

Chapter 8 – Thermodynamics

1. Introduction – Energy, Heat, and the Rules That Govern Them

Every activity in nature — from the burning of the Sun to the beating of your heart — involves energy transformation.

Thermodynamics is the branch of physics that studies heat, work, and energy and the laws that connect them.

While mechanics deals with motion, thermodynamics deals with energy transfer between systems and surroundings. It doesn't care what atoms do individually — it focuses on the overall behavior of macroscopic systems, like gases, engines, or refrigerators.

The word *thermodynamics* comes from two Greek words:

- “Therme” → heat
- “Dynamis” → power

So literally, thermodynamics is “the study of heat and power.”

2. Basic Concepts and Definitions

◆ System

A system is the part of the universe we are studying. Everything else is the surroundings.

Types of systems:

- Open system: exchanges both energy and matter (e.g., boiling water in open pan).
- Closed system: exchanges energy but not matter (e.g., gas in sealed cylinder).
- Isolated system: exchanges neither energy nor matter (e.g., thermos flask).

◆ State Variables

These describe the condition of a system — pressure (P), volume (V), temperature (T), and internal energy (U).

If these are fixed, the system's state is fixed.

◆ Process

When a system changes from one state to another, it undergoes a thermodynamic process.

- Isothermal: temperature constant ($\Delta T = 0$)
- Adiabatic: no heat exchange ($Q = 0$)
- Isobaric: pressure constant
- Isochoric: volume constant

◆ Quasi-static Process

A process so slow that the system remains nearly in equilibrium throughout. Only these processes are *reversible* and measurable.

🔥 3. Work, Heat, and Internal Energy

◆ Work (W)

In thermodynamics, work is done by or on a gas when its volume changes under pressure.

For infinitesimal change:

$$\delta W = P dV \quad \delta W = P dV$$

- Expansion → work done by gas (+ve)
- Compression → work done on gas (-ve)

Total work:

$$W = \int_{V_1}^{V_2} P dV = \int_{V_1}^{V_2} P dV$$

Graphically, area under P–V curve = work done.

◆ Heat (Q)

Heat is the energy transfer due to temperature difference.

Heat always flows from higher to lower temperature.

◆ Internal Energy (U)

It's the total microscopic energy (kinetic + potential) of all molecules.

It changes when heat is added or work is done.

ΔU depends only on initial and final states, not on the path — hence it's a state function.

4. The First Law of Thermodynamics

This is the law of conservation of energy in thermal form.

Energy can neither be created nor destroyed — it can only change form.

$$\Delta Q = \Delta U + \Delta W \quad \Delta Q = \Delta U + \Delta W$$

or equivalently,

$$\Delta U = \Delta Q - \Delta W \quad \Delta U = \Delta Q - \Delta W$$

where

- Q = heat supplied to system
 - W = work done by system
 - U = internal energy change
- ◆ Sign Conventions
- Heat supplied to system $\rightarrow +Q$
 - Heat rejected $\rightarrow -Q$
 - Work done by system $\rightarrow +W$
 - Work done on system $\rightarrow -W$

◆ Applications

1. Isochoric Process ($\Delta V = 0$):

No work done $\rightarrow \Delta Q = \Delta U$

e.g., heating gas in rigid container.

2. Isothermal Process ($\Delta T = 0$):

For ideal gas, $\Delta U = 0 \rightarrow \Delta Q = \Delta W$

All heat converts to work.

3. Adiabatic Process ($Q = 0$):

$\Delta U = -\Delta W \rightarrow$ internal energy decreases if gas expands.

4. Cyclic Process:

System returns to initial state $\rightarrow \Delta U = 0 \rightarrow Q = W$

5. Work Done in Various Processes

(a) Isothermal Process (for ideal gas)

$$PV=nRT \Rightarrow P=nRT/V \Rightarrow P=VnRT$$

$$W=\int_{V_1}^{V_2} P dV = nRT \ln(V_2/V_1)$$

$$W=nRT \ln(V_2/V_1)$$

(b) Adiabatic Process

Using Poisson's law: $PV^\gamma = \text{constant}$ $PV^\gamma = \text{constant}$

$$W=P_1V_1 - P_2V_2 = \frac{P_1V_1 - P_2V_2}{\gamma - 1}$$

and

$$T^\gamma = \text{constant}, PV^\gamma = \text{constant} \quad T^\gamma = \text{constant}, PV^\gamma = \text{constant}$$

$$\gamma = C_p/C_v$$

For monoatomic gas, $\gamma = 5/3$; diatomic ≈ 1.4 .

