Unit 16

First Law of Thermodynamics

Teaching Periods 08 **Weightage %** 05

The Thar Coal Power Plant, located in the Thar Desert of Sindh, Pakistan, is a significant project aimed at utilizing the large reserves of lignite coal found in the region. The Thar Coal Power Plant is an example of how the first law of thermodynamics is applied in real-world energy systems.

In this unit student should be able to:

- Describe that heat flow and work are two forms of energy transfer between systems and calculate heat being transferred.
- Relate rise in temperature of a system increases its internal energy.
- Explain that internal energy is determined by the state of the system and that it can be expressed as the sum of the random distribution of kinetic and potential energies associated with the molecules of the system.
- Define thermodynamics and various terms associated with it.
- Calculate work done by a thermodynamic system during a volume change.
- Describe the first law of thermodynamics expressed in terms of the change in internal energy, the heating of the system and work done on the system.
- Explain that first law of thermodynamics expresses the conservation of energy.
- Describe the applications of first law of thermodynamics with the help of equations and graphs
- Solve the problems using the equations of first law of thermodynamics
- Define the terms specific heat and molar specific heats of a gas.
- Apply first law of thermodynamics to derive $C_p - C_v = R$.
- Solve the problems using equations of specific heat

Introduction:

In this unit, we explore key concepts in thermodynamics: heat, work, internal energy, the First Law of Thermodynamics, and molar specific heat. These concepts form the basis for understanding energy transfer and conservation within systems. Heat and work are fundamental modes of energy transfer, influencing changes in a system's internal energy. This unit provides a comprehensive overview, essential for grasping energy transformations and their applications in diverse fields.

16.1.1 Heat and Work:

Work and thermal energy are linked by the broader framework of energy conservation. The study of heat has evolved significantly over centuries. As shown in table 16.1.

Table 16.1: History of heat energy

18th Century	Isaac Newton	➢ Proposed the concept of "caloric" (a hypothetical fluid thought to be the substance of heat). ➢ Influenced future research.
18th Century	Joseph Black	➢ Introduced the concept of latent heat. ➢ Distinguished between heat and temperature.
18th Century	Count Rumford (Benjamin Thompson)	➢ Challenged the caloric theory of heat ➢ Demonstrated that heat is a form of motion, not a material substance
19th Century	Sadi Carnot	➢ Laid the foundations of thermodynamics ➢ Described the Carnot cycle
19th Century	James Prescott Joul	➢ Established the mechanical equivalent of heat ➢ Showed that energy can be converted from one form to another
19th Century	Rudolf Clausius and William Thomson (Lord Kelvin)	➢ Developed the concept of entropy ➢ Formulated the second law of thermodynamics
20th Century	Albert Einstein	➢ Provided evidence for the atomic theory of matter ➢ Helped establish the kinetic theory of heat
Modern Developments:		➢ Advances in statistical mechanics and quantum mechanics ➢ Connected heat to the behavior of atoms and molecules

These discoveries laid the foundation for modern thermodynamics, which explains heat as the kinetic energy of particles in motion.

Work can lead to changes in the internal energy of a system, contributing to thermal energy. The first law of thermodynamics encapsulates this relationship, emphasizing the interplay between heat transfer, work and changes in internