycles: The working fluid is continuously recirculated within a closed loop (e.g., steam in a power plant boiler-turbine-condenser loop). Heat transfer occurs through boundaries without mass transfer.
- Open Cycles: The working fluid is drawn from the surroundings, undergoes thermodynamic changes, and is discharged back to the environment (e.g., internal combustion engines, open gas turbines).
- Analysis of Open Cycles: Often modeled as equivalent closed cycles using the 'Air-Standard Assumptions' to simplify analysis by replacing combustion and exhaust steps with heat exchangers.

Ideal vs. Actual Cycles I

- Cycles can be categorized by the degree of thermodynamic reversibility of their processes.
- Idealized (Reversible) Cycles: Composed entirely of internally reversible processes (e.g., Carnot cycle, ideal Otto cycle). Serve as upper limits for efficiency and COP.
- Friction, rapid expansion, and heat transfer across finite temperature differences are neglected in ideal cycles to simplify mathematical modeling.
- Actual (Irreversible) Cycles: Real-world cycles containing irreversibilities such as fluid friction, heat loss, and mechanical friction.
- Engineers analyze actual cycles using process-level efficiencies (e.g., isentropic efficiency of turbines and compressors) relative to their ideal equivalents.

System Boundary Interpretations of Cycles I

- Thermodynamic cycle definition: A sequence of processes that returns the system to its initial state ($u_{final} = u_{initial}$, $s_{final} = s_{initial}$).
- In a closed cycle, the system boundary is fixed around the mass of the working fluid, meaning $\oint dE = 0$, leading to $Q_{net} = W_{net}$.
- In an open cycle (e.g., open-system gas turbine), the thermodynamic cycle is completed by the atmosphere acting as a heat sink.
- Applying the First Law to any complete cycle yields: $\sum Q_{in} - \sum Q_{out} = \sum W_{out} - \sum W_{in}$.

Summary Matrix of Cycle Classifications I

- Power Generation Cycles: Categorized into Gas (Otto, Diesel, Dual, Brayton, Stirling, Ericsson) and Vapor (Rankine, Carnot vapor).
- Refrigeration/Heat Pump Cycles: Categorized into Vapor Compression, Absorption, and Gas Refrigeration (reversed Brayton).
- Theoretical Limits: The Carnot cycle represents the maximum theoretical efficiency limit for any power cycle operating between T_H and T_L.
- Design Selection: Choosing a cycle depends on heat source temperature, available heat sink, working fluid constraints, and desired output profile (stationary vs. mobile power).

Classifications of Thermodynamic cycles I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 5: Thermodynamic Cycles
- Lecture 37: Classifications of Thermodynamic Cycles
- Instructor: Professor of Mechanical Engineering
- Overview: Fundamental concepts, definitions, and taxonomic classification of power and refrigeration cycles.

Lecture Agenda I

- Introduction to Thermodynamic Cycles & Definitions
- Classification based on Net Work: Power vs. Refrigeration Cycles
- Classification based on Phase of Working Fluid: Gas vs. Vapor Cycles
- Classification based on System Boundary: Open vs. Closed Cycles
- Carnot Cycle as the Thermodynamic Standard
- Idealized vs. Actual Cycles & Air-Standard Assumptions
- Summary and Comparative Matrix

Thermodynamic Cycle Fundamentals I

- Definition: A thermodynamic cycle is a sequence of processes that begins and ends at the same thermodynamic state.
- State Postulate: Since the initial and final states are identical, the net change in any state property (e.g., internal energy, enthalpy, entropy) over a complete cycle is zero ($dE = 0$, $dS = 0$).
- First Law of Thermodynamics for a Cycle: The net work output equals the net heat input: $W_{net} = Q_{in} - Q_{out}$.
- Thermal Efficiency: For power-producing cycles, efficiency is defined as $\eta_{th} = W_{net}/Q_{in} = 1 - Q_{out}/Q_{in}$.
- Significance: Engineering systems operate on cycles to continuously convert heat into work (e.g., steam power plants) or transfer heat against temperature gradients (e.g., refrigerators).

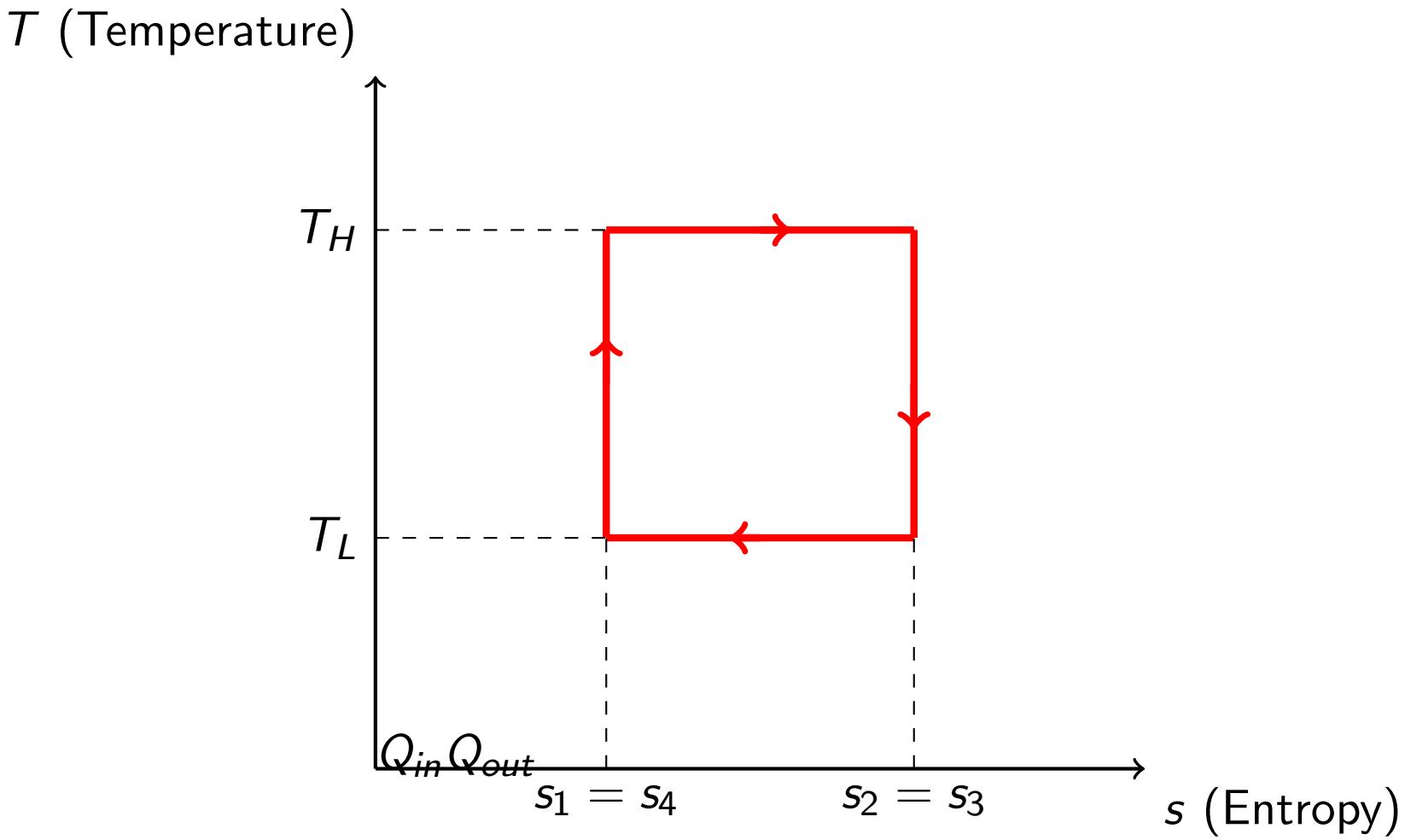
Power vs. Refrigeration Cycles I

- Power Cycles: Designed to produce a net work output ($W_{net} > 0$) from a heat input (Q_{in}) from a high-temperature source. Examples include Rankine, Brayton, Otto, and Diesel cycles.
- Refrigeration/Heat Pump Cycles: Designed to transfer heat from a low-temperature medium ($Q_{in} = Q_L$) to a high-temperature medium ($Q_{out} = Q_H$) by consuming external work ($W_{in} > 0$).
- Direction of Processes: Power cycles run clockwise on a T-s or P-v diagram, whereas refrigeration/heat pump cycles run counter-clockwise.
- Performance Metric: Refrigeration systems are evaluated by Coefficient of Performance ($COP_R = Q_L/W_{in}$), and Heat Pumps by ($COP_{HP} = Q_H/W_{in}$).

Power vs. Refrigeration Cycles II

- Thermodynamic Limit: The Second Law dictates that no power cycle can have 100% efficiency, and no refrigeration cycle can have an infinite COP.

The Carnot Power Cycle T-s Diagram



Gas vs. Vapor Power Cycles I

- Working Fluid Behavior: Cycles are categorized based on whether the working fluid remains a gas or undergoes phase changes.
- Gas Power Cycles: The working fluid remains in the gaseous phase throughout the entire cycle (e.g., Otto, Diesel, Stirling, Brayton cycles).
- Vapor Power Cycles: The working fluid undergoes phase changes between liquid and vapor phases (e.g., Rankine steam cycle, vapor-compression refrigeration cycle).
- Component Differences: Gas cycles use compressors and turbines/pistons; Vapor cycles use pumps, boilers, turbines, and condensers.
- Thermal Characteristics: Vapor cycles can achieve high heat transfer rates during isobaric-isothermal phase transitions (condensation and boiling).

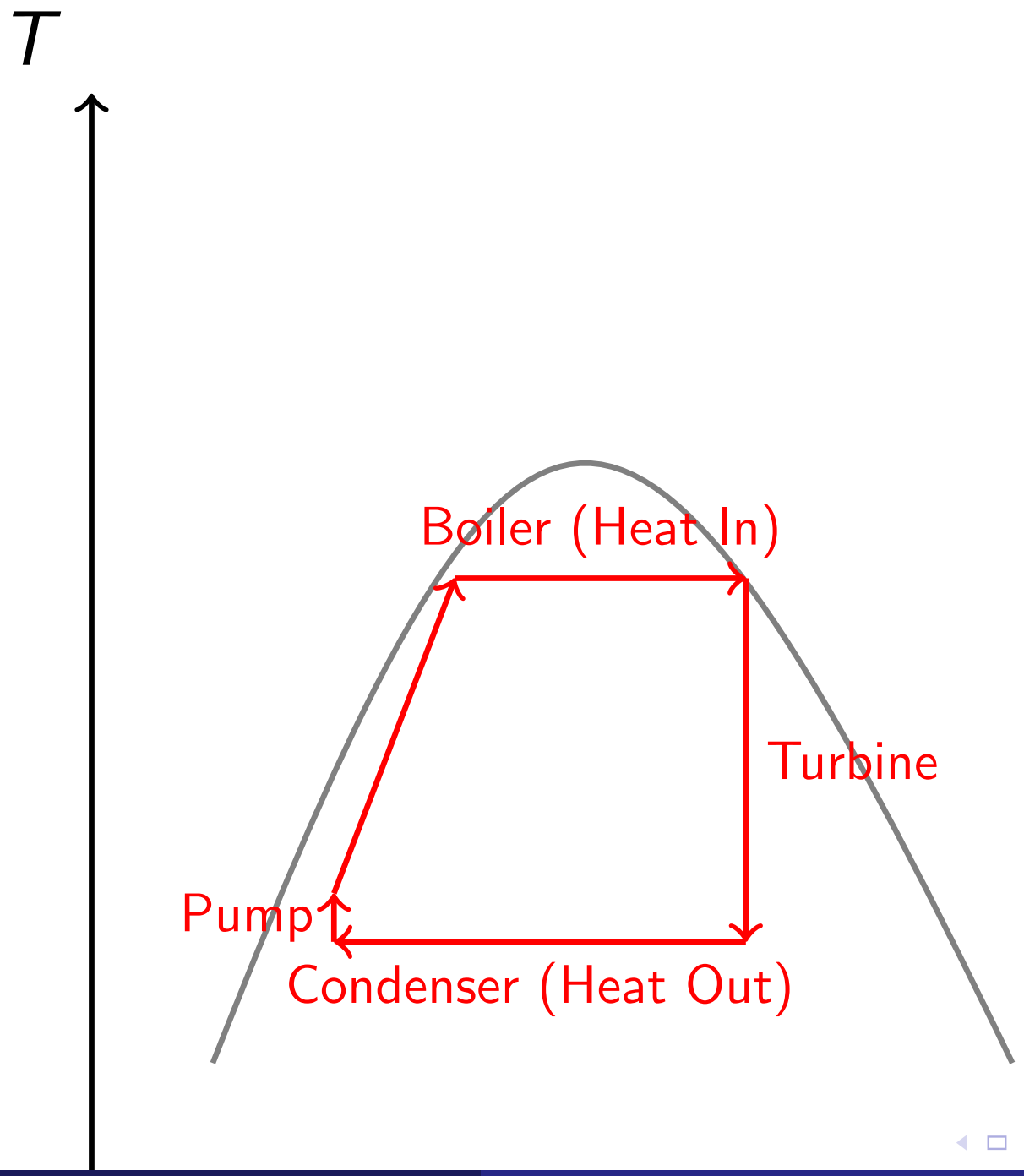
Open vs. Closed Cycles I

- Closed Cycles: The working fluid is continuously recirculated within a closed loop without mass exchange with the surroundings (e.g., Steam power plants, domestic refrigerators).
- Open Cycles: The working fluid is drawn from the surroundings, undergoes thermodynamic processes, and is exhausted back to the atmosphere (e.g., Internal Combustion Engines, open gas turbine systems).
- Thermodynamic Idealization: To analyze open cycles using classical thermodynamics, they are often modeled as closed cycles by replacing the exhaust/intake processes with a heat rejection process back to the inlet state (Air-Standard assumptions).

Open vs. Closed Cycles II

- Mass Flow Characteristics: Closed cycles operate under fixed mass systems ($m = const$), while open cycles are analyzed using steady-flow control volumes with mass flow rates ($\dot{m}$).

Rankine Cycle T-s Diagram with Vapor Dome



Idealized Cycles & Air-Standard Assumptions I

- Idealization Goal: Eliminate complex real-world irreversibilities (friction, pressure drops, heat losses) to establish the upper thermodynamic limit of performance.
- Air-Standard Assumption 1: The working fluid is air, which behaves as an ideal gas throughout the entire cycle.
- Air-Standard Assumption 2: All processes are internally reversible.
- Air-Standard Assumption 3: The combustion process is replaced by a heat addition process from an external source.
- Cold Air-Standard Assumption: Air properties are assumed constant at room temperature ($C_p = 1.005$ kJ/kg$\cdot$ K, $C_v = 0.718$ kJ/kg$\cdot$ K, $k = 1.4$).

Summary of Cycle Classifications I

- Purpose-Based: Power producing ($W_{net} > 0$) vs. Refrigeration/Heat Pump ($W_{net} < 0$).
- Phase-Based: Gas cycles (no phase change, e.g. Otto, Brayton) vs. Vapor cycles (phase change, e.g. Rankine).
- System-Based: Open cycles (exchanging mass, e.g. IC Engines) vs. Closed cycles (fixed mass, e.g. Steam plants).
- Idealization: Actual cycles have irreversibilities (entropy generation, pressure drops); idealized cycles establish theoretical maximum performance.
- Next Steps: Moving forward, we will analyze each major cycle category individually, starting with the Gas Power Cycles (Otto and Diesel).

Carnot cycle and its representation on P-V and T-s diagram. I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 5: Thermodynamic Cycles
- Lecture 38: Carnot cycle and its representation on P-V and T-s diagram.
- Instructor: Professor of Mechanical Engineering

Agenda I

- Introduction to the Carnot cycle and its historical significance in thermodynamics.
- Step-by-step description of the four reversible processes composing the cycle.
- Graphical representation and analysis on the Pressure-Volume (P-V) plane.
- Graphical representation and analysis on the Temperature-Entropy (T-s) plane.
- Mathematical derivation of the Carnot cycle thermal efficiency.
- Implications of the Carnot efficiency as the theoretical upper limit for heat engines.
- Practical limitations preventing the physical realization of a Carnot engine.

Introduction to the Carnot Cycle I

- The Carnot cycle was proposed in 1824 by French physicist Nicolas Léonard Sadi Carnot.
- It is a theoretical thermodynamic cycle that provides the maximum possible thermal efficiency for any engine operating between two temperatures.
- The cycle is completely reversible, meaning all internal processes are reversible and there are no external irreversibilities like friction or unrestrained expansion.
- It serves as the ultimate benchmark or standard of comparison for all actual heat engines.
- Any engine operating on the Carnot cycle is called a Carnot Heat Engine.
- The cycle consists of four distinct processes: two isothermal processes and two adiabatic (isentropic) processes.

