

# **PHY101: Introduction to Physics I**

**Monsoon Semester 2024**

**Lecture 33**

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## **Previous Lectures**

**Introduction to thermodynamics**

## **This Lecture**

**First law of thermodynamics**

# FIRST LAW OF THERMODYNAMICS

- First law of thermodynamics is basically the statement of Energy conservation in a broader sense.
- The total energy of an isolated system is constant; energy can be transformed from one form to another, but cannot be created or destroyed.

$$\textit{Energy}(E) = \text{constant}$$

# FIRST LAW OF THERMODYNAMICS

- Using **Clausius sign convention** this can be expressed as,

$$\Delta U = Q - W$$



- Here  $\Delta U$  signifies the change in internal energy of the system.  $Q$  is the heat transferred **to** the system by the environment and  $W$  is the work done **by** the system on the environment.
- If  $W$  is +ve ( $-W$  is negative), that means work is being done by the system on the environment. If  $W$  is -ve ( $-W$  is +ve), that means work is being done by the environment on the system.
- Note that for  $Q = 0$ ,  $\Delta U$  is negative if  $W$  is +ve, i.e., if work is done by the system on the environment. This is reasonable as the internal energy of the system must decrease in this process. Similarly, if  $W$  is -ve,  $\Delta U$  is +ve, i.e., system's internal energy increases.
- This sign convention is well suited for studying heat engines, which provide useful work that is regarded as positive.

**We will use this convention throughout.**

# FIRST LAW OF THERMODYNAMICS

- Using **IUPAC sign convention** first law is written as

$$\Delta U = Q + W$$

- Here  $\Delta U$  signifies the change in internal energy of the system.  $Q$  is the heat transferred **to** the system by the environment and  $W$  is the work done **on** the system by the environment.
- This convention considers all net energy transfers to the system as positive and all net energy transfers from the system as negative, irrespective of any use for the system as an engine or other device.
- If  $W$  is +ve, that means work is being done on the system by the environment. If  $W$  is -ve, that means work is being done by the system on the environment.
- Note that for  $Q = 0$ ,  $\Delta U$  is positive if  $W$  is +ve, i.e., if work is done  the system  the environment. This is reasonable as the internal energy of the system must increase in this process. Similarly, if  $W$  is -ve,  $\Delta U$  is -ve, i.e., system's internal energy decreases.



# FIRST LAW OF THERMODYNAMICS

In the differential form the first law is written as follows:

