Thermodynamics Fundamentals – Unit 1

(BME301 Mechanical Engineering Notes)

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Contents

1 Introduction to Thermodynamics

Definition: Thermodynamics

Thermodynamics is the science of energy transformations, studying relationships between heat, work, and macroscopic properties like pressure (P), volume (V), and temperature (T). It predicts system behavior in mechanical applications like engines and compressors, governed by four fundamental laws.

Example

In a gas turbine, thermodynamics analyzes heat input and work output to optimize efficiency.

2 Basic Concepts

2.1 System, Control Volume, Surroundings, Boundary, Universe

Definition

A **system** is a specified quantity of matter or region under study, separated by a **boundary** (real or imaginary, fixed or movable) from its **surroundings** (everything external). A **control volume** is a fixed spatial region allowing mass and energy flow. The **universe** is the system plus surroundings.

Example

In a piston-cylinder, the gas is the system, the cylinder walls form the boundary, and the external environment is the surroundings.

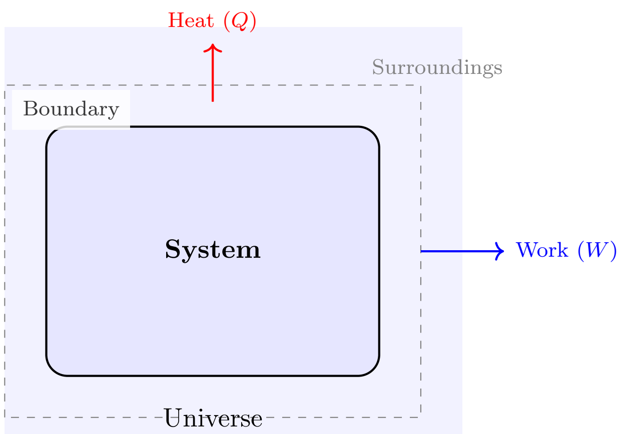

Figure 1: System, Surroundings, Boundary, and Universe with Heat and Work Transfer

2.2 Types of Systems

Definition

Systems are classified by their interactions: **Closed system** (fixed mass, allows energy transfer, $\Delta U = Q - W$); **Open system** (control volume, allows mass and energy transfer, $\dot{m}_{in} - \dot{m}_{out} = \frac{dm}{dt}$); **Isolated system** (no mass or energy transfer).

Example

A sealed piston-cylinder is a closed system, allowing heat/work transfer. A gas turbine is an open system with air flow.

Summary Table

System Type	Mass Transfer	Energy Transfer	Example
Closed	No	Yes (Heat, Work)	Piston-cylinder
Open	Yes	Yes (Mass, Heat, Work)	Gas turbine
Isolated	No	No	Insulated tank

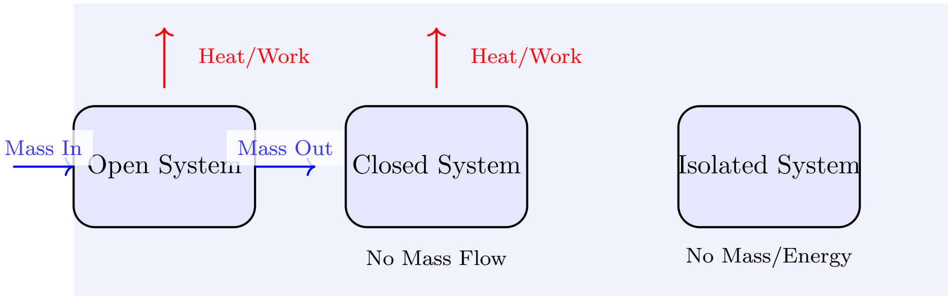

Figure 2: Types of Systems: Open, Closed, Isolated

2.3 Macroscopic and Microscopic Viewpoints

Definition

The **macroscopic viewpoint** analyzes systems using bulk properties (P, V, T) without molecular details. The **microscopic viewpoint** studies molecular behavior (kinetic and potential energies), forming the basis of statistical mechanics.

Example

Macroscopically, a compressor's gas is described by P and T. Microscopically, its behavior arises from molecular collisions.

2.4 Concept of Continuum

Definition

The **continuum concept** assumes matter is continuous, ignoring molecular structure, allowing use of differential calculus for properties like density ($\rho = \frac{dm}{dV}$). It is valid when system size $L \gg \lambda$ (mean free path), with Knudsen number $Kn = \frac{\lambda}{L} \ll 1$.