6. Specific Heats of a Gas

Specific heat (C) = amount of heat required to raise temperature of 1 mole by 1°C .

- C_p : at constant pressure
- C_v : at constant volume

Relation (Mayer's relation):

$$C_p - C_v = R$$

and

$$\gamma = \frac{C_p}{C_v}$$

7. Second Law of Thermodynamics

The first law tells *how energy is conserved*, the second law tells *why all processes have direction*.

"It is impossible to construct a machine which converts all absorbed heat into work without leaving any effect on surroundings."

Statements:

- Kelvin-Planck: No engine can be 100% efficient.
- Clausius: Heat cannot flow from cold to hot body without external work.

This introduces entropy, the measure of randomness or disorder.

$$\Delta S = \frac{\Delta Q_{rev}}{T} \quad T \text{ in K}$$

For isolated systems, $\Delta S \geq 0$ (entropy always increases).

8. Reversible and Irreversible Processes

- Reversible: infinitely slow, no friction, system always in equilibrium. (Ideal, not real.)
 - Irreversible: spontaneous, frictional losses, entropy increases. (All real processes.)
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9. Heat Engine and Efficiency

A heat engine converts heat into work by operating between two reservoirs:

Component	Function
Source	provides heat Q_1 at temperature T_1
Sink	absorbs remaining heat Q_2 at T_2

Working Substance does work $W = Q_1 - Q_2$

Efficiency:

$$\eta = \frac{W}{Q_1} = \frac{Q_1 - Q_2}{Q_1} = 1 - \frac{Q_2}{Q_1}$$

For perfect (Carnot) engine:

$$\eta = 1 - \frac{T_2 - T_1}{T_2} = 1 - \frac{T_2 - T_1}{T_2} \eta = 1 - \frac{T_2 - T_1}{T_2}$$

Higher temperature difference → higher efficiency.

10. Carnot Engine

Invented by Sadi Carnot, it's an ideal reversible engine operating in four steps:

1. Isothermal Expansion at T_1
2. Adiabatic Expansion ($T_1 \rightarrow T_2$)
3. Isothermal Compression at T_2
4. Adiabatic Compression ($T_2 \rightarrow T_1$)

It represents the maximum possible efficiency any engine can have.

No real engine can beat a Carnot engine.

11. Refrigerator or Heat Pump

Reverse of heat engine: it absorbs heat from cold reservoir and releases it to hot reservoir using external work.

$$Q_1 = Q_2 + W \quad Q_1 = Q_2 + W$$

Coefficient of Performance (COP):

$$\beta = \frac{Q_2}{W} = \frac{T_2 - T_1}{T_2} = \frac{T_2 - T_1}{T_2} \beta = \frac{W}{Q_2} = \frac{T_1 - T_2}{T_2}$$

Higher COP → better refrigerator.

12. Applications in Real Life

- Steam engines and internal combustion engines apply thermodynamic cycles.
 - Refrigerators and air conditioners are reversed heat engines.
 - Human metabolism converts chemical energy into work and heat.
 - Stars and the Sun obey thermodynamic balance between radiation and gravity.
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13. Important Graphs

- P–V diagram of isothermal → hyperbola.
 - Adiabatic → steeper curve.
 - Work done → area under curve.
 - Carnot cycle → closed loop (area = work).
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14. Entropy and the Arrow of Time

Entropy is what gives direction to time — systems evolve from order to disorder.
For reversible processes, $\Delta S = 0$; for irreversible, $\Delta S > 0$.

The entire universe moves toward maximum entropy — this is the second law on a cosmic scale.

15. Summary

1. Thermodynamics connects heat, work, and energy.
2. First law → $\Delta Q = \Delta U + \Delta W$.
3. For ideal gases:
 - Isothermal → $PV = \text{constant}$
 - Adiabatic → $PV^\gamma = \text{constant}$
4. Second law introduces entropy and defines efficiency limits.
5. No machine is 100% efficient.
6. Carnot cycle → ideal reversible engine.
7. Refrigerator COP = $T_2 / (T_1 - T_2)$.
8. Energy is conserved; entropy always increases.