**Clausius sign convention:**

$$dU = \delta Q - \delta W$$

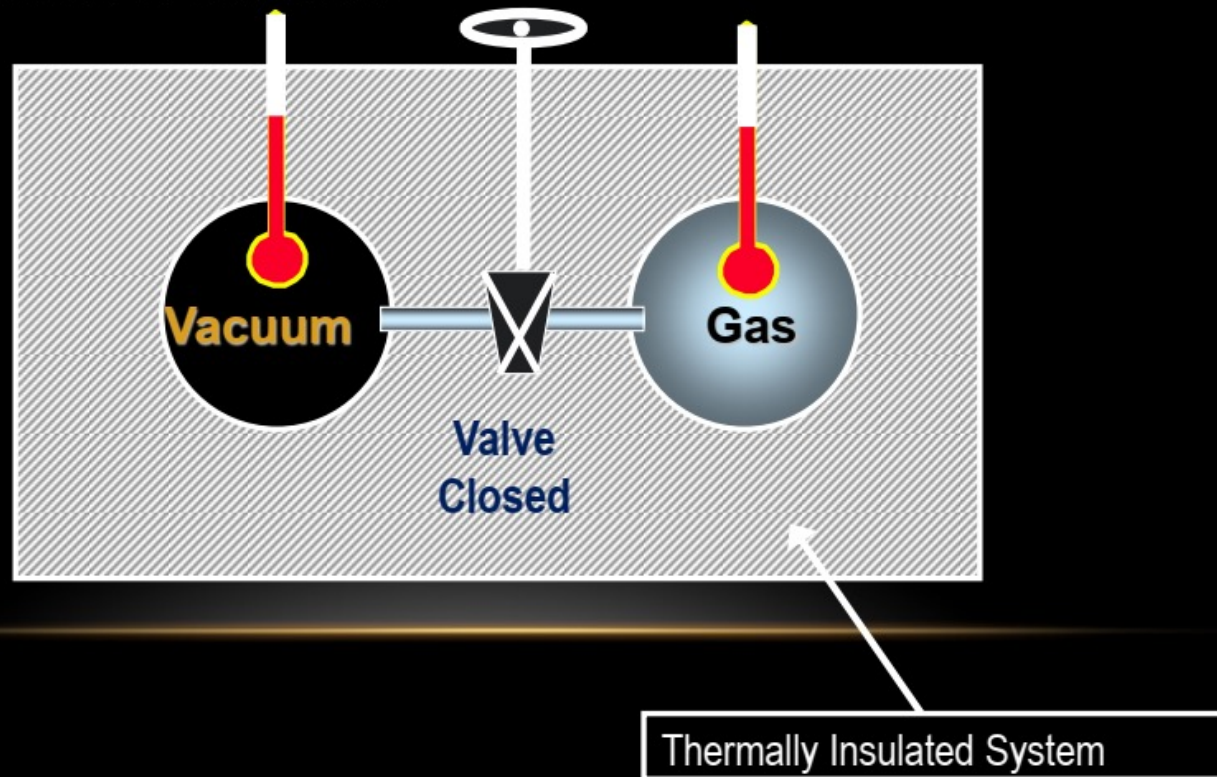
**IUPAC sign convention:**

$$dU = \delta Q + \delta W$$

The notation  $\delta$  with  $Q$  and  $W$  signify that they are **path functions**, i.e., they depend on how the thermodynamic system changes from the initial state to final state. While  $U$  is a **state function**– it is independent of the path take to reach the final state from the initial.

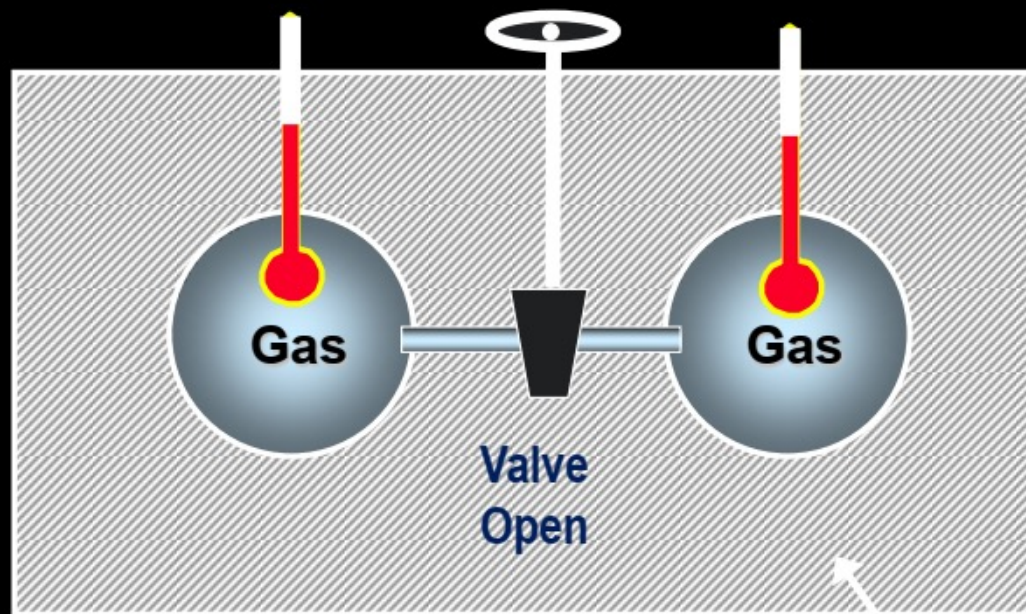
# JOULE EXPANSION

- Much of the early progress in thermodynamics was made in the study of the properties of gases. One of the early questions was whether or not gases cool on expansion.
- Joule designed an experiment to find out whether or not gases cool on expansion and if so how much.