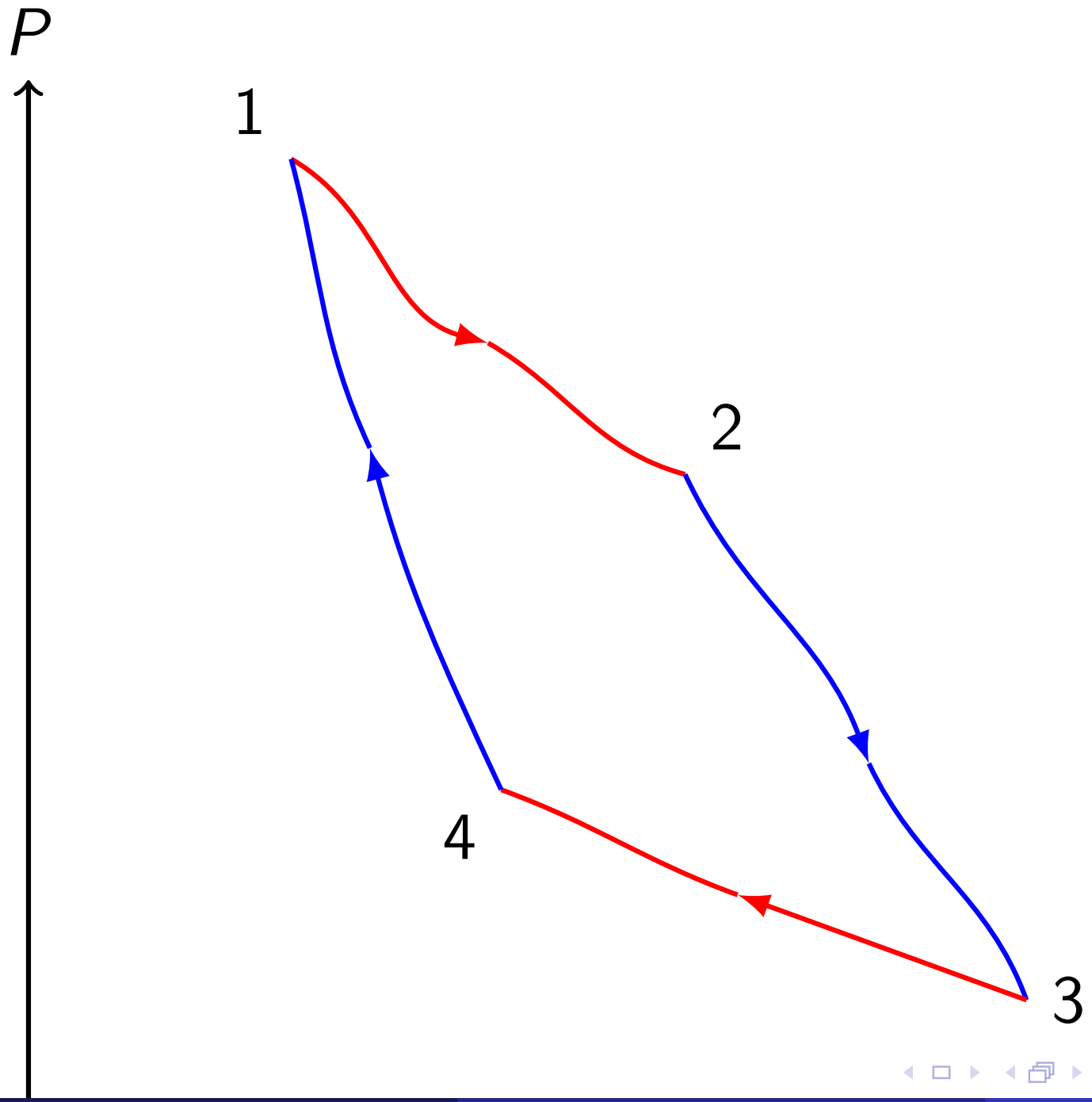
Processes 1-2 and 2-3: Expansion Phase I

- Process 1-2: Reversible Isothermal Expansion. The working fluid (e.g., gas in a piston-cylinder) is placed in contact with a heat source at constant temperature T_H.
- As the gas expands, it does work on the surroundings. To maintain a constant temperature T_H, heat Q_H (Q_{in}) is transferred from the source to the gas.
- Process 2-3: Reversible Adiabatic (Isentropic) Expansion. The cylinder is thermally insulated from both the source and the sink.
- The gas continues to expand, doing work on the surroundings, causing its temperature to drop from T_H to the sink temperature T_L.
- Since the process is reversible and adiabatic, the entropy of the fluid remains constant ($s_2 = s_3$).

Processes 1-2 and 2-3: Expansion Phase II

- The work output during these expansion phases forms the primary energy delivery mechanism of the cycle.

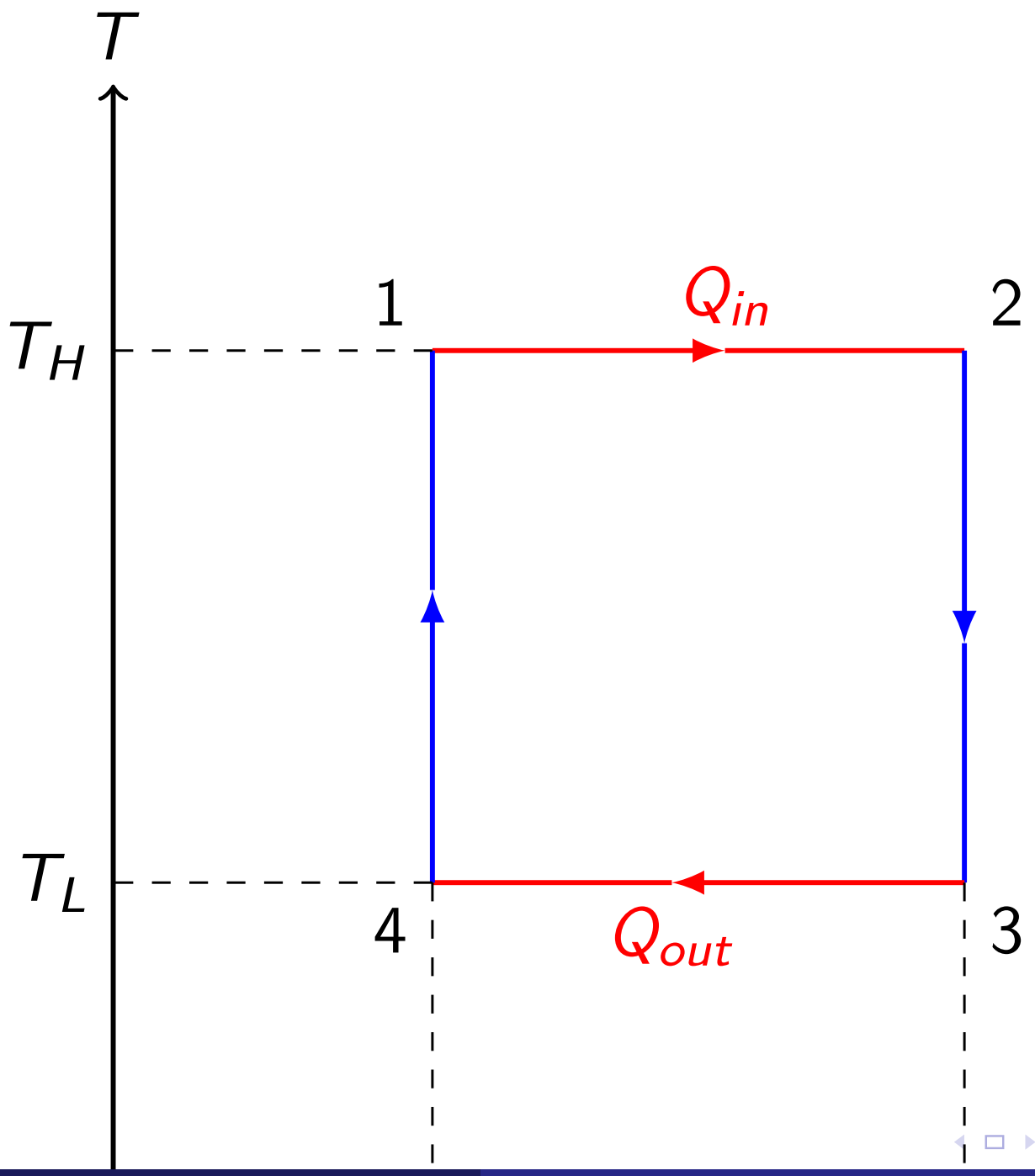
Carnot Cycle on P-V Diagram



Processes 3-4 and 4-1: Compression Phase I

- Process 3-4: Reversible Isothermal Compression. The cylinder is brought into contact with a heat sink at constant temperature T_L.
- The gas is compressed slowly, doing work on the gas. Heat Q_L (Q_{out}) is rejected from the gas to the sink to maintain a constant temperature T_L.
- Process 4-1: Reversible Adiabatic (Isentropic) Compression. The cylinder is insulated again, preventing any heat transfer.
- The gas is compressed further in a reversible manner, causing its temperature to rise from T_L to T_H, completing the cycle.
- Entropy remains constant during this step ($s_4 = s_1$).
- This stage returns the working fluid to its initial state, preparing it for the next cycle of operation.

Carnot Cycle on T-s Diagram



Efficiency of the Carnot Cycle I

- The thermal efficiency of any heat engine is defined as: $\eta_{th} = \frac{W_{net}}{Q_{in}} = 1 - \frac{Q_{out}}{Q_{in}}$.
- For a reversible Carnot cycle, heat transfer is directly proportional to absolute temperature.
- During isothermal expansion (1-2), heat transfer is $Q_{in} = T_H(s_2 - s_1)$.
- During isothermal compression (3-4), heat transfer is $Q_{out} = T_L(s_2 - s_1)$ (in magnitude).
- Substituting these yields the Carnot efficiency: $\eta_{Carnot} = 1 - \frac{T_L}{T_H}$.
- This shows that the efficiency of a Carnot cycle depends solely on the absolute temperatures of the source and the sink, not on the working fluid.

Carnot Principles & Practical Limits I

- First Carnot Principle: The efficiency of an irreversible heat engine is always less than the efficiency of a totally reversible one operating between the same two reservoirs.
- Second Carnot Principle: The efficiencies of all reversible heat engines operating between the same two reservoirs are the same.
- Isothermal heat transfer is practically impossible to achieve because it requires infinite time (infinitesimally slow piston movement) or infinite heat transfer area.
- Adiabatic processes require perfect thermal insulation and high-speed operation, which contradicts the slow speed required for isothermal processes.
- Due to these conflicting speed requirements, a real engine cannot execute the Carnot cycle.

Carnot Principles & Practical Limits II

- However, the Carnot cycle remains a vital theoretical limit that engineers strive to approach in design.

Summary & Key Takeaways I

- The Carnot cycle consists of four reversible processes: two isothermal and two isentropic.
- On a P-V diagram, the cycle is enclosed by two isotherms (flatter curves) and two adiabats (steeper curves).
- On a T-s diagram, the Carnot cycle is represented by a perfect rectangle, simplifying the analysis of heat and work.
- The net work output of the cycle is represented by the enclosed area on both the P-V and T-s diagrams.
- Carnot efficiency ($\eta = 1 - T_L/T_H$) represents the ultimate thermodynamic limit for converting heat into work.
- Increasing T_H or decreasing T_L maximizes the efficiency of power-generation systems.

Carnot Cycle and its Representation on P-V and T-s Diagram I

- Course: DI03019011 - Engineering Thermodynamics
- Unit 5: Thermodynamic Cycles
- Lecture 39
- Instructor: Professor of Mechanical Engineering

Agenda I

- Introduction to Sadi Carnot and the Ideal Reversible Cycle
- The Four Basic Processes of the Carnot Cycle
- Representation of Carnot Cycle on Pressure-Volume (P-V) Diagram
- Representation of Carnot Cycle on Temperature-Entropy (T-s) Diagram
- Thermodynamic Analysis & Thermal Efficiency Derivation
- The Two Carnot Principles (Corollaries of the Second Law)
- Practical Engineering Limitations of the Carnot Cycle

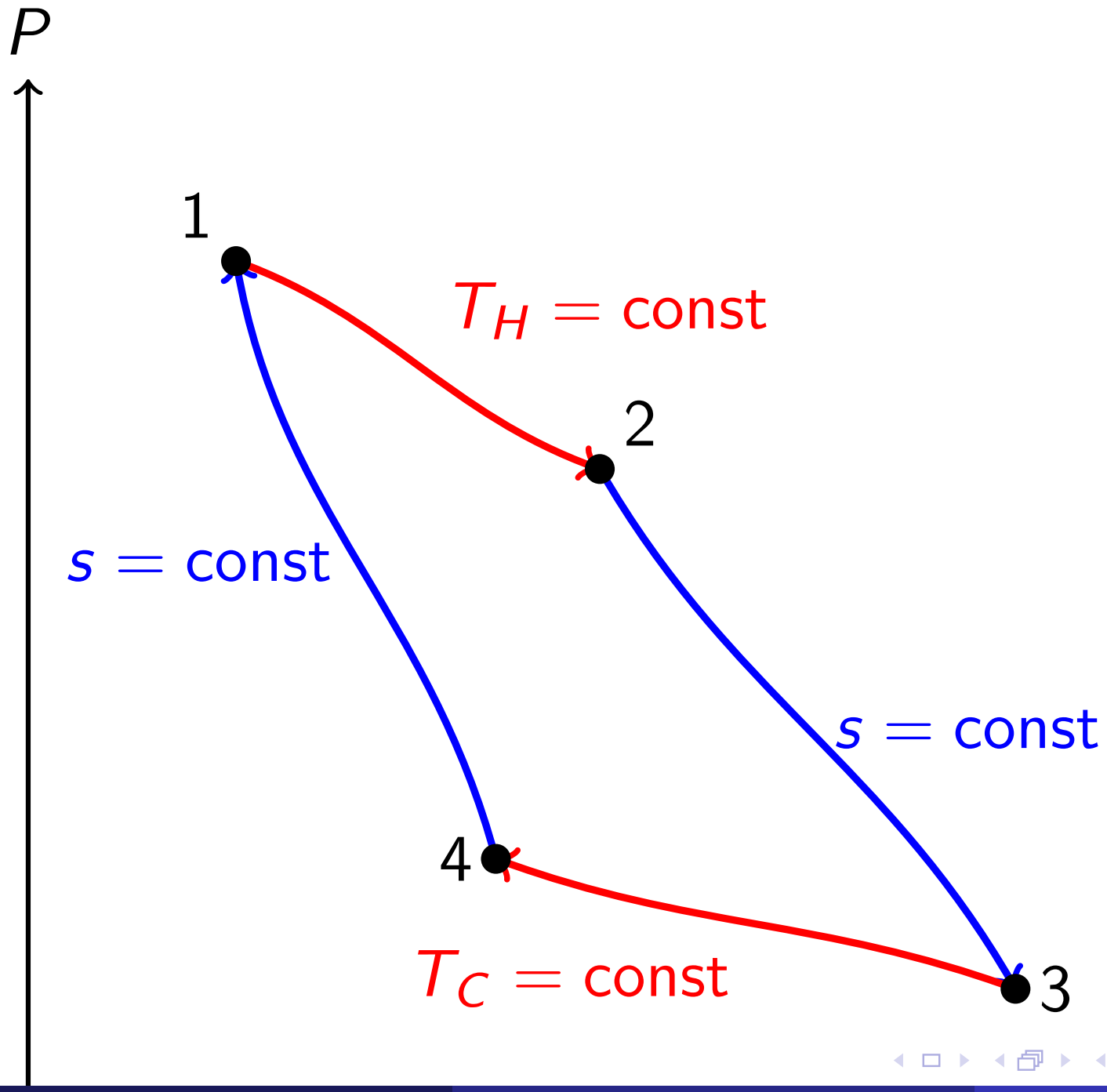
Introduction to the Carnot Cycle I

- Proposed in 1824 by French engineer Nicolas Léonard Sadi Carnot.
- It is a theoretical, idealized thermodynamic cycle that operates between two thermal reservoirs: a source at T_H and a sink at T_C.
- Consists entirely of reversible processes, making it both internally and externally reversible.
- Provides the maximum possible thermal efficiency that any heat engine can achieve operating between the same temperature limits.
- Serves as the absolute standard of comparison for all real-world power cycles.