Example

In a gas turbine blade, the continuum assumption enables calculation of pressure and density gradients.

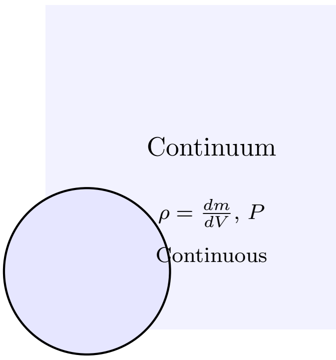

Figure 3: Continuum Concept: Homogeneous Matter

2.5 Thermodynamic Equilibrium

Definition

Thermodynamic equilibrium occurs when a system has uniform temperature (thermal equilibrium), uniform pressure (mechanical equilibrium), and no net chemical reactions or diffusion (chemical equilibrium), resulting in no spontaneous changes.

Example

A gas in a sealed piston-cylinder reaches thermodynamic equilibrium when temperature and pressure are uniform with no reactions.

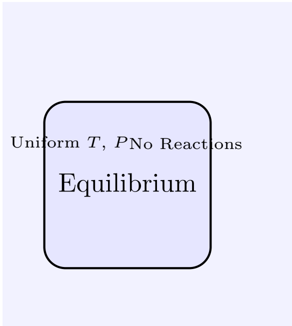

Figure 4: Thermodynamic Equilibrium: Uniform Properties

2.6 State, Property, Process, and Cycle

Definition

State is defined by a system's properties at an instant, fixed by two independent properties (state postulate). **Properties** are measurable, either **intensive** (e.g., P, T) or **extensive** (e.g., V, U). A **process** is a state change (e.g., isothermal: $T = \text{const}$, isobaric: $P = \text{const}$). A **cycle** is a sequence of processes returning to the initial state, with $\Delta U = 0$, efficiency $\eta = \frac{W_{net}}{Q_{in}}$.

Example

In a steam engine, a cycle involves isothermal expansion and adiabatic compression, producing net work.

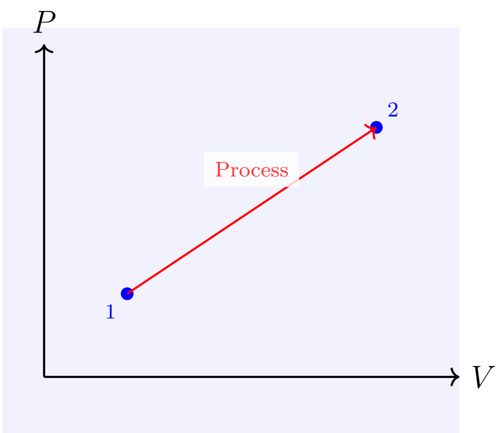

Figure 5: P–V Diagram: Process Between States

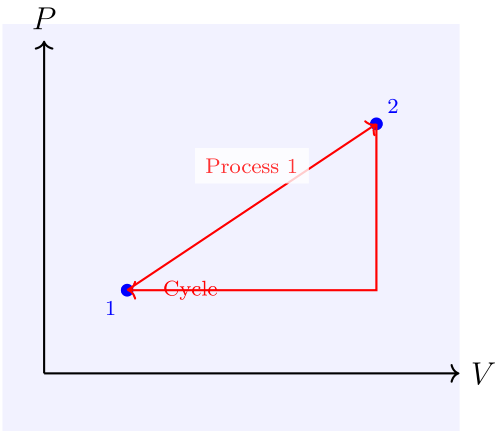

Figure 6: P–V Diagram: Cycle with $\Delta U = 0$

2.7 Exact and Inexact Differentials

Definition

An **exact differential** represents a state function (e.g., U, T), independent of path: $dU = \left(\frac{\partial U}{\partial T}\right)_V dT + \left(\frac{\partial U}{\partial V}\right)_T dV$. An **inexact differential** (e.g., δQ, δW) depends on the process path. For $f(x, y) = M dx + N dy$, exactness requires: $\left(\frac{\partial M}{\partial y}\right)_x = \left(\frac{\partial N}{\partial x}\right)_y$.

Example

Internal energy U has an exact differential, while work $\delta W = P dV$ in a piston-cylinder is inexact, varying with path.