The Joule expansion is an irreversible process in thermodynamics in which a volume of gas is kept in one side of a thermally isolated container (via a small partition), with the other side of the container being evacuated; the partition between the two parts of the container is then opened, and the gas fills the whole container.

# JOULE EXPANSION



Thermally Insulated System



# JOULE EFFECT

- Gas (ideal) expanded to fill the container:

*Change in volume*

*Change in pressure*

- No external work was done:

*Air expanded into a vacuum*

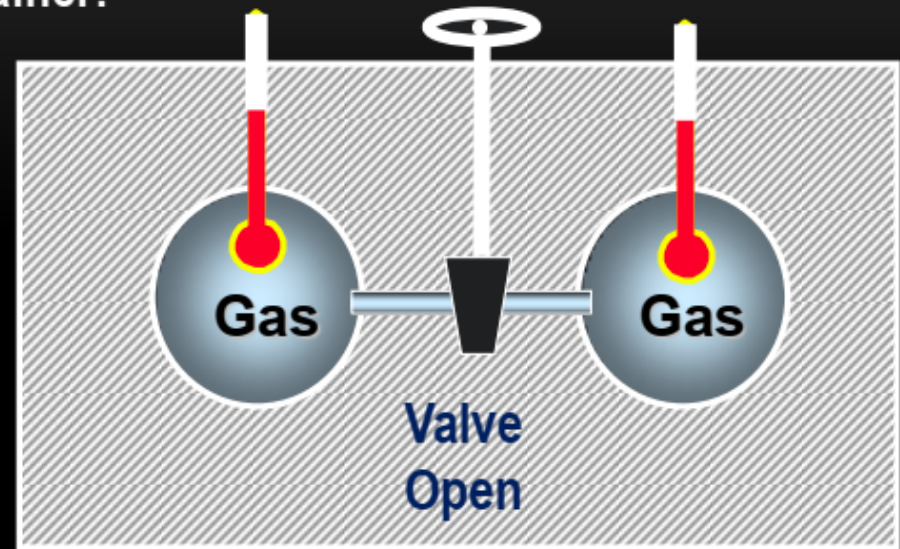
*within the system*

- No heat was added or subtract:

*Thermally insulated system*

- No change in internal energy
- No change in temperature

Internal energy is a function of only temperature.



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

$$\Delta U = 0$$

# THERMAL CAPACITIES (HEAT CAPACITY)

If  $dQ$  is a small amount of heat **absorbed** by a substance, and  $dT$  is the accompanying small change in temperature, then heat capacity is defined as

$$C = \frac{dQ}{dT}.$$

The heat capacity depends on the detailed nature of the system and meaning it is a physical property that scales with the size of a physical system. Thus it is an **extensive property**.

Two types of specific heat can be defined depending on the conditions under which the material receives heat.

**Constant volume:**

$$C_V = \left( \frac{dQ}{dT} \right)_V$$

**Constant pressure:**

$$C_P = \left( \frac{dQ}{dT} \right)_P$$

# SPECIFIC HEAT & MOLAR HEAT CAPACITY

**Specific Heat** of a substance is the heat capacity per unit mass of the substance. Consequently specific heat is an **intensive property** - an intrinsic characteristic of a particular substance. It is usually represented by lower case  $c$ , in contrast to upper case  $C$  for the heat capacity. Sometimes  $s$  is also used to represent it.

As in the case of heat capacity, we have two kinds of specific heat: Specific heat at constant volume ( $c_V$ ) and Specific heat at constant pressure ( $c_P$ ).