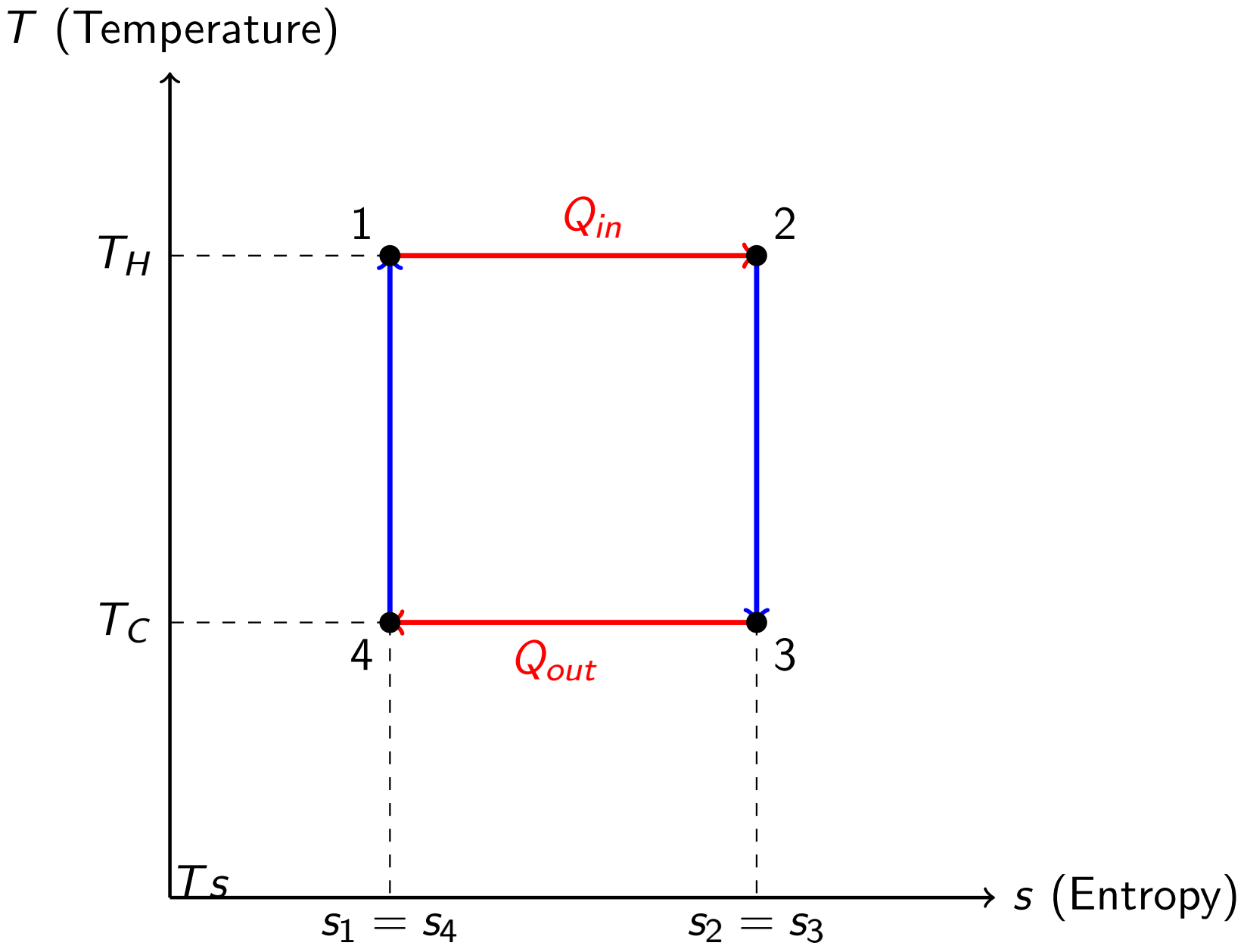
The Four Processes of the Carnot Cycle I

- Process 1-2: Reversible Isothermal Expansion. The system expands while maintaining constant high temperature T_H, absorbing heat Q_{in} from the source.
- Process 2-3: Reversible Adiabatic (Isentropic) Expansion. The working fluid continues to expand without heat exchange ($Q = 0$). The temperature drops from T_H to T_C as work is performed.
- Process 3-4: Reversible Isothermal Compression. The fluid is compressed while maintaining constant low temperature T_C, rejecting heat Q_{out} to the sink.
- Process 4-1: Reversible Adiabatic (Isentropic) Compression. The fluid is compressed without heat exchange ($Q = 0$), raising its temperature back from T_C to T_H to close the cycle.

Carnot Cycle on P-V Diagram



Carnot Cycle on T-s Diagram



Thermodynamic Analysis & Efficiency I

- Net work output is given by the area enclosed in the cycle: $W_{net} = Q_{in} - Q_{out}$.
- Heat addition during Process 1-2: $Q_{in} = T_H(s_2 - s_1)$.
- Heat rejection during Process 3-4: $Q_{out} = T_C(s_2 - s_1)$ (in magnitude).
- Thermal efficiency $\eta_{th} = \frac{W_{net}}{Q_{in}} = 1 - \frac{Q_{out}}{Q_{in}}$.
- Substituting the entropy expressions yields the famous relation: $\eta_{Carnot} = 1 - \frac{T_C}{T_H}$.
- Note: Temperature values must be in absolute scale (Kelvin or Rankine).

The Carnot Principles I

- Principle 1: The efficiency of an irreversible heat engine is always less than the efficiency of a totally reversible engine operating between the same two reservoirs.
- Principle 2: The efficiencies of all totally reversible heat engines operating between the same two reservoirs are identical.
- These principles are direct consequences of the Second Law of Thermodynamics.
- They prove that maximum possible efficiency is independent of the working fluid and depends solely on the reservoir temperatures.

Practical Limitations of the Carnot Cycle I

- Isothermal processes (1-2 and 3-4) require infinite time (extremely slow piston motion) to maintain thermal equilibrium.
- Adiabatic processes (2-3 and 4-1) require high-speed operations to prevent any heat transfer with the environment.
- Running an engine that is both infinitely slow and extremely fast in the same cycle is physically impossible.
- Isentropic expansion and compression require high pressure changes and massive cylinder sizes, leading to low work output per unit engine weight.
- Consequently, the Carnot cycle remains a purely theoretical limit and cannot be realized in practice.

Summary and Key Takeaways I

- The Carnot cycle establishes the upper efficiency bound for all thermodynamic engines.
- It consists of two reversible isothermal and two reversible isentropic processes.
- On the P-V diagram, isothermals are less steep than adiabatics due to the difference between $PV = \text{const}$ and $PV^{\gamma} = \text{const}$.
- On the T-s diagram, the Carnot cycle forms a simple rectangle, which simplifies thermodynamic evaluation.
- Maximum thermal efficiency depends exclusively on the absolute temperature limits: $\eta = 1 - T_C/T_H$.

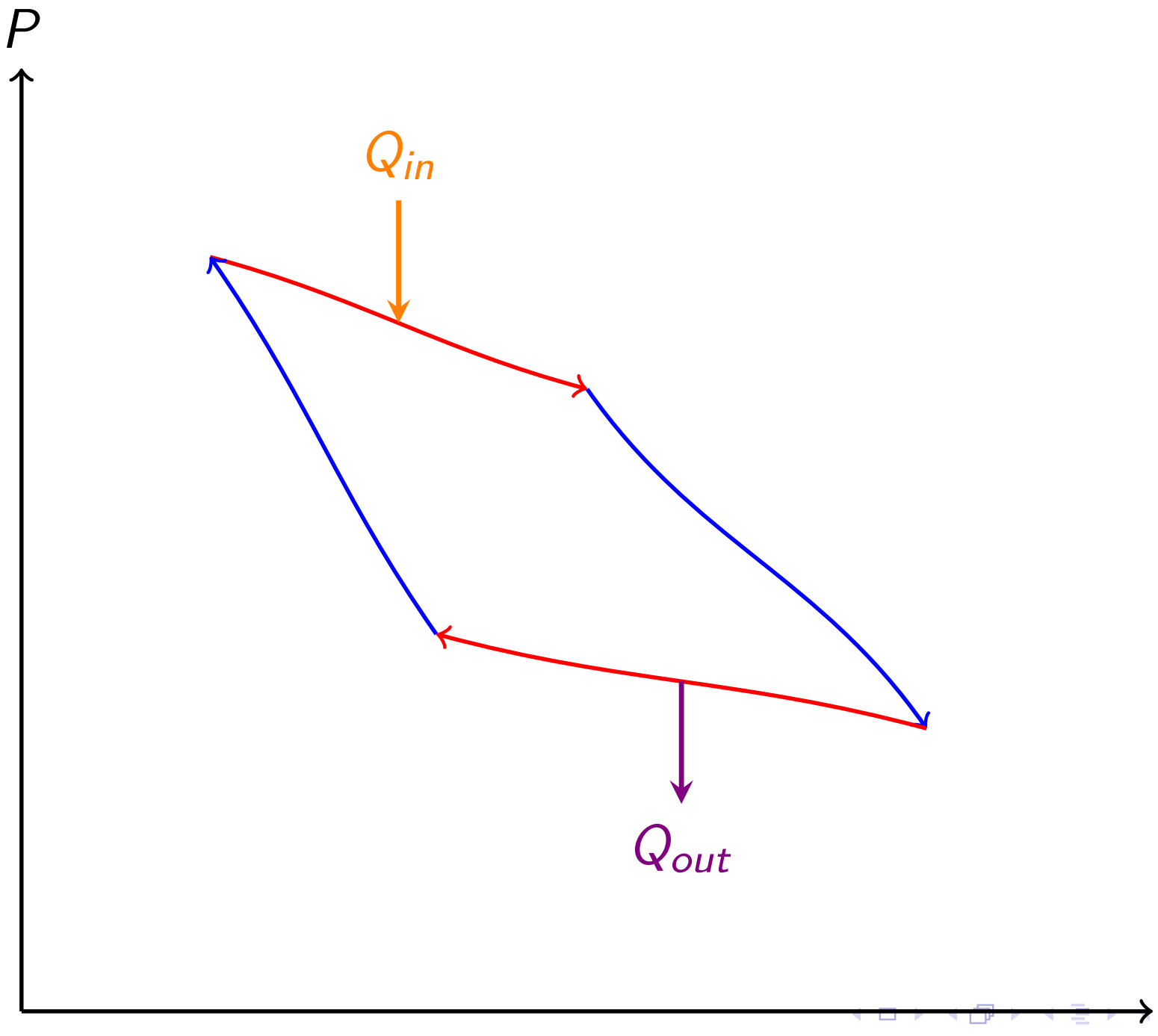
Derivation of thermal efficiency of Carnot cycle and simple numerical based on it. I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 5: Thermodynamic Cycles - Lecture 40
- Instructor: Professor of Mechanical Engineering
- Topic: Detailed derivation of the Carnot cycle's thermal efficiency and its practical application through numerical problems.

Lecture Agenda I

- Introduction to the Carnot Cycle and its four constituent processes.
- Thermodynamic analysis of each process (isothermal expansion, reversible adiabatic expansion, isothermal compression, reversible adiabatic compression).
- Net work output (W_{net}) and heat addition (Q_{in}) derivations.
- Derivation of the Carnot thermal efficiency (η_{Carnot}) in terms of temperatures T_H and T_C.
- Physical significance of Carnot efficiency and the Carnot Theorem.
- Step-by-step resolution of a simple numerical problem.
- Key takeaways and summary.

The Carnot Cycle on a P-V Diagram



Detailed Analysis of the Four Processes I

- Process 1-2: Reversible isothermal expansion of the gas at high temperature T_H. Heat absorbed is $Q_{in} = W_{1-2} = mRT_H \ln(V_2/V_1)$. Internal energy change is zero ($\Delta U_{1-2} = 0$).
- Process 2-3: Reversible adiabatic (isentropic) expansion. The system is thermally insulated. Temperature decreases from T_H to T_C. Work is done by the gas at the expense of its internal energy: $W_{2-3} = mC_v(T_H - T_C)$.
- Process 3-4: Reversible isothermal compression of the gas at low temperature T_C. Heat rejected to the sink is $Q_{out} = -W_{3-4} = mRT_C \ln(V_3/V_4)$. Internal energy change is zero ($\Delta U_{3-4} = 0$).

Detailed Analysis of the Four Processes II

- Process 4-1: Reversible adiabatic (isentropic) compression. The system is thermally insulated. Temperature increases from T_C to T_H. Work is done on the gas, increasing its internal energy: $W_{4-1} = -mC_v(T_H - T_C)$.

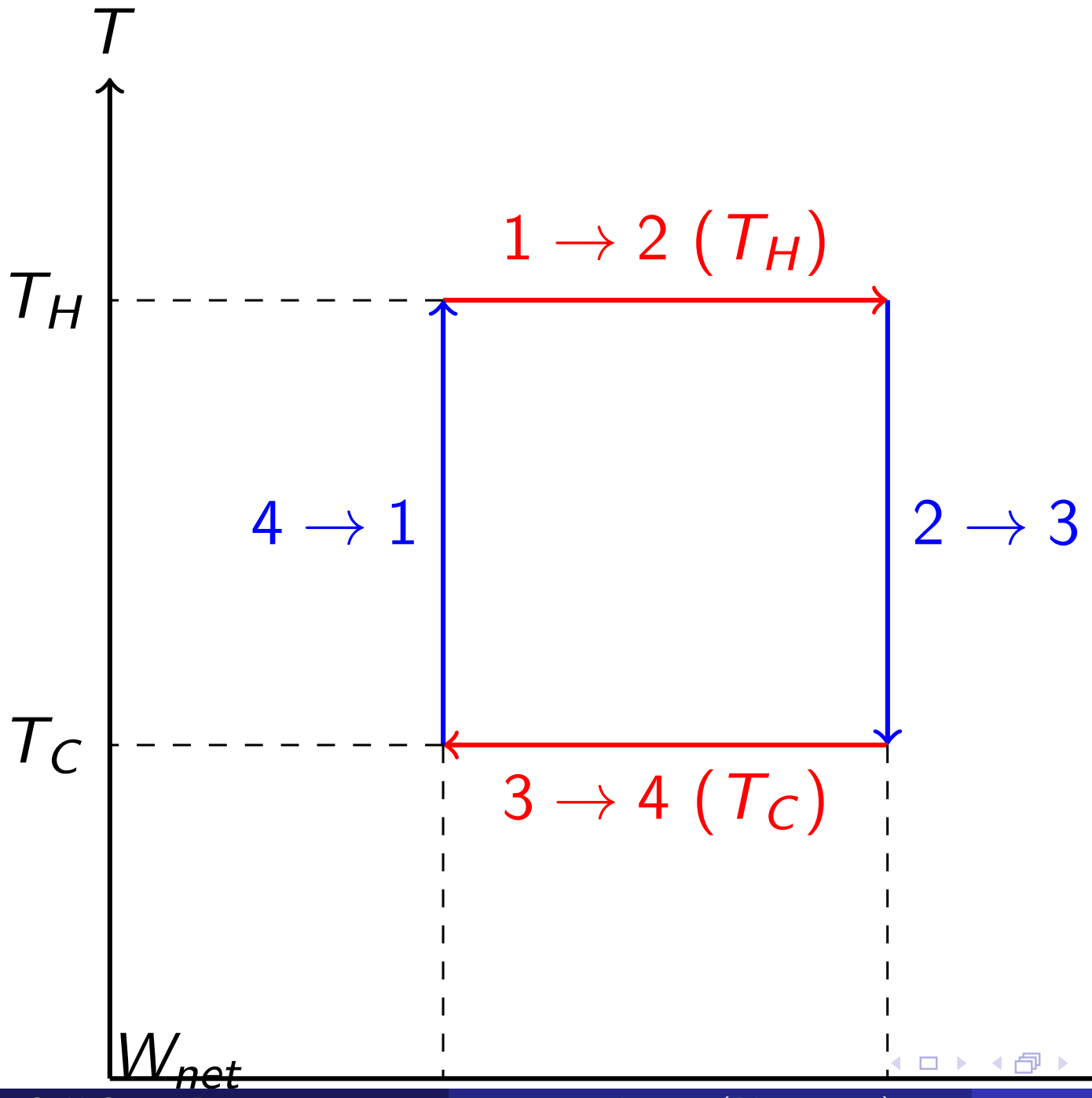
Isentropic Relations and Volume Ratios I

- To simplify the thermal efficiency expression, we must establish a relationship between the volume ratios of the isothermal processes.
- For the isentropic expansion process 2-3:
$$T_H V_2^{\gamma-1} = T_C V_3^{\gamma-1} \implies \frac{T_H}{T_C} = \left(\frac{V_3}{V_2}\right)^{\gamma-1}.$$
- For the isentropic compression process 4-1:
$$T_H V_1^{\gamma-1} = T_C V_4^{\gamma-1} \implies \frac{T_H}{T_C} = \left(\frac{V_4}{V_1}\right)^{\gamma-1}.$$
- Equating the temperature ratios:
$$\left(\frac{V_3}{V_2}\right)^{\gamma-1} = \left(\frac{V_4}{V_1}\right)^{\gamma-1} \implies \frac{V_3}{V_2} = \frac{V_4}{V_1}.$$
- Rearranging the volume ratios gives: $\frac{V_2}{V_1} = \frac{V_3}{V_4}$. This is a key simplification for the efficiency derivation.

Derivation of Thermal Efficiency (η) I

- Thermal efficiency of any heat engine is defined as: $\eta = \frac{W_{net}}{Q_{in}} = \frac{Q_{in} - Q_{out}}{Q_{in}} = 1 - \frac{Q_{out}}{Q_{in}}$.
- Substitute the expressions for heat transferred during the isothermal processes: $Q_{in} = mRT_H \ln\left(\frac{V_2}{V_1}\right)$ and $Q_{out} = mRT_C \ln\left(\frac{V_3}{V_4}\right)$.
- Substitute these into the efficiency formula: $\eta = 1 - \frac{mRT_C \ln(V_3/V_4)}{mRT_H \ln(V_2/V_1)}$.
- Since $\frac{V_2}{V_1} = \frac{V_3}{V_4}$, the logarithmic terms cancel out: $\ln\left(\frac{V_2}{V_1}\right) = \ln\left(\frac{V_3}{V_4}\right)$.
- Thus, the thermal efficiency of the Carnot cycle is: $\eta_{Carnot} = 1 - \frac{T_C}{T_H}$.
- This shows that Carnot efficiency depends solely on the absolute temperatures of the hot source (T_H) and cold sink (T_C).

The Carnot Cycle on a T-s Diagram



Carnot's Theorem and Key Observations I

- Carnot's Theorem: No heat engine operating between two heat reservoirs can be more efficient than a reversible (Carnot) engine operating between the same reservoirs.
- All reversible engines operating between the same two reservoirs have the same efficiency, regardless of the working fluid used.
- To maximize Carnot efficiency, one must either maximize the source temperature T_H (limited by material stress/metallurgical limits) or minimize the sink temperature T_C (limited by ambient environment temperature).
- If $T_C = 0$ K (Absolute Zero) or $T_H = \infty$, efficiency becomes 100% (1.0). Since neither condition is physically achievable (Third Law of Thermodynamics), a 100% efficient heat engine is impossible.

Carnot's Theorem and Key Observations II

- Real cycles (Otto, Diesel, Rankine) are less efficient than the Carnot cycle due to irreversibilities like friction, turbulence, and non-isothermal heat transfer.

Numerical Problem: Setup and Solution I

- Problem Statement: A Carnot heat engine operates between a source at 800°C and a sink at 30°C. If the engine receives 500 kW of heat from the source, calculate: (a) The thermal efficiency of the engine, and (b) The power output of the engine.
- Step 1: Convert temperatures to Kelvin (absolute scale). This is critical, as Carnot formula uses absolute temperature.
  $T_H = 800 + 273.15 = 1073.15$ K and
  $T_C = 30 + 273.15 = 303.15$ K or rounding to nearest integer:
  $T_H = 1073$ K, $T_C = 303$ K. Let's use the precise conversion.
- Step 2: Calculate thermal efficiency (η_{Carnot}).
  $\eta_{Carnot} = 1 - \frac{T_C}{T_H} = 1 - \frac{303.15}{1073.15}$.
- $\eta_{Carnot} = 1 - 0.2825 = 0.7175$ or 71.75%.

Numerical Problem: Setup and Solution II

- Step 3: Calculate the power output ($\dot{W}_{net}$). We know $\eta = \frac{\dot{W}_{net}}{\dot{Q}_{in}} \implies \dot{W}_{net} = \eta \times \dot{Q}_{in}$.
- $\dot{W}_{net} = 0.7175 \times 500 \text{ kW} = 358.75 \text{ kW}$.
- Therefore, the Carnot efficiency is 71.75% and the net power output is 358.75 kW.

Summary & Review Questions I

- The Carnot cycle represents the thermodynamic upper limit for the thermal efficiency of any heat engine operating between two constant-temperature reservoirs.
- The efficiency is purely a function of reservoir temperatures: $\eta_{Carnot} = 1 - T_C/T_H$, and is independent of the working substance (ideal gas, steam, etc.).
- Real-world engines are always less efficient than Carnot engines due to unavoidable irreversibilities.
- Review Question 1: Why is the Carnot cycle impossible to implement in practice? (Hint: Isothermal process requires infinitely slow speed; adiabatic requires infinitely fast speed).
- Review Question 2: A Carnot engine has an efficiency of 40% with a sink at 27°C. By how many Kelvin must the source temperature be increased to raise the efficiency to 50%?

Concept of air standard efficiency. I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 5: Thermodynamic Cycles
- Lecture 41: Concept of Air Standard Efficiency
- Instructor: Department of Mechanical Engineering

Agenda I

- Introduction to Gas Power Cycles & Reciprocating Engines
- The Air Standard Assumptions & Cold-Air-Standard Assumptions
- Definition and Thermodynamic Meaning of Air Standard Efficiency
- Carnot Cycle on P-V Coordinates (Diagram & Analysis)
- Derivation of Carnot Efficiency
- Otto Cycle on P-V Coordinates (Diagram & Analysis)
- Derivation of Otto Cycle Air Standard Efficiency
- Parametric Effects: Compression Ratio & Specific Heat Ratio
- Mean Effective Pressure (MEP) & Practical Application Limits

Air Standard Assumptions I

- To simplify the highly complex transient, chemical, and fluid dynamic processes in internal combustion engines, idealized cycles are analyzed.
- Air Standard Assumption 1: The working fluid is a fixed mass of air that continuously behaves as an ideal gas throughout the entire closed loop.
- Air Standard Assumption 2: All processes in the cycle are internally reversible (no friction, turbulence, or unrestrained expansions).
- Air Standard Assumption 3: The actual combustion process is modeled as an equivalent heat-addition process from an external high-temperature source.
- Air Standard Assumption 4: The exhaust stroke is replaced by an equivalent heat-rejection process to an external sink, returning the air to its initial state.

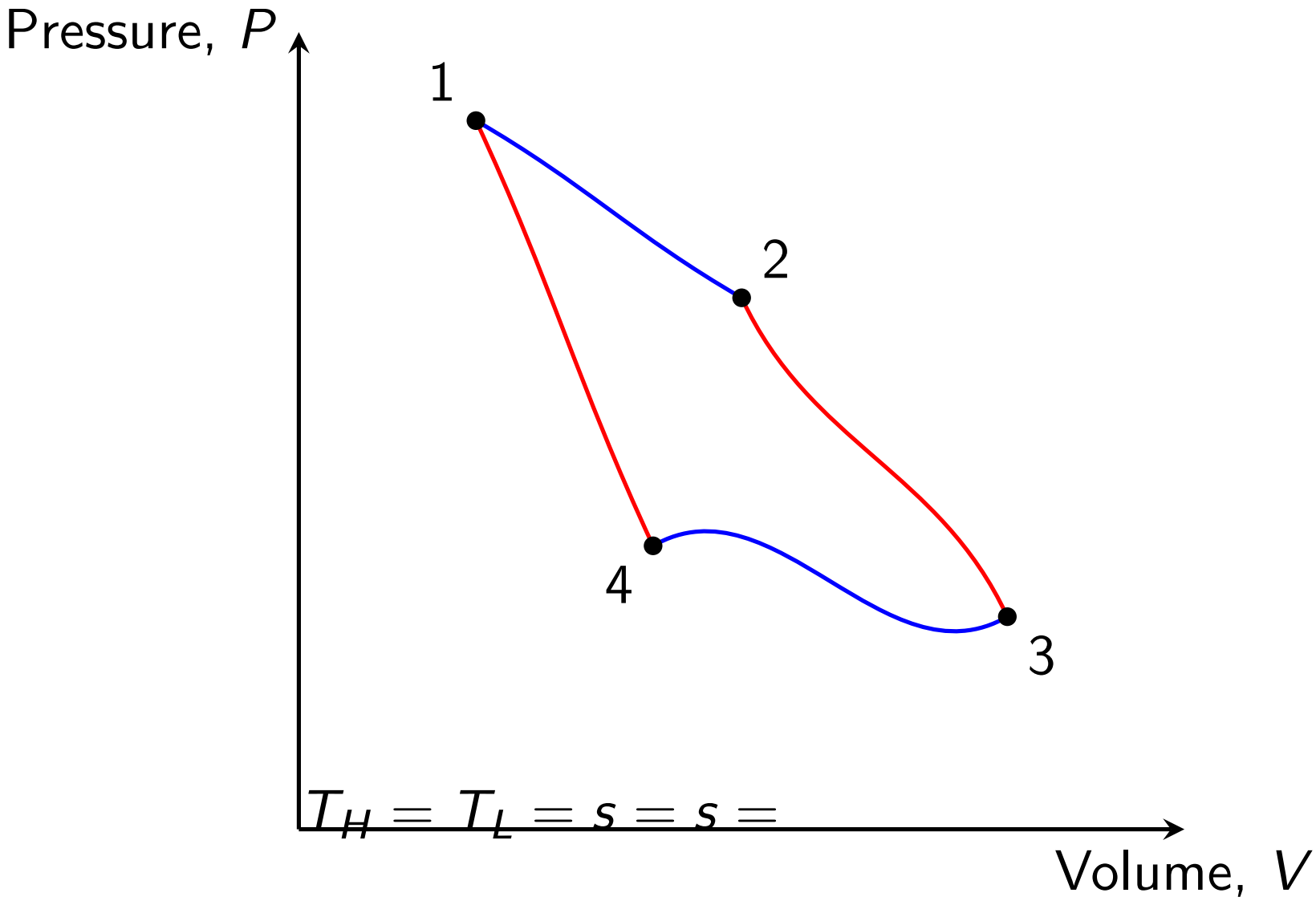
Air Standard Assumptions II

- Cold Air Standard Assumption: Air has constant specific heats ($C_p = 1.005$

Thermodynamic Definition of Air Standard Efficiency I

- Air Standard Efficiency ($\eta_{th,air}$) is the thermal efficiency of an idealized cycle operating under the air-standard assumptions.
- It measures how effectively a thermodynamic cycle converts heat input into net useful work output.
- Mathematical definition: $\eta_{th,air} = \frac{W_{net}}{Q_{in}} = \frac{Q_{in}-Q_{out}}{Q_{in}} = 1 - \frac{Q_{out}}{Q_{in}}$
- Here, W_{net} is the net power output, Q_{in} is the heat supplied from the external source, and Q_{out} is the heat rejected to the environment.
- This efficiency serves as the ultimate theoretical benchmark. Actual engine thermal efficiencies are always lower due to irreversibilities like friction and heat loss.

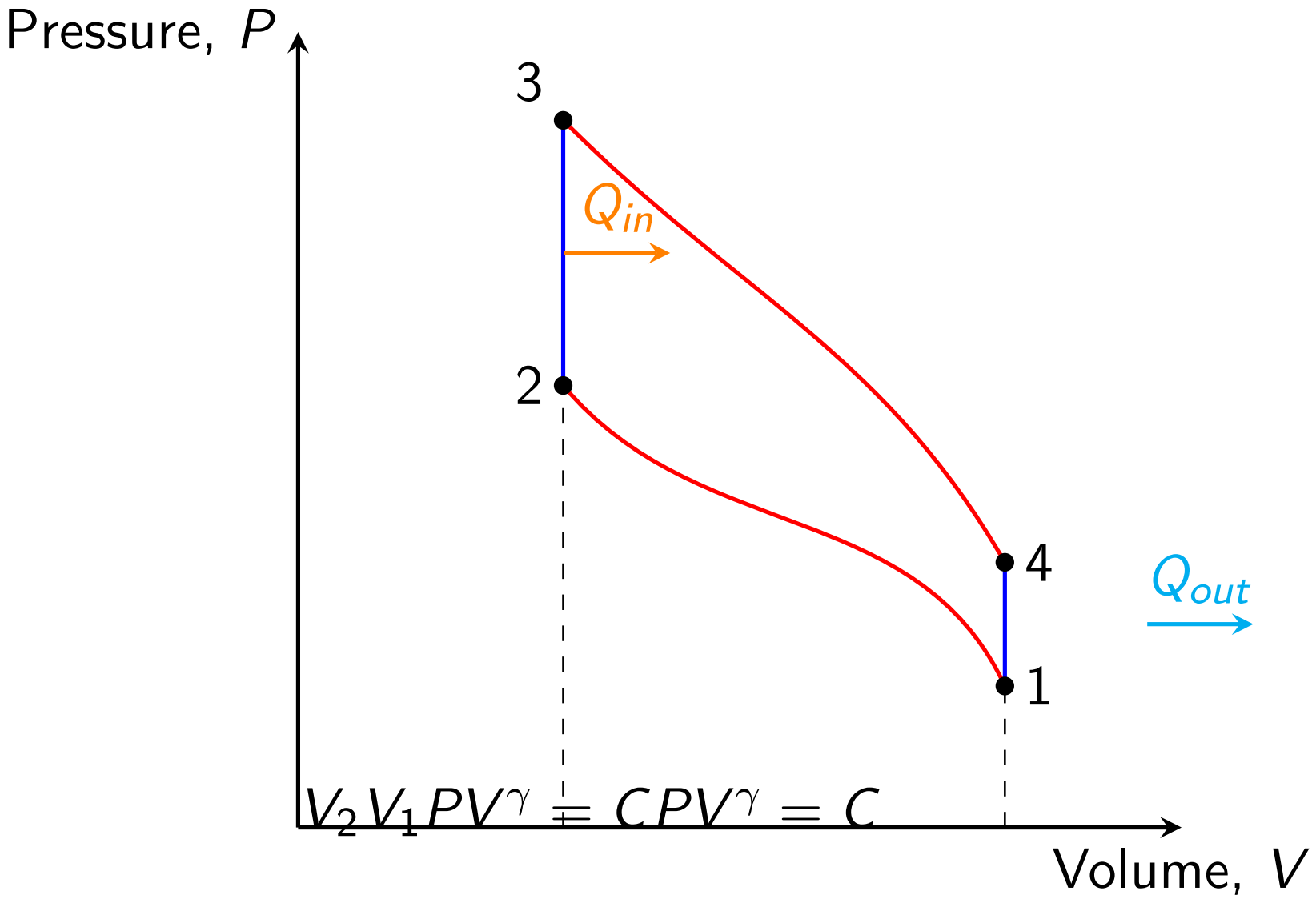
P-V Diagram of Carnot Cycle



Efficiency Analysis of Carnot Cycle I

- Heat added (Q_{in}) during isothermal expansion (1-2):
  $Q_{in} = mRT_H \ln\left(\frac{V_2}{V_1}\right)$
- Heat rejected (Q_{out}) during isothermal compression (3-4):
  $Q_{out} = mRT_L \ln\left(\frac{V_3}{V_4}\right)$
- For isentropic processes 2-3 and 4-1: $\frac{T_H}{T_L} = \left(\frac{V_3}{V_2}\right)^{\gamma-1}$ and $\frac{T_H}{T_L} = \left(\frac{V_4}{V_1}\right)^{\gamma-1}$
- Equating temperature ratios yields the volume relation:
  $\frac{V_3}{V_2} = \frac{V_4}{V_1} \implies \frac{V_2}{V_1} = \frac{V_3}{V_4}$
- Carnot efficiency simplifies strictly to:
  $\eta_{th,Carnot} = 1 - \frac{Q_{out}}{Q_{in}} = 1 - \frac{T_L}{T_H}$
- This establishes that Carnot efficiency is independent of the working fluid and depends only on absolute temperatures.

P-V Diagram of Air Standard Otto Cycle



Derivation of Otto Cycle Efficiency I

- Under constant volume processes: Heat input is $Q_{in} = mc_v(T_3 - T_2)$ and heat rejected is $Q_{out} = mc_v(T_4 - T_1)$.
- Thermal efficiency: $\eta_{th,Otto} = 1 - \frac{Q_{out}}{Q_{in}} = 1 - \frac{mc_v(T_4 - T_1)}{mc_v(T_3 - T_2)} = 1 - \frac{T_4 - T_1}{T_3 - T_2}$
- Using isentropic relations for 1-2 and 3-4: $\frac{T_2}{T_1} = \left(\frac{V_1}{V_2}\right)^{\gamma-1} = r^{\gamma-1}$ and $\frac{T_3}{T_4} = \left(\frac{V_4}{V_3}\right)^{\gamma-1} = r^{\gamma-1}$.
- Since $V_4 = V_1$ and $V_3 = V_2$, we find $\frac{T_2}{T_1} = \frac{T_3}{T_4} \implies \frac{T_4}{T_1} = \frac{T_3}{T_2}$.
- Factoring out temperatures: $\eta_{th,Otto} = 1 - \frac{T_1(T_4/T_1 - 1)}{T_2(T_3/T_2 - 1)} = 1 - \frac{T_1}{T_2}$.
- Final efficiency formulation: $\eta_{th,Otto} = 1 - \frac{1}{r^{\gamma-1}}$, where $r = V_1/V_2$ is the compression ratio.

Parameters Governing Air Standard Efficiency I

- Compression Ratio (r): As r increases, the thermal efficiency increases monotonically because the fuel-air charge is compressed to a higher pressure and temperature before ignition.
- Due to metallurgical limits and gasoline engine knock (pre-ignition), actual compression ratios for spark-ignition engines are restricted to $8 \leq r \leq 11$.
- Specific Heat Ratio (γ): Monatomic gases (e.g., helium, $\gamma = 1.667$) yield higher efficiencies than diatomic air ($\gamma = 1.4$) or triatomic combustion products.
- Mean Effective Pressure (MEP): $\text{MEP} = \frac{W_{net}}{V_1 - V_2}$. It represents the fictitious constant pressure that delivers the same net work output.

Parameters Governing Air Standard Efficiency II

- An engine with a larger MEP produces more net work per cycle for the same displacement volume, allowing for more compact design.

Summary of Engineering Takeaways I

- Air standard efficiency establishes the theoretical thermodynamic limits of reciprocating engine design.
- The Otto cycle serves as the standard for spark-ignition (petrol) engines, while the Diesel cycle serves compression-ignition systems.
- Maximizing the compression ratio (r) and utilizing high γ working fluids are key design drivers for efficiency.
- Engineers must balance high theoretical efficiencies against mechanical stresses, thermal loads, and knock phenomena.
- Real engine cycles operate at 50% to 70% of air standard efficiency due to friction, heat transfer, and finite combustion speed.