2.8 Reversibility and Irreversibility

Definition

A **quasi-static process** occurs infinitely slowly, maintaining equilibrium, enabling calculations like $W = \int P dV$. A **reversible process** restores system and surroundings to initial states, requiring quasi-static conditions, no friction, infinitesimal temperature gradients, and no unrestrained expansion. An **irreversible process** generates entropy due to friction, finite ΔT heat transfer, free expansion, or mixing, reducing work output.

Causes of Irreversibility

- **Friction**: Converts work to heat (e.g., piston friction).
- **Finite ΔT Heat Transfer**: Causes entropy generation.
- **Free Expansion**: Increases entropy without work.
- **Mixing or Deformation**: Leads to irreversible changes.

Example

Reversible compression in a piston-cylinder maximizes work via a smooth P–V curve. Irreversible expansion with friction reduces work.

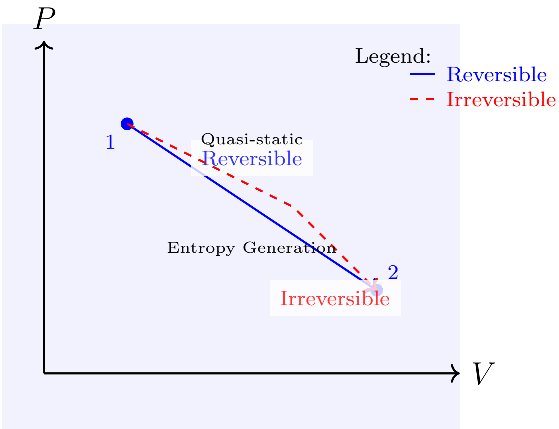

Figure 7: Reversible (Solid) vs. Irreversible (Dashed) Processes on P–V Diagram

3 Energy and Its Forms

Definition

Energy is the capacity to do work, comprising **macroscopic** forms (kinetic: $\frac{1}{2}mv^2$, potential: mgz) and **microscopic** internal energy (U, from molecular kinetic and potential energies). Changes in internal energy follow the first law: $\Delta U = Q - W$.

Example

In a turbine, internal energy of steam increases with heat, while kinetic energy changes with blade motion.

$$\boxed{E_{total} = U + \frac{1}{2}mv^2 + mgz}$$

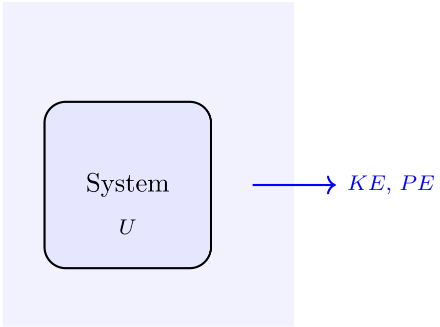

Figure 8: Energy Forms: Internal (U), Kinetic (KE), Potential (PE)

4 Work and Heat

Definition: Work

Work is energy transfer not due to temperature difference, with forms: **Boundary** ($W = \int PdV$), **Shaft** ($W = 2\pi NT$), **Electrical** ($W = \int Vdq$), **Magnetic** ($W = \int HdM$), **Gravitational** ($W = mg\Delta z$), **Spring** ($W = \frac{1}{2}k(x_2^2 - x_1^2)$). Positive work is done by the system.

Definition: Heat

Heat is energy transfer due to a temperature difference, occurring via conduction, convection, or radiation. It is path-dependent, positive when added to the system.

Example

In a piston-cylinder, boundary work is done during gas expansion, while heat addition in a boiler increases internal energy.

$$\boxed{\Delta U = Q - W}$$

Work Forms Table

Work Type	Formula	Example
Boundary	$W = \int PdV$	Piston-cylinder expansion
Shaft	$W = 2\pi NT$	Turbine rotation
Electrical	$W = \int Vdq$	Electric motor
Magnetic	$W = \int HdM$	Magnetic actuator
Gravitational	$W = mg\Delta z$	Lifting a weight
Spring	$W = \frac{1}{2}k(x_2^2 - x_1^2)$	Spring compression