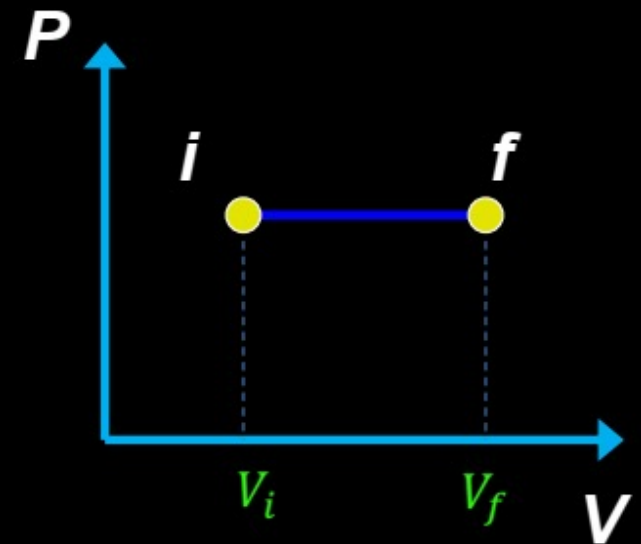
**Molar heat capacity** of a substance is the heat capacity per mole of the substance. It is an **intensive property**. It may be represented by the symbol  $C_{\text{mol}}$ .

# PV Diagrams and Types of Processes

## Isobaric process:

- The pressure remains constant during the process.
- The curve representing the process is parallel to the  $V$ -axis.

- Work done =  $\int_{V_i}^{V_f} P dV$   
$$= P \int_{V_i}^{V_f} dV = P(V_f - V_i)$$

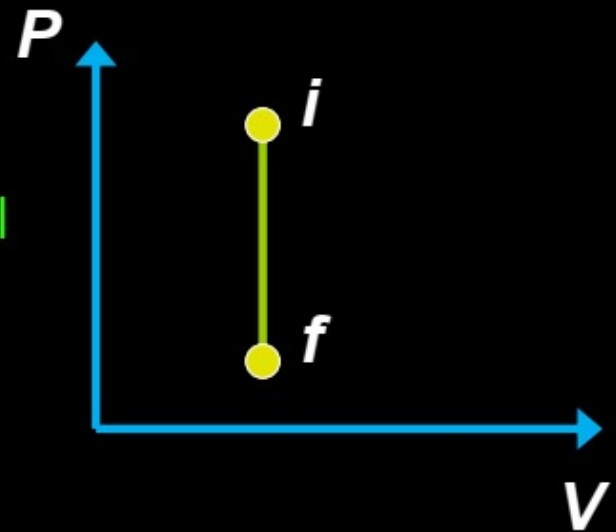


# PV Diagrams and Types of Processes

Isochoric process:

(Also called an isometric process or an isovolumetric process)

- The volume remains constant during the process.
- The curve representing the process is parallel to the  $P$ -axis.
- Work done =  $\int_{V_i}^{V_f} P dV = 0$   
since  $V_i = V_f$ .





# PV Diagrams and Types of Processes

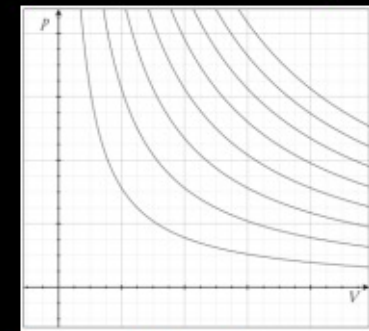
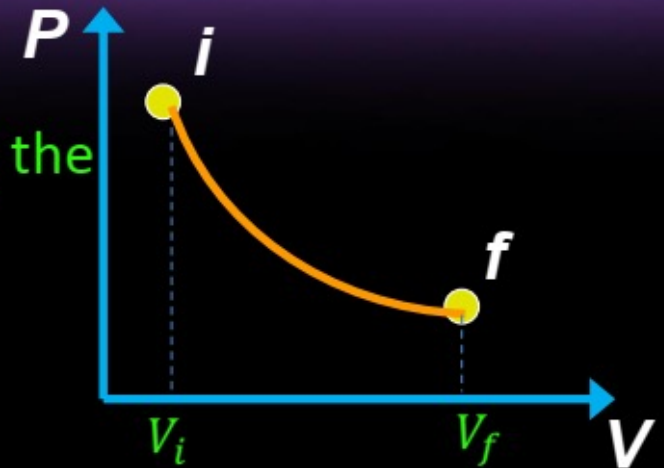
## Isothermal process:

- The temperature remains constant during the process.

$$P = \frac{nRT}{V} = \frac{\text{Constant}}{V}$$

( $n$  represents the number of moles.)

- Work done =  $\int_{V_i}^{V_f} P dV$   
$$= nRT \int_{V_i}^{V_f} \frac{1}{V} dV = nRT \ln(V_f/V_i)$$



Several Isotherms

# PV Diagrams and Types of Processes

Adiabatic process:

The process in which no exchange of heat takes place between the environment and the system. Thus,

$$dQ = 0.$$

First law of thermodynamics , therefore gives,

$$dU = -dW = -PdV$$

# PV Diagrams and Types of Processes

Adiabatic process:

In an adiabatic process,

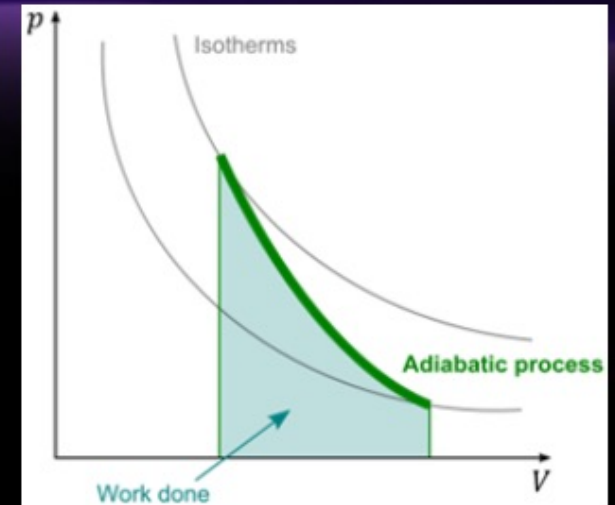
$$PV^\gamma = C(\text{constant}), \text{ where } \gamma = \frac{C_P}{C_V}.$$

$$\text{Thus Work done} = \int_{V_i}^{V_f} P dV$$

$$= \int_{V_i}^{V_f} \frac{C}{V^\gamma} dV = \frac{1}{1-\gamma} (C V_f^{1-\gamma} - C V_i^{1-\gamma})$$

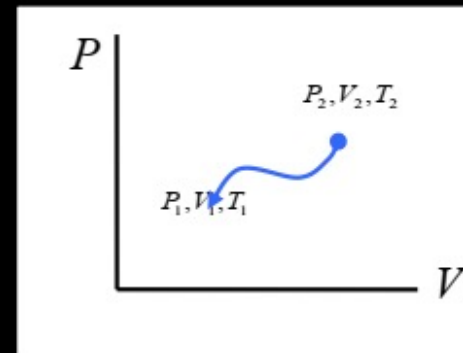
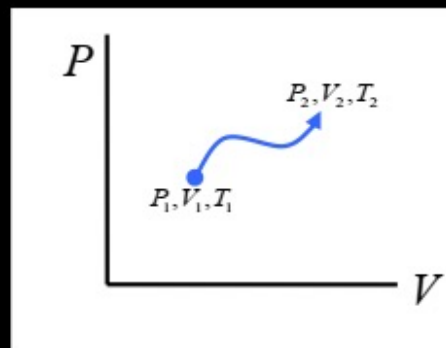
$$= \frac{1}{1-\gamma} (P_f V_f - P_i V_i) = -\frac{C_V}{R} (P_f V_f - P_i V_i) = -nC_V(T_f - T_i)$$

$$(\text{Since } C V_f^{-\gamma} = P_f, C V_i^{-\gamma} = P_i \text{ and } C_P - C_V = R)$$



# Reversible Processes

- In a reversible process, the system undergoing the process can be returned to its initial conditions along the same path shown on a  $PV$  diagram, and every point along this path is an equilibrium state.
- Therefore it involves infinitesimally small change in each step and involves no dissipation of energy.
- It would take infinite amount of time for a reversible process to complete.