Concept of air standard efficiency. I

- Course: Engineering Thermodynamics
- Unit: 5 - Thermodynamic Cycles
- Lecture 42: Concept of Air Standard Efficiency
- Instructor: Professor of Mechanical Engineering

Agenda I

- Introduction to Gas Power Cycles & Air Standard Assumptions
- Definition of Air Standard Efficiency
- The Otto Cycle PV Diagram Analysis
- The Diesel Cycle PV Diagram Analysis
- Derivation of Air Standard Efficiency for Otto Cycle
- Derivation of Air Standard Efficiency for Diesel Cycle
- Comparison of Otto & Diesel Cycles
- Mean Effective Pressure (MEP) & Performance Metrics

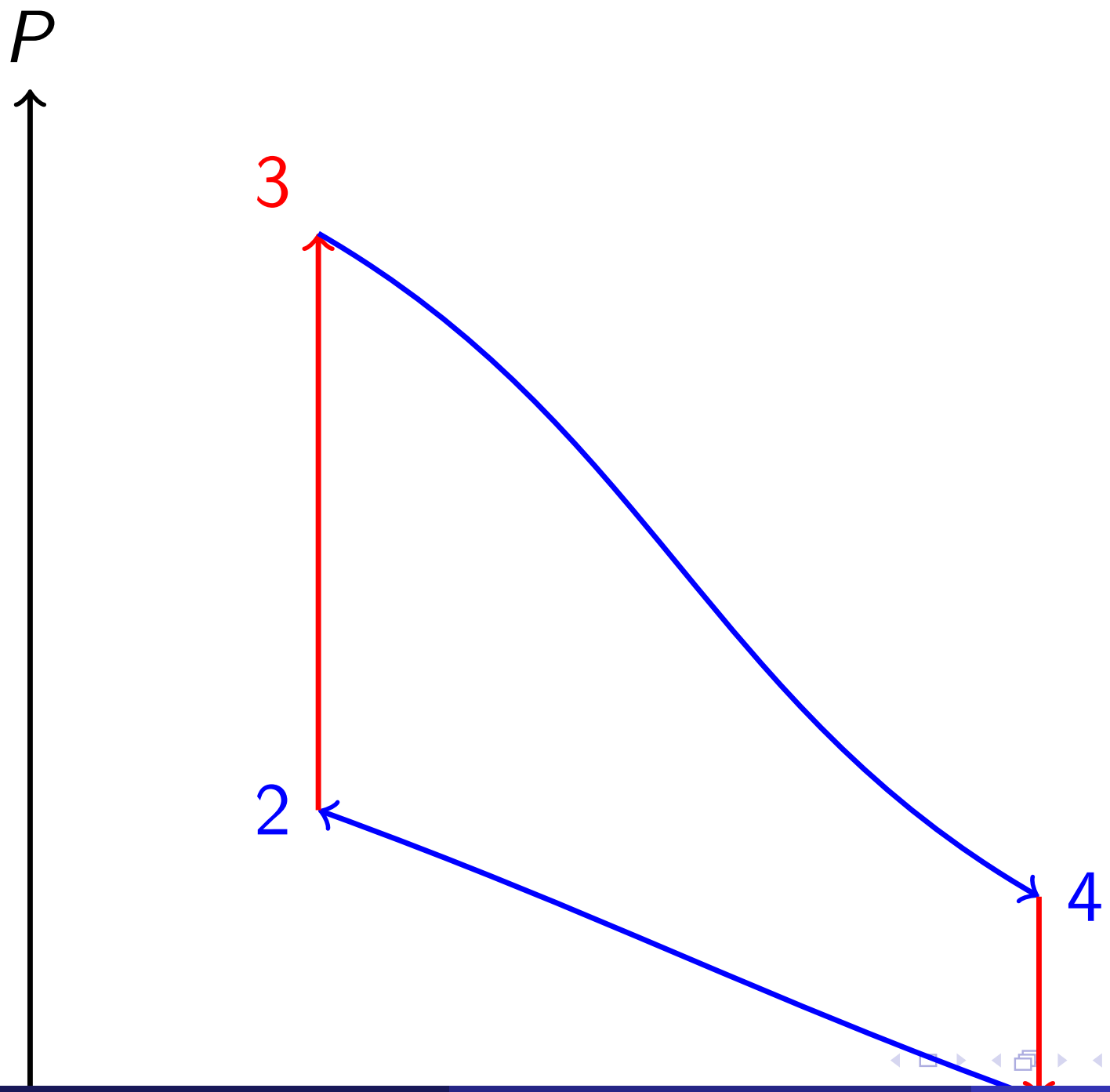
Air Standard Assumptions I

- Working fluid is a fixed mass of air that behaves as an ideal gas throughout the cycle.
- All processes are internally reversible.
- Combustion process is replaced by an external heat-addition process from a high-temperature source.
- Exhaust process is replaced by an external heat-rejection process to a low-temperature sink.
- Cold-air-standard assumption: Air has constant specific heats evaluated at room temperature ($C_p = 1.005$ kJ/kg $\cdot$ K, $C_v = 0.718$ kJ/kg $\cdot$ K, $\gamma = 1.4$).

Definition of Air Standard Efficiency I

- Air standard efficiency (η_{th}) is the thermal efficiency of a cycle operating under air standard assumptions.
- Defined as the ratio of net work output to the total heat supplied: $\eta_{th} = \frac{W_{net}}{Q_{in}}$.
- Using the First Law of Thermodynamics for a cycle: $W_{net} = Q_{in} - Q_{out}$.
- Therefore, the efficiency is expressed as: $\eta_{th} = 1 - \frac{Q_{out}}{Q_{in}}$.
- It represents the upper thermodynamic limit for internal combustion engines operating under these simplified conditions.

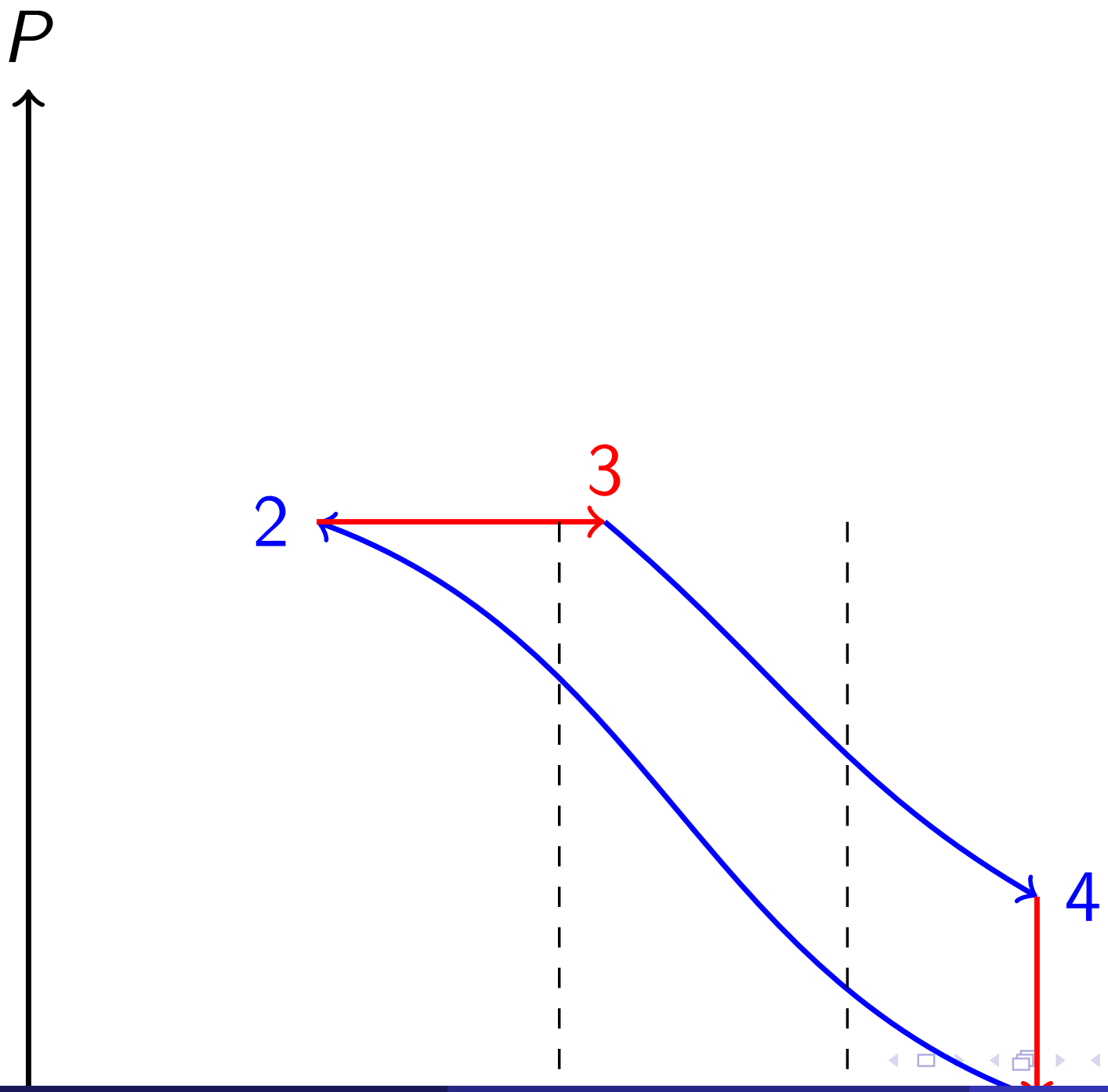
The Air-Standard Otto Cycle: P-V Diagram



Otto Cycle Efficiency Derivation I

- Heat addition at constant volume (2 to 3):
  $Q_{\text{in}} = mC_v(T_3 - T_2)$
- Heat rejection at constant volume (4 to 1):
  $Q_{\text{out}} = mC_v(T_4 - T_1)$
- Substitute into efficiency formula:
  $\eta_{\text{otto}} = 1 - \frac{Q_{\text{out}}}{Q_{\text{in}}} = 1 - \frac{T_4 - T_1}{T_3 - T_2} = 1 - \frac{T_1(T_4/T_1 - 1)}{T_2(T_3/T_2 - 1)}$
- For isentropic processes 1-2 and 3-4, with compression ratio $r = V_1/V_2$: $\frac{T_2}{T_1} = \left(\frac{V_1}{V_2}\right)^{\gamma-1} = r^{\gamma-1}$ and $\frac{T_3}{T_4} = \left(\frac{V_4}{V_3}\right)^{\gamma-1} = r^{\gamma-1}$
- Since $V_4 = V_1$ and $V_3 = V_2$, we get $\frac{T_2}{T_1} = \frac{T_3}{T_4} \implies \frac{T_4}{T_1} = \frac{T_3}{T_2}$
- Thus, the term $(T_4/T_1 - 1)$ cancels with $(T_3/T_2 - 1)$, leaving the final expression: $\eta_{\text{otto}} = 1 - \frac{1}{r^{\gamma-1}}$

The Air-Standard Diesel Cycle: P-V Diagram



Diesel Cycle Efficiency Derivation I

- Heat addition at constant pressure (2 to 3):
  $Q_{\text{in}} = mC_p(T_3 - T_2)$
- Heat rejection at constant volume (4 to 1):
  $Q_{\text{out}} = mC_v(T_4 - T_1)$
- Substitute into efficiency formula:
  $\eta_{\text{diesel}} = 1 - \frac{Q_{\text{out}}}{Q_{\text{in}}} = 1 - \frac{C_v(T_4 - T_1)}{C_p(T_3 - T_2)} = 1 - \frac{1}{\gamma}\frac{T_4 - T_1}{T_3 - T_2}$
- Using compression ratio $r = V_1/V_2$ and cut-off ratio $r_c = V_3/V_2$.
- Expressing temperatures in terms of T_1, r, r_c, and γ yields the final efficiency equation: $\eta_{\text{diesel}} = 1 - \frac{1}{r^{\gamma-1}}\left[\frac{r_c^{\gamma} - 1}{\gamma(r_c - 1)}\right]$
- Since the bracketed term is always greater than 1 for $r_c > 1$, the Diesel cycle is less efficient than the Otto cycle for the same compression ratio.

Factors Affecting Air Standard Efficiency I

- Compression Ratio (r): As r increases, efficiency of both Otto and Diesel cycles increases due to higher expansion temperatures.
- Specific Heat Ratio (γ): Monatomic gases (higher $\gamma \approx 1.67$) yield higher efficiency than diatomic gases like air ($\gamma \approx 1.4$).
- Cut-off Ratio (r_c): For a Diesel cycle, increasing the cut-off ratio decreases the thermal efficiency because heat addition occurs at a later stage of expansion.
- Practical Limits: High compression ratios lead to engine knock (auto-ignition) in gasoline engines, limiting r to about 10-12, whereas diesel engines can operate at r of 15-22.

Mean Effective Pressure (MEP) I

- Mean Effective Pressure (MEP) is a theoretical constant pressure that, if acting on the piston during the power stroke, would produce the same net work.
- Mathematically defined as: $\text{MEP} = \frac{W_{\text{net}}}{V_{\text{swept}}} = \frac{W_{\text{net}}}{V_1 - V_2}$.
- It acts as a parameter to compare the power output performance of engines of different sizes.
- A higher MEP indicates that the engine produces more net work output per unit displacement volume.
- For a given displacement, maximizing MEP is key to optimizing engine performance and power density.

Otto, Diesel, Dual, and Brayton Cycles: Diagrams, Efficiency Equations & Solved Examples I

- Course Code: DI03019011
- Course: Engineering Thermodynamics
- Unit 5: Thermodynamic Power Cycles
- Lecture 43: Comparative Study of Air-Standard Cycles

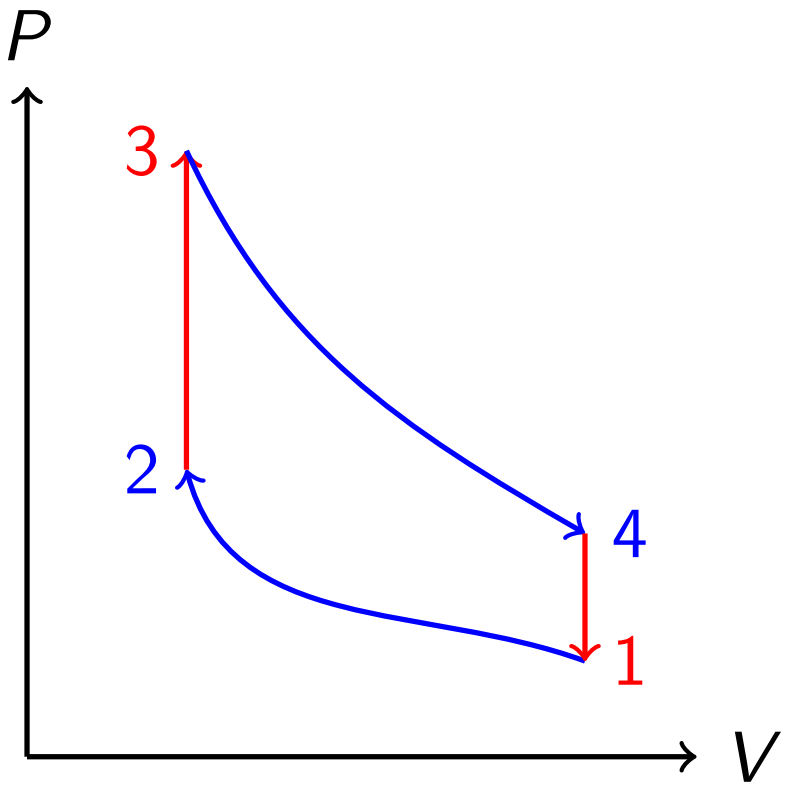
Lecture Agenda I

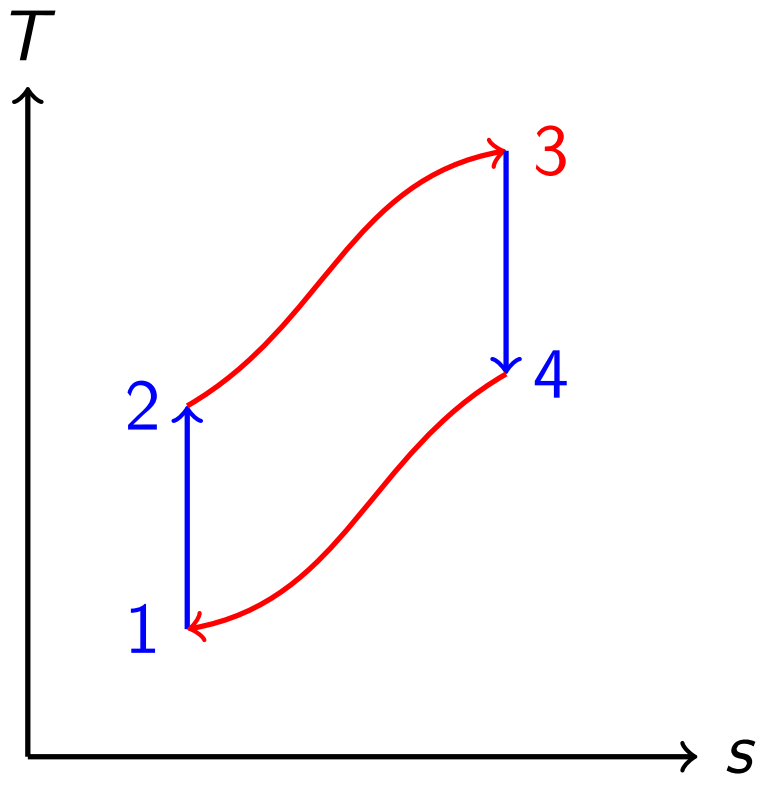
- Fundamental Air-Standard Assumptions and Idealizations
- Otto Cycle: P-V & T-s Representations, Thermal Efficiency Equation
- Diesel Cycle: P-V & T-s Representations, Thermal Efficiency Equation
- Dual Cycle (Limited Pressure Cycle): P-V & T-s Representations, Thermal Efficiency Equation
- Brayton Cycle: Gas Turbine Operations, P-V & T-s Diagrams, Thermal Efficiency Equation
- Thermodynamic Comparison of Otto, Diesel, and Dual Cycles
- Practical Numerical Examples & Problem Solving

Air-Standard Assumptions & Idealizations I

- The working fluid is a fixed mass of air that behaves as an ideal gas throughout the entire cycle.
- All combustion processes are replaced by equivalent heat-addition processes from an external source.
- The exhaust process is replaced by a heat-rejection process to an external sink, restoring the air to its initial thermodynamic state.
- All processes are assumed to be internally reversible, meaning friction and other irreversibilities are neglected.
- Cold Air-Standard Assumptions: Specific heats of air ($C_p = 1.005$
- The specific heat ratio (adiabatic index) for air is assumed constant: $\gamma = C_p/C_v = 1.4$.

Otto Cycle: P-V and T-s Diagrams

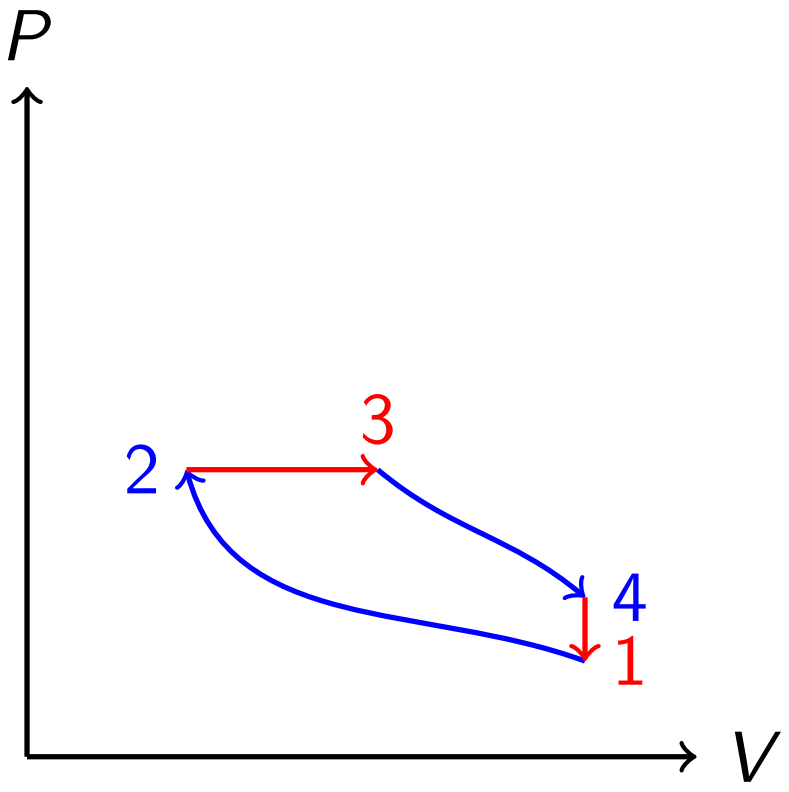


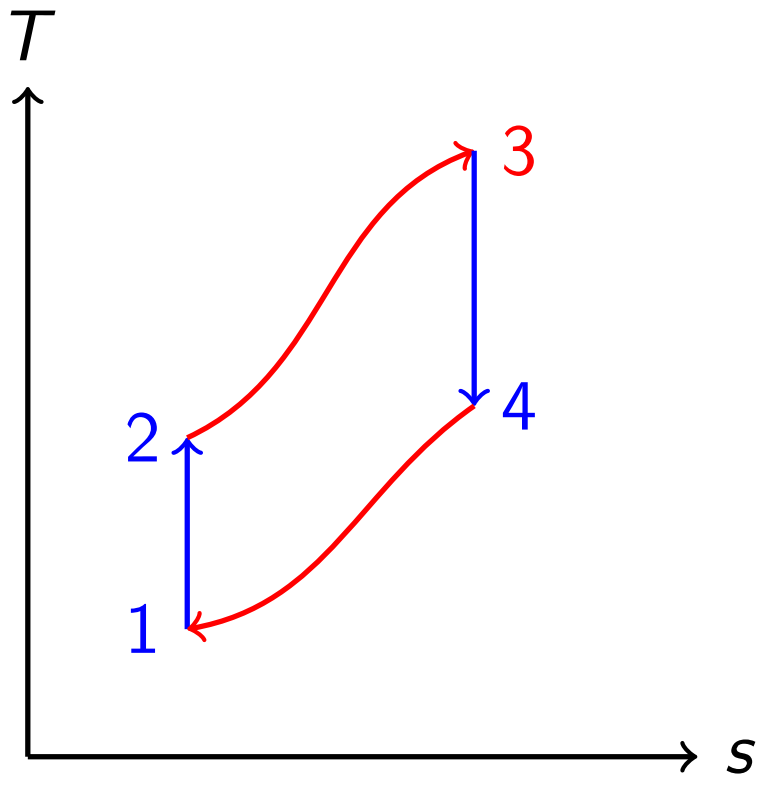


Otto Cycle: Efficiency Formulation I

- The Otto cycle is the ideal thermodynamic cycle for spark-ignition (SI) internal combustion engines.
- Compression ratio (r): Defined as the ratio of maximum to minimum cylinder volume, $r = V_1/V_2 = V_4/V_3$.
- Process 2-3 (Constant-volume heat addition): Combustion is simulated by heat input $Q_{in} = mC_v(T_3 - T_2)$.
- Process 4-1 (Constant-volume heat rejection): Heat is rejected to the atmosphere as $Q_{out} = mC_v(T_4 - T_1)$.
- Thermal efficiency: $\eta_{th,Otto} = 1 - \frac{Q_{out}}{Q_{in}} = 1 - \frac{T_4 - T_1}{T_3 - T_2}$.
- Air-Standard Efficiency equation (without derivation): $\eta_{th,Otto} = 1 - \frac{1}{r^{\gamma-1}}$.
- Key insight: Efficiency increases asymptotically with compression ratio r and is limited by engine knock.

Diesel Cycle: P-V and T-s Diagrams





Diesel and Dual Cycles: Efficiency Formulation I

- The Diesel cycle is the ideal cycle for compression-ignition (CI) engines, where fuel is injected into hot compressed air.
- Compression ratio: $r = V_1/V_2$. Cut-off ratio (combustion duration): $r_c = V_3/V_2$.
- Heat addition is at constant pressure: $Q_{in} = mC_p(T_3 - T_2)$. Heat rejection is at constant volume: $Q_{out} = mC_v(T_4 - T_1)$.
- Air-Standard Efficiency of Diesel Cycle: $\eta_{th,Diesel} = 1 - \frac{1}{r^{\gamma-1}}\left[\frac{r_c^{\gamma}-1}{\gamma(r_c-1)}\right]$.
- The Dual Cycle (Sabathe Cycle) represents modern CI engines where heat is added in two stages: partly at constant volume and partly at constant pressure.
- Pressure ratio: $\alpha = P_3/P_2$. Cut-off ratio: $r_c = V_4/V_3$.
- Air-Standard Efficiency of Dual Cycle: $\eta_{th,Dual} = 1 - \frac{1}{r^{\gamma-1}}\left[\frac{\alpha r_c^{\gamma}-1}{(\alpha-1)+\gamma\alpha(r_c-1)}\right]$.

Brayton Cycle: Gas Turbine Operations I

- The Brayton cycle is the ideal thermodynamic cycle for simple gas turbine power plants and aircraft jet engines.
- Process 1-2: Isentropic compression in a compressor, raising the pressure from P_1 to P_2 ($s_1 = s_2$).
- Process 2-3: Constant-pressure heat addition ($P_2 = P_3$) via combustion, where $Q_{in} = mC_p(T_3 - T_2)$.
- Process 3-4: Isentropic expansion in a turbine ($s_3 = s_4$), generating mechanical work output.
- Process 4-1: Constant-pressure heat rejection ($P_4 = P_1$), cooling the fluid via a heat exchanger or exhaust release.
- Pressure ratio (r_p): Defined as $r_p = P_2/P_1 = P_3/P_4$.
- Air-Standard Efficiency of Brayton Cycle: $\eta_{th,Brayton} = 1 - \frac{1}{(r_p)^{(\gamma-1)/\gamma}}$.

Thermodynamic Comparisons of the Cycles I

- Comparison at the same compression ratio (r) and heat input: $\eta_{th,Otto} > \eta_{th,Dual} > \eta_{th,Diesel}$. Constant-volume heat addition allows maximum peak pressure.
- Comparison at the same maximum pressure (P_{max}) and temperature (T_{max}): $\eta_{th,Diesel} > \eta_{th,Dual} > \eta_{th,Otto}$. Under this condition, the Diesel cycle operates at a much higher compression ratio.
- Brayton cycle efficiency behaves identically to the Otto cycle when expressed in terms of volumetric compression ratio: $\eta_{th,Brayton} = 1 - 1/r^{\gamma-1}$.
- Dual cycle efficiency simplifies to Diesel cycle efficiency when $\alpha = 1$ (no constant-volume heat addition stage).
- Dual cycle efficiency simplifies to Otto cycle efficiency when $r_c = 1$ (no constant-pressure heat addition stage).

Simple Numerical Examples I

- Example 1 (Otto): An engine runs on an ideal Otto cycle with a compression ratio of $r = 8.5$. Using cold air-standard values ($\gamma = 1.4$), find its thermal efficiency. Solution: $\eta_{th} = 1 - 1/(8.5^{0.4}) = 1 - 1/2.355 \approx 57.5\%$.
- Example 2 (Diesel): A Diesel engine has a compression ratio of $r = 18$. Combustion heat is added until the volume doubles ($r_c = 2$). Find the thermal efficiency. Solution: $\eta_{th} = 1 - \frac{1}{18^{0.4}} \left[\frac{2^{1.4}-1}{1.4(2-1)} \right] \approx 1 - 0.315 \cdot [1.17] \approx 63.2\%$.
- Example 3 (Brayton): A gas turbine power plant operates on the Brayton cycle with a pressure ratio of $r_p = 8$. Find the efficiency. Solution: $\eta_{th} = 1 - 1/(8^{(1.4-1)/1.4}) = 1 - 1/(8^{0.286}) \approx 1 - 1/1.811 \approx 44.8\%$.
- Practical Summary: Thermodynamic cycles specify upper bounds. Real engine cycles have lower efficiencies due to heat loss, friction, and fluid dissociation.

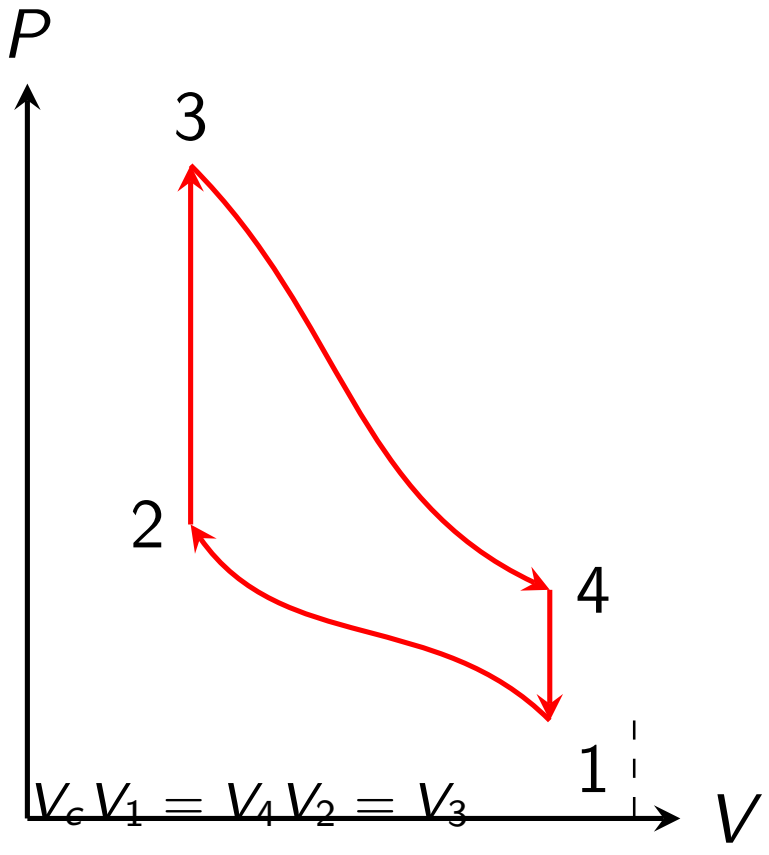
Otto, Diesel, Dual, and Brayton Cycles I

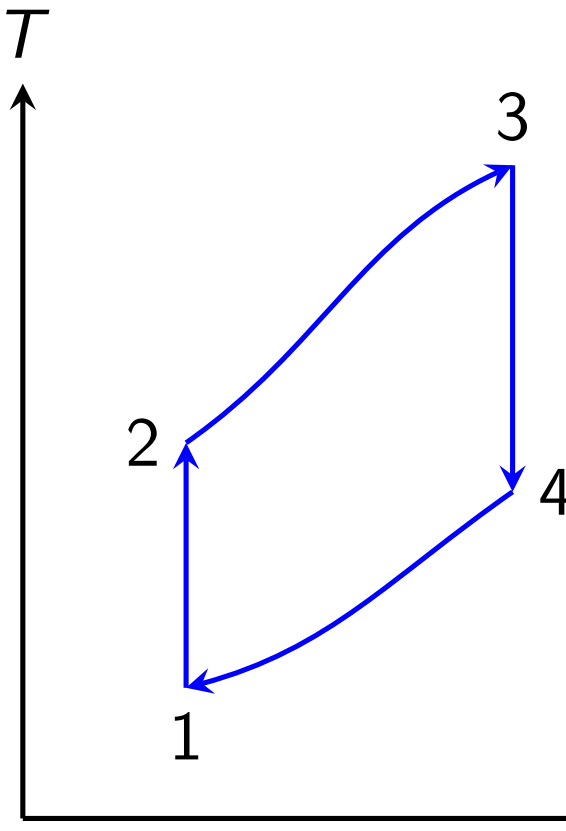
- Course Code: DI03019011
- Course Title: Engineering Thermodynamics
- Lecture 44: Otto, Diesel, dual and Brayton: Representation on P-V & T-s diagram, Equation of air standard efficiency (Without derivations) and simple examples.
- Unit 5: Gas Power Cycles

Agenda I

- Air-Standard Assumptions: Simplified modeling parameters for gas power cycles.
- The Otto Cycle: Four-stroke spark-ignition engine ideal cycle, diagrams, and efficiency equation.
- The Diesel Cycle: Four-stroke compression-ignition engine ideal cycle, diagrams, and efficiency equation.
- The Dual Combustion Cycle: Representation of modern high-speed engines (Sabathé cycle).
- The Brayton Cycle: Basic gas turbine engine cycle under closed/open configurations.
- Summary of key thermodynamic parameters (Compression ratio r, cut-off ratio rc, pressure ratio rp).
- Practical numerical examples on air-standard efficiency calculations.

Otto Cycle: Representation on P-V and T-s Diagrams

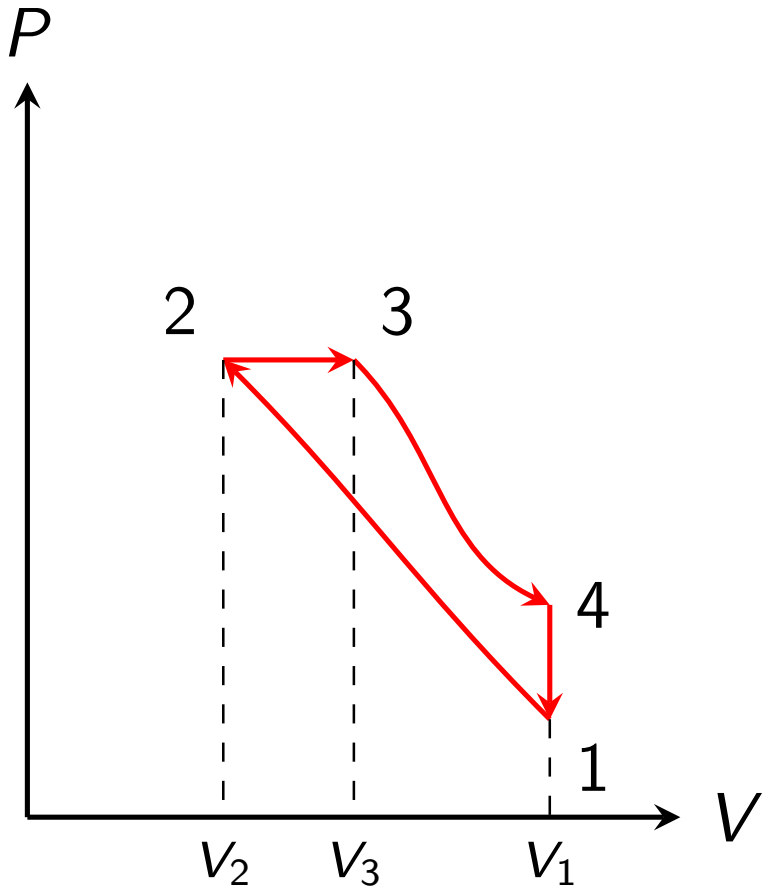


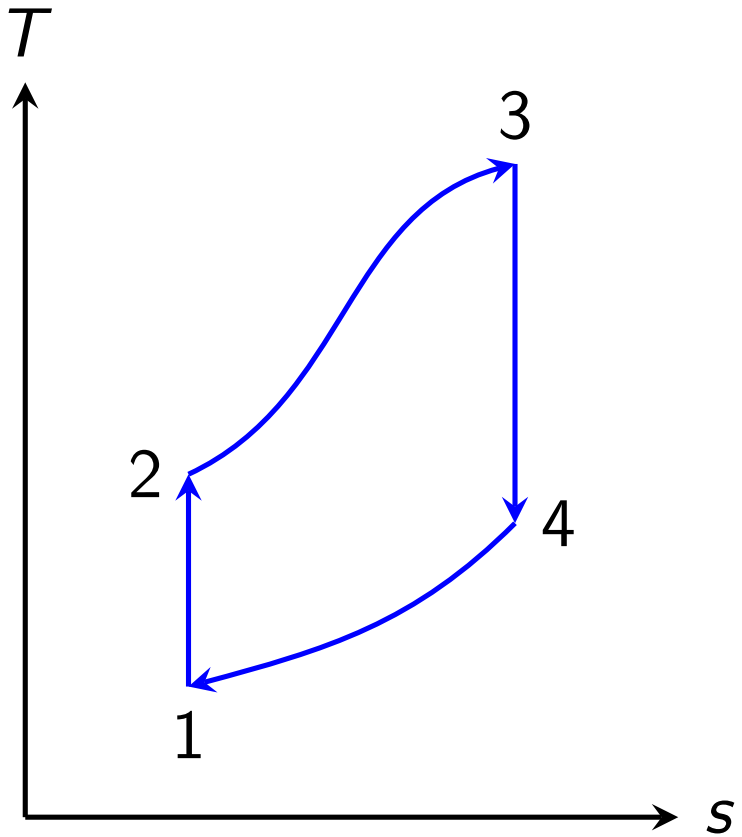


Otto Cycle: Equation of Air Standard Efficiency I

- Compression ratio is defined as: $r = V_1/V_2 = V_4/V_3$, which represents the ratio of maximum to minimum cylinder volume.
- Air-Standard Efficiency of the Otto Cycle: $\eta_{th,Otto} = 1 - \frac{1}{r^{\gamma-1}}$.
- Here, γ represents the specific heat ratio (C_p/C_v), taken as approximately 1.4 for ambient air under cold air-standard conditions.
- The thermal efficiency is independent of heat added or temperature limits; it depends solely on compression ratio (r) and specific heat ratio (γ).
- As the compression ratio r increases, the thermal efficiency increases, but is practically limited to $8 \leq r \leq 12$ due to fuel auto-ignition (knocking).
- Mean Effective Pressure (MEP): $MEP = W_{net}/(V_1 - V_2)$, indicating the theoretical constant pressure that would produce the same net work.