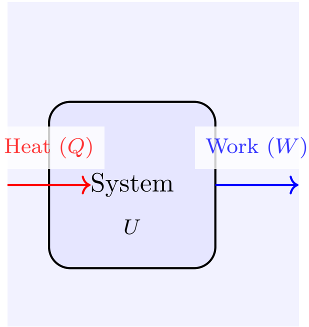

Figure 9: Heat and Work Transfers Affecting Internal Energy

5 Gas Laws and Behavior

5.1 Ideal Gas Laws

Definition

An **ideal gas** follows: **Boyle's Law** (PV = constant, isothermal), **Charles's Law** ($\frac{V}{T}$ = constant, isobaric), **Gay-Lussac's Law** ($\frac{P}{T}$ = constant, isochoric), and **Avogadro's Law** (equal volumes, same T, P have equal moles). Combined: $PV = nRT$, where $R = 8.314\,\text{J/mol·K}$.

Example

In a compressor, Boyle's law predicts increased pressure as volume decreases at constant temperature.

$$\boxed{PV = nRT}$$

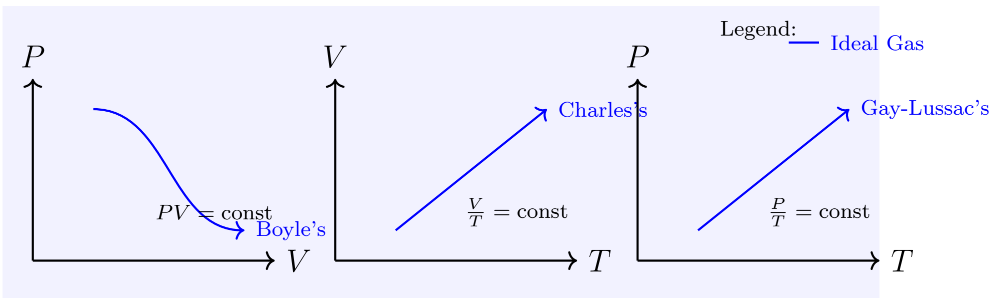

Figure 10: Gas Laws: Boyle's, Charles's, Gay-Lussac's

5.2 Real Gas Behavior

Definition

A **real gas** deviates from ideal behavior at high pressure or low temperature due to intermolecular forces and molecular volume. The van der Waals equation corrects this: $\left(P + \frac{a}{V^2}\right)(V - b) = RT$, where a accounts for attraction, b for volume.

Example

In a high-pressure gas cylinder, the van der Waals equation predicts real gas behavior.

5.3 Law of Corresponding States

Definition

The **law of corresponding states** states that gases show similar behavior when expressed in reduced properties: $P_r = P/P_c$, $T_r = T/T_c$, $V_r = V/V_c$ (critical point values). The compressibility factor $Z = \frac{PV}{nRT}$ is similar for gases at same P_r, T_r.

Example

In gas storage tanks, reduced properties predict deviations from ideal gas behavior.

5.4 Properties of Gas Mixtures

Definition

For gas mixtures, **Dalton's Law** states total pressure is the sum of partial pressures: $P = \sum P_i$, where $P_i = x_i P$ (mole fraction). **Amagat's Law** states total volume is the sum of partial volumes: $V = \sum V_i$, where $V_i = \frac{n_i RT}{P}$.

Example

In a gas turbine, Dalton's law calculates partial pressures of air and fuel gases.

6 Zeroth Law of Thermodynamics

Definition

The **Zeroth Law** states that if two systems (A and B) are each in thermal equilibrium with a third system (C), they are in thermal equilibrium with each other, establishing temperature as a measurable, transitive property.

Example

A thermometer (system C) ensures two gas tanks (A and B) have equal temperatures, used in engine calibration.

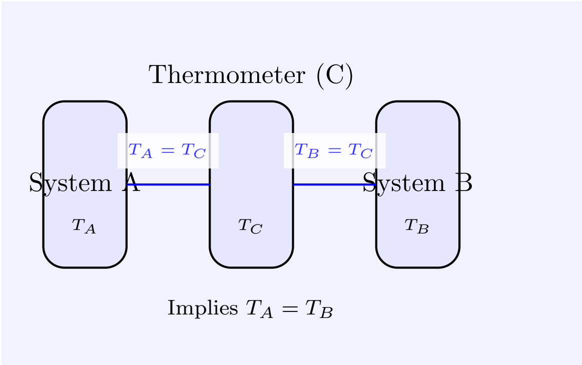

Figure 11: Zeroth Law: Thermal Equilibrium Implies $T_A = T_B$

6.1 Temperature Measurement and Scales

Definition

Temperature is measured using thermometers (e.g., mercury, thermocouple) based on properties like thermal expansion or electrical resistance. Standard scales include:

- **Celsius (°C)**: Ice point at 0 °C, steam point at 100 °C (1 atm).
- **Fahrenheit (°F)**: Ice point at 32 °F, steam point at 212 °F (1 atm).
- **Kelvin (K)**: Absolute scale, zero at -273.15 °C (absolute zero), no negative values.
- **Rankine (°R)**: Absolute scale, zero at -459.67 °F, used in engineering.

Kelvin is the SI unit for thermodynamic calculations.

Conversion Formulas

- Celsius to Fahrenheit: $T(F) = \frac{9}{5}T(C) + 32$
- Fahrenheit to Celsius: $T(C) = \frac{5}{9}[T(F) - 32]$
- Celsius to Kelvin: $T(K) = T(C) + 273.15$
- Kelvin to Celsius: $T(C) = T(K) - 273.15$
- Fahrenheit to Rankine: $T(R) = T(F) + 459.67$
- Rankine to Fahrenheit: $T(F) = T(R) - 459.67$
- Kelvin to Rankine: $T(R) = \frac{9}{5}T(K)$
- Rankine to Kelvin: $T(K) = \frac{5}{9}T(R)$

Derivation: Celsius to Fahrenheit

The Celsius and Fahrenheit scales are linear, with ice points at 0 °C, 32 °F, and steam points at 100 °C, 212 °F. The temperature difference between ice and steam points is 100 °C or 180 °F, so the ratio is $\frac{180}{100} = \frac{9}{5}$. For a temperature $T(C)$, the Fahrenheit equivalent is:

$$T(F) = \frac{9}{5}T(C) + 32$$

The +32 accounts for the Fahrenheit ice point offset.

Example

In an engine, a thermocouple reads 300 °C. Converting to Fahrenheit: $T(F) = \frac{9}{5} \cdot 300 + 32 = 572F$; to Kelvin: $T(K) = 300 + 273.15 = 573.15\,\text{K}$; to Rankine: $T(R) = \frac{9}{5} \cdot 573.15 = 1031.67\,°\text{R}$.

Conversion Table

Conversion	**Formula**
Celsius to Fahrenheit	$T(F) = \frac{9}{5}T(C) + 32$
Fahrenheit to Celsius	$T(C) = \frac{5}{9}[T(F) - 32]$
Celsius to Kelvin	$T(K) = T(C) + 273.15$
Kelvin to Celsius	$T(C) = T(K) - 273.15$
Fahrenheit to Rankine	$T(R) = T(F) + 459.67$
Rankine to Fahrenheit	$T(F) = T(R) - 459.67$
Kelvin to Rankine	$T(R) = \frac{9}{5}T(K)$
Rankine to Kelvin	$T(K) = \frac{5}{9}T(R)$

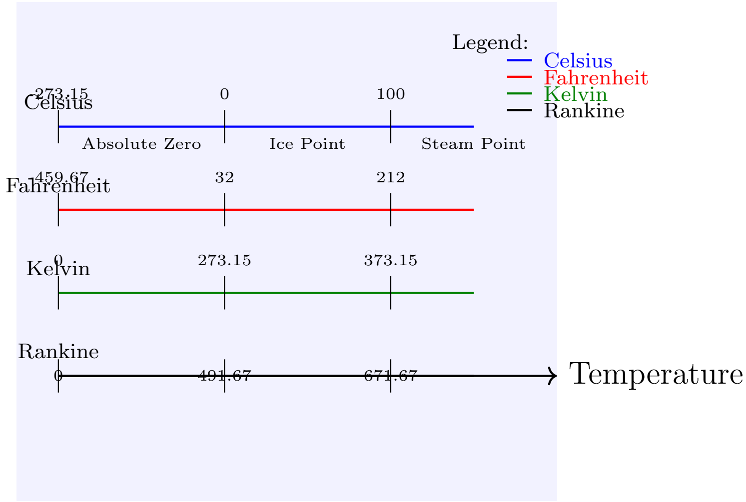

Figure 12: Comparison of Temperature Scales with Reference Points