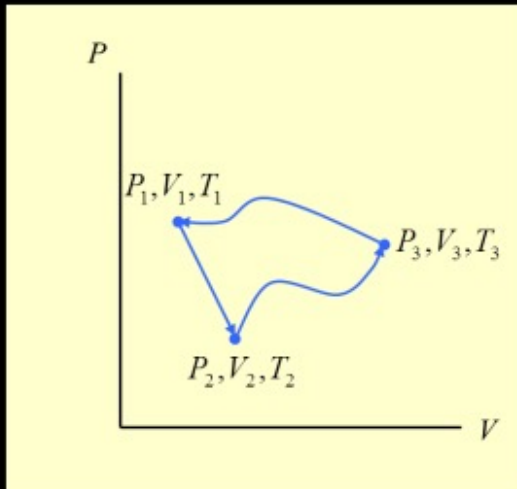
# Irreversible Processes

- A process that does not satisfy these requirements (associated with the reversible process) is irreversible.
- In an irreversible process, the changes involved are finite. As a consequence the system is not at equilibrium throughout the process.
- Heat loss is associated with it.
- Irreversible processes cannot be drawn on a PV diagram.
- Real world processes are irreversible.



# Reversible Cycle

- A series of reversible processes can be plotted on a PV diagram.
- A reversible cycle is constituted by a set of reversible processes that return to the initial state.



## Practice Problems

1. 5000 J of heat are added to two moles of an ideal monatomic gas, initially at a temperature of 500 K, while the gas performs 7500 J of work. What is the final temperature of the gas?

### Solution

$$\Delta U = Q - W = 5000 \text{ J} - 7500 \text{ J} = -2500 \text{ J}$$

$$\Delta U = -2500 \text{ J} = (3/2)nR\Delta T = (3/2)(2)(8.31)\Delta T$$

$$\rightarrow \Delta T = -100 \text{ K}$$

$$\rightarrow T_f = \mathbf{500 \text{ K} - 100 \text{ K} = 400 \text{ K}}$$

*comment*: the gas does more work than it takes in as heat, so it must use 2500 J of its internal energy.

## Practice Problems

2. Compute the internal energy change and temperature change for the two processes involving 1 mole of an ideal monatomic gas.
- (a) 1500 J of heat are added to the gas and the gas does no work and no work is done on the gas
  - (b) 1500 J of work are done on the gas and the gas does no work and no heat is added or taken away from the gas

### Solution

(a)

$$\begin{aligned}\Delta U &= Q - W = 1500\text{ J} - 0 = 1500\text{ J} \\ \Delta U &= 1500\text{ J} = \left(\frac{3}{2}\right)nR\Delta T = \left(\frac{3}{2}\right)(1)(8.31)\Delta T \\ &\rightarrow \Delta T = 120\text{ K}\end{aligned}$$

(b)

$$\begin{aligned}\Delta U &= Q - W = 0 - (-1500\text{ J}) = +1500\text{ J} \\ \Delta U &= 1500\text{ J} = \left(\frac{3}{2}\right)nR\Delta T = \left(\frac{3}{2}\right)(1)(8.31)\Delta T \\ &\rightarrow \Delta T = 120\text{ K}\end{aligned}$$

*Notice that in both processes, the change in internal energy is the same. We say that the internal energy is a “state function”. A state function depends only on the state of the system and not on the process that brings the system to that particular state.*