Diesel Cycle: Representation on P-V and T-s Diagrams

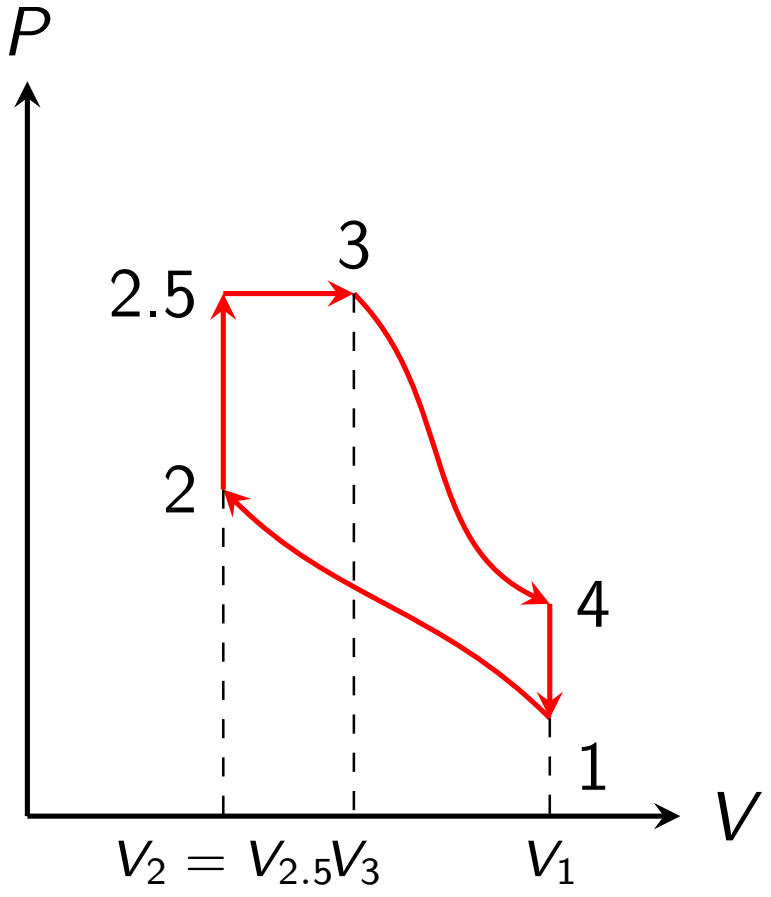


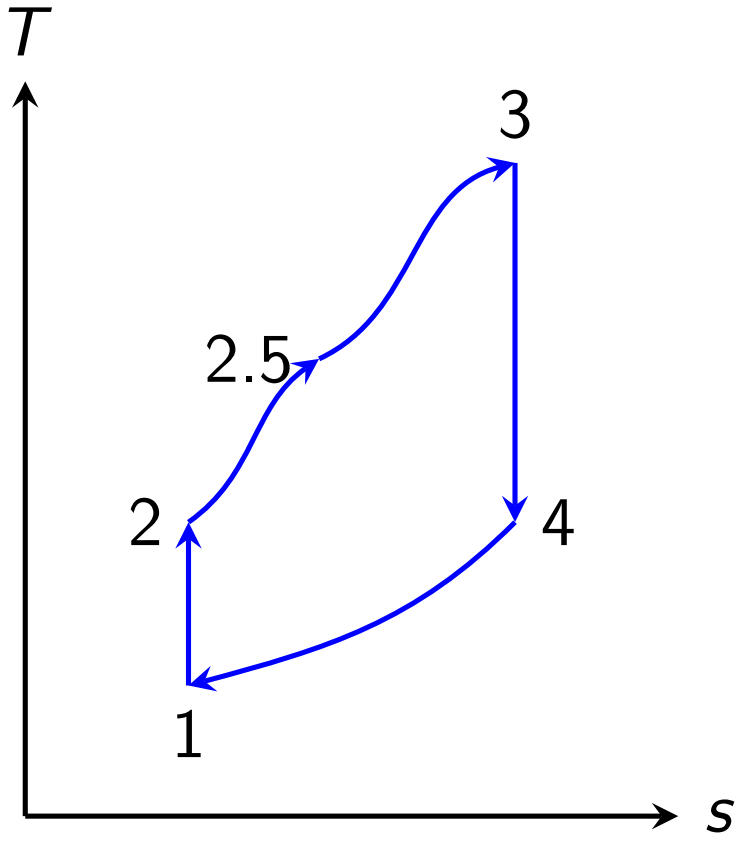


Diesel Cycle: Equation of Air Standard Efficiency I

- Compression ratio is defined as: $r = V_1/V_2$. Cut-off ratio is defined as: $r_c = V_3/V_2$, indicating volume ratio during combustion.
- Air-Standard Efficiency of the Diesel Cycle: $\eta_{th,Diesel} = 1 - \frac{1}{r^{\gamma-1}}\left[\frac{r_c^{\gamma}-1}{\gamma(r_c-1)}\right]$.
- Because $r_c > 1$, the bracketed factor is always greater than 1, meaning that for the same compression ratio r, the Diesel cycle is less efficient than the Otto cycle: $\eta_{Diesel} < \eta_{Otto}$.
- However, Diesel engines compress only air, allowing compression ratios from 15 to 22, resulting in higher operational efficiencies in practice.
- The constant-pressure heat addition simulates fuel injection occurring over a portion of the stroke in slower-speed engines.
- Expansion ratio during the power stroke: $r_e = V_4/V_3 = r/r_c$.

Dual Cycle: Representation on P-V and T-s Diagrams

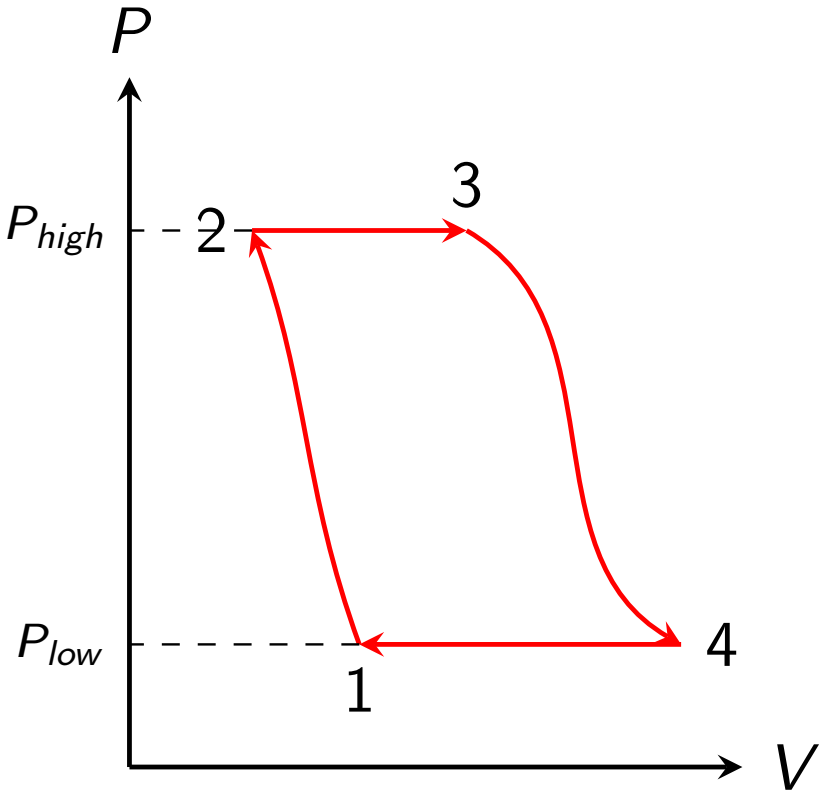


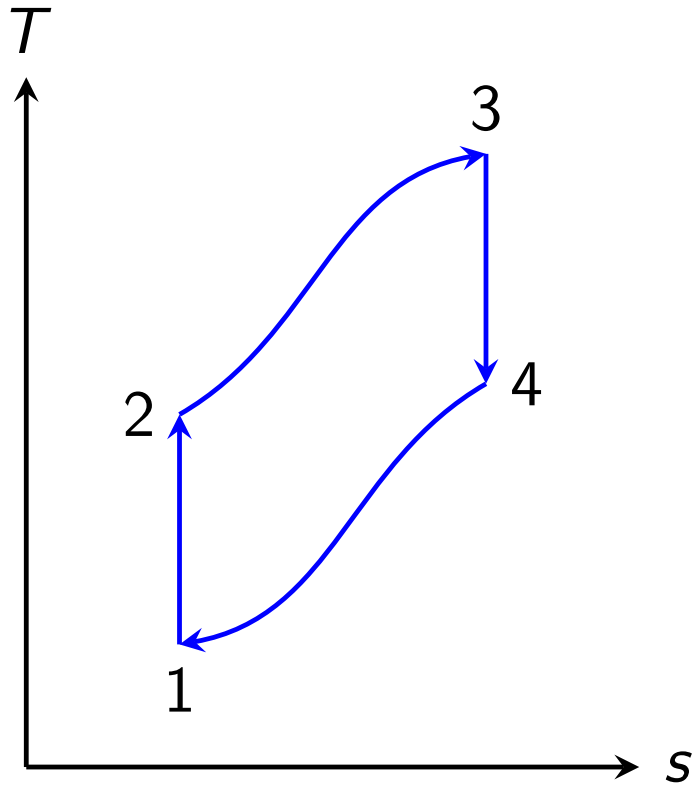


Dual Cycle: Equation of Air Standard Efficiency I

- Key parameter definitions: Compression ratio $r = V_1/V_2$, explosion (pressure) ratio $r_p = P_{2.5}/P_2$, and cut-off ratio $r_c = V_3/V_{2.5}$.
- Air-Standard Efficiency of the Dual Cycle: $\eta_{th,Dual} = 1 - \frac{1}{r^{\gamma-1}} \left[\frac{r_p r_c^{\gamma} - 1}{(r_p - 1) + \gamma r_p (r_c - 1)} \right]$.
- The Dual cycle represents a realistic combination, matching modern high-speed compression-ignition engine performance.
- Limiting behaviors: Setting $r_p = 1$ reduces this equation to the Diesel efficiency formula; setting $r_c = 1$ reduces it to the Otto efficiency formula.
- For identical compression ratio and heat input, thermodynamic cycle efficiency ranks as: $\eta_{Otto} > \eta_{Dual} > \eta_{Diesel}$.
- For identical maximum cycle pressure and heat input, efficiency ranks as: $\eta_{Diesel} > \eta_{Dual} > \eta_{Otto}$.

Brayton Cycle: Representation on P-V and T-s Diagrams





Brayton Cycle Efficiency & Simple Illustrative Examples I

- The pressure ratio is defined as: $r_p = P_2/P_1 = P_3/P_4$.
  Air-Standard Efficiency of Brayton Cycle:
  $\eta_{th,Brayton} = 1 - \frac{1}{r_p^{(\gamma-1)/\gamma}}$.
- Example 1: An Otto cycle engine operates with a compression ratio of 8. Estimate its thermal efficiency (take $\gamma = 1.4$).
  Solution: $\eta_{Otto} = 1 - 1/8^{0.4} = 1 - 0.435 = 0.565$ or 56.5%.
- Example 2: A Diesel engine has a compression ratio of 16 and cut-off occurs at 6% of the stroke. Find its thermal efficiency.
  Solution: Stroke volume $V_s = V_1 - V_2 = 15V_2$. Cut-off volume difference
  $V_3 - V_2 = 0.06V_s = 0.9V_2 \implies V_3 = 1.9V_2 \implies r_c = 1.9$.
  $\eta_{Diesel} = 1 - \frac{1}{16^{0.4}}\left[\frac{1.9^{1.4}-1}{1.4\times(1.9-1)}\right] \approx 1 - 0.330 \times 1.171 \approx 61.4\%$.

Brayton Cycle Efficiency & Simple Illustrative Examples II

- Example 3: A stationary gas turbine operating on the Brayton cycle has a pressure ratio of 8. Compute its air-standard efficiency. Solution:
  $\eta_{Brayton} = 1 - 1/8^{(1.4-1)/1.4} = 1 - 1/8^{0.2857} = 1 - 0.552 = 0.448$ or 44.8%.
- Example 4: Calculate the efficiency of a Dual combustion cycle with compression ratio $r = 15$, explosion ratio $r_p = 1.4$, and cut-off ratio $r_c = 1.2$. Solution:
  $\eta_{Dual} = 1 - \frac{1}{15^{0.4}} \left[\frac{1.4(1.2)^{1.4} - 1}{(1.4-1)+1.4\times 1.4(1.2-1)} \right] =$
  $1 - 0.338 \left[\frac{1.4(1.291)-1}{0.4+1.96(0.2)} \right] \approx 1 - 0.338 \left[\frac{0.807}{0.792} ight \right] \approx 65.5\%.$

Brayton Cycle Efficiency & Simple Illustrative Examples III

- Summary of parameters: Spark-ignition engines (Otto) are limited by fuel pre-ignition; compression-ignition engines (Diesel/Dual) avoid this limitation; gas turbines (Brayton) operate continuously with distinct pressure ratios.

Representation of Reversed Carnot cycle and Reversed Brayton cycle on P-V and T-s diagram respectively. I

- Course: Engineering Thermodynamics (DI03019011)
- Unit 5: Gas Power and Vapor Cycles
- Lecture 45: Representation of Reversed Carnot & Reversed Brayton Cycles
- Mechanical Engineering Department

Lecture Agenda I

- Introduction to Gas Refrigeration Systems
- Reversed Carnot Cycle: Thermodynamic Principles and Processes
- Reversed Carnot Cycle on P-V and T-s Diagrams
- COP of Reversed Carnot Cycle and Practical Limitations
- Reversed Brayton (Bell-Coleman) Cycle: Principles and Processes
- Reversed Brayton Cycle on P-V and T-s Diagrams
- Comparison of Carnot and Brayton Refrigeration Cycles

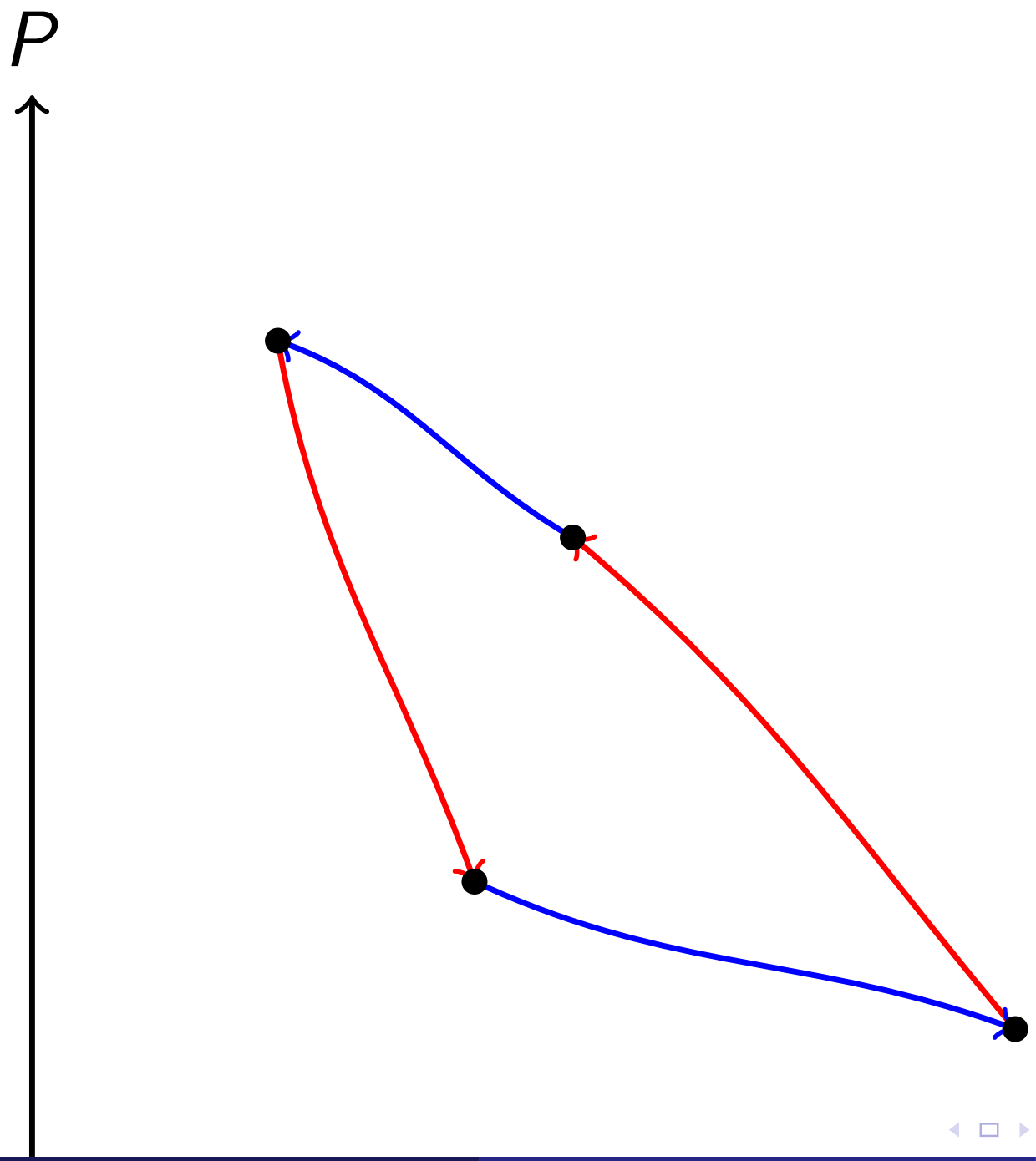
The Reversed Carnot Cycle: Operating Principles I

- The Carnot cycle is completely reversible, meaning all four processes can be reversed to serve as a heat pump or refrigerator.
- It defines the upper limit of efficiency (COP) for any refrigeration system operating between two temperature limits.
- Process 1-2 (Isentropic Compression): Working fluid is compressed reversibly and adiabatically from state 1 to state 2, raising its temperature from TL to TH.
- Process 2-3 (Isothermal Compression): Heat QH is rejected reversibly to the hot reservoir at constant temperature TH while the fluid is compressed.
- Process 3-4 (Isentropic Expansion): Working fluid expands reversibly and adiabatically, dropping its temperature from TH to TL.

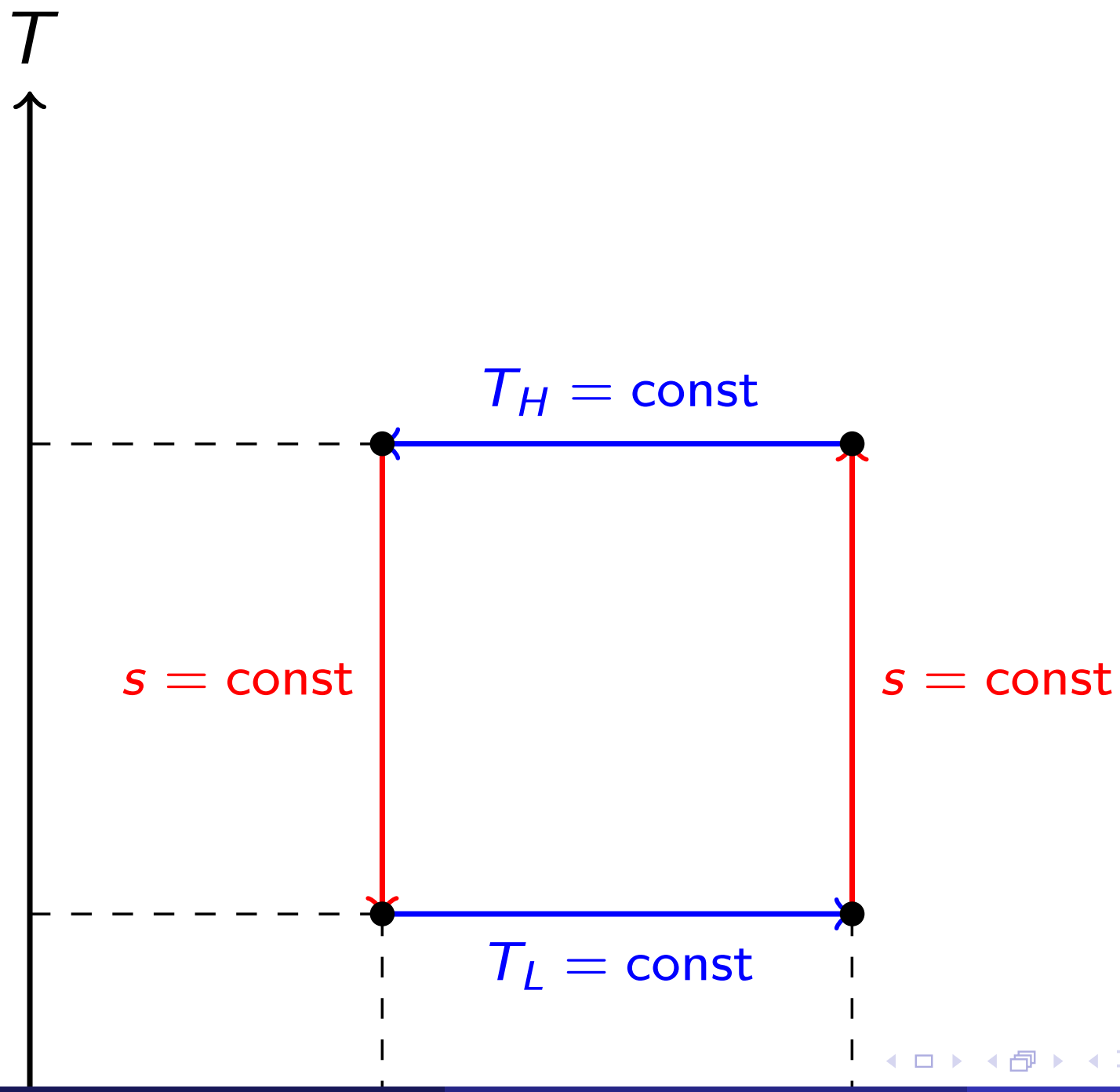
The Reversed Carnot Cycle: Operating Principles II

- Process 4-1 (Isothermal Expansion): Heat QL is absorbed reversibly from the cold space at constant temperature TL, completing the cycle.

Reversed Carnot Cycle on P-V Diagram



Reversed Carnot Cycle on T-s Diagram



Limitations of the Reversed Carnot Cycle I

- Isothermal heat transfer processes (2-3 and 4-1) require very low fluid velocities to allow heat exchange, meaning high cycle durations.
- Isentropic compression and expansion (1-2 and 3-4) require high velocities to achieve adiabatic conditions, presenting a mechanical contradiction.
- Wet compression issue: At state 1, the refrigerant is a mixture of liquid and vapor. Compressing a liquid-vapor mixture damages compressor parts.
- Low work output from expander: The work generated during isentropic expansion (3-4) is very low for liquids/vapors and doesn't justify the cost of an expander.
- As a result, practical systems use vapor-compression cycles with a simple expansion valve rather than an expander.

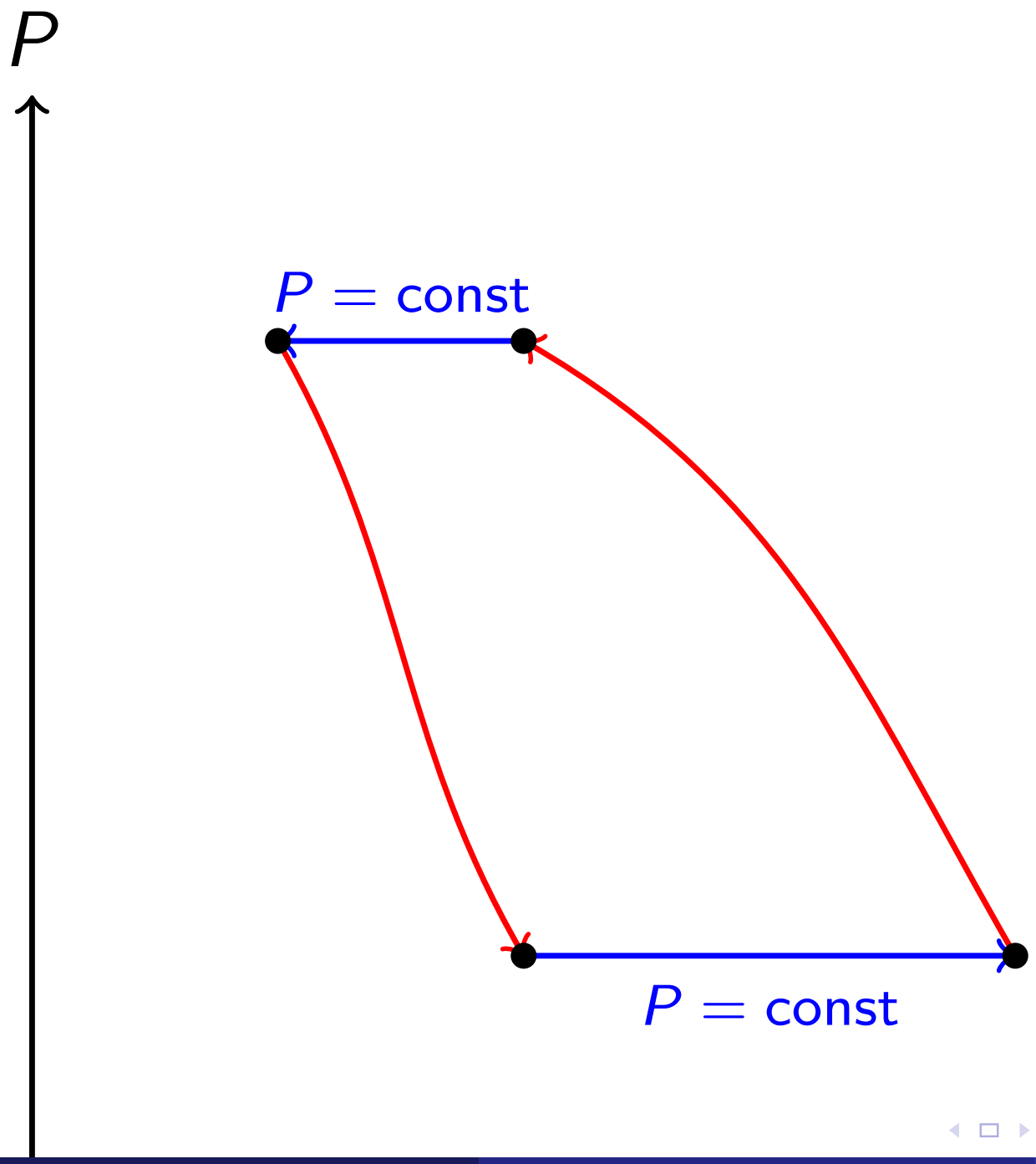
The Reversed Brayton Cycle: Introduction I

- The Reversed Brayton cycle (also known as the Bell-Coleman or Air Refrigeration cycle) is a practical cycle using a gas working fluid.
- Unlike the Carnot cycle, the working fluid (typically air) remains in the gaseous state throughout all processes.
- Process 1-2 (Isentropic Compression): Air is compressed in a compressor, raising its pressure from P1 to P2 and temperature from T1 to T2.
- Process 2-3 (Constant Pressure Cooling): Heat QH is rejected at constant high pressure P2, cooling the air from temperature T2 to T3.
- Process 3-4 (Isentropic Expansion): Air is expanded in an expander/turbine, dropping the pressure to P1 and temperature to a very low T4.

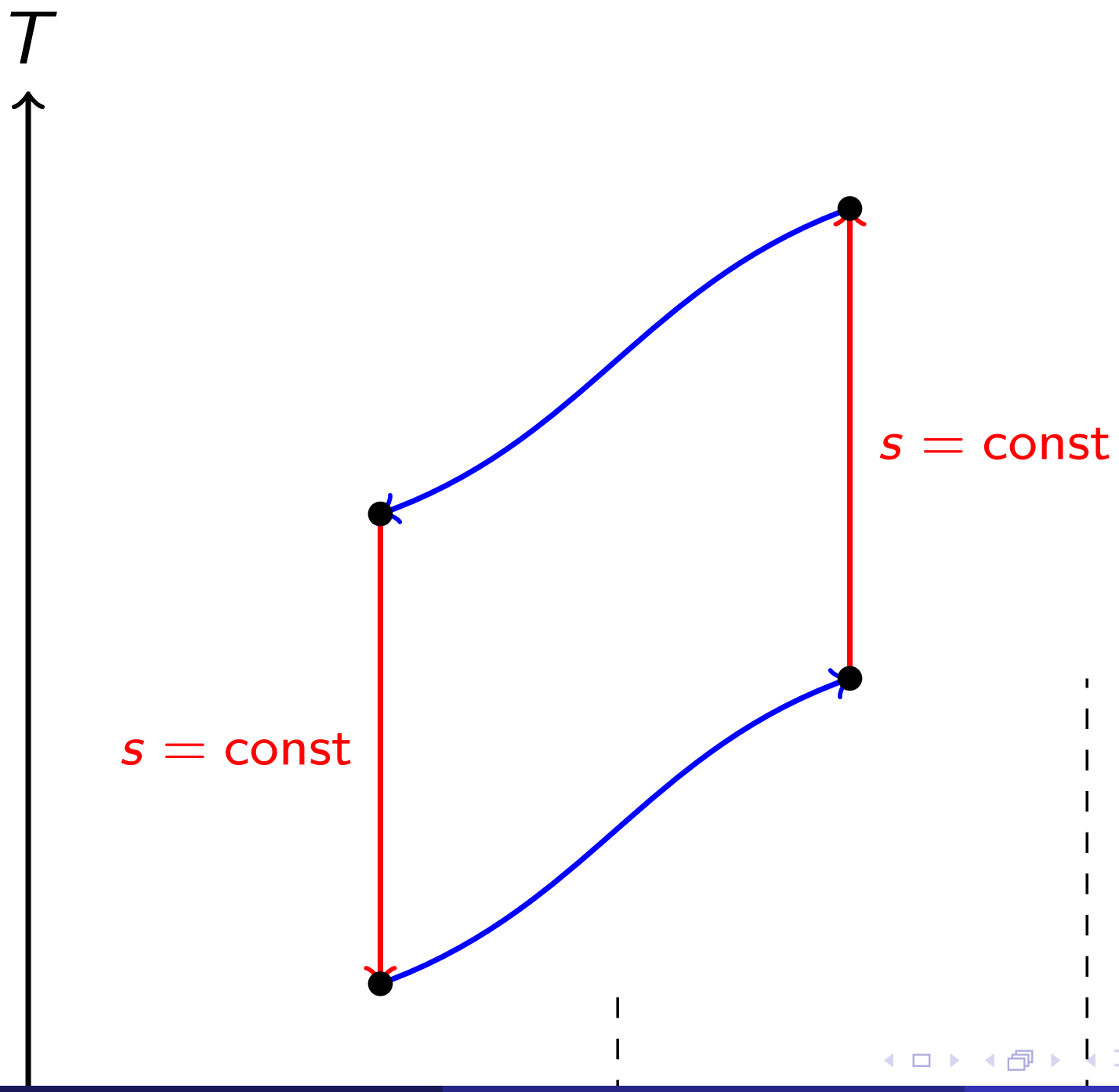
The Reversed Brayton Cycle: Introduction II

- Process 4-1 (Constant Pressure Heating): Air absorbs heat QL from the cold space at constant pressure P1, rising in temperature to T1.

Reversed Brayton Cycle on P-V Diagram



Reversed Brayton Cycle on T-s Diagram



Comparison of Reversed Carnot and Reversed Brayton Cycles I

- Efficiency: Under identical maximum and minimum temperature limits, the Carnot COP is always higher than the Brayton COP.
- Phase Change: Reversed Carnot involves two-phase flow (wet vapor), while Reversed Brayton uses single-phase gaseous working fluid.
- Heat Transfer Mode: Carnot uses isothermal heat transfer, whereas Brayton uses constant-pressure heat transfer, which is faster and easier to implement.
- System Complexity: Reversed Brayton cycle requires a gas compressor and an expander turbine, which makes it lightweight and ideal for aircraft cooling.

Comparison of Reversed Carnot and Reversed Brayton Cycles II

- Practical Application: Brayton cycles are viable for gas-cycle refrigeration, while Carnot serves only as a theoretical thermodynamic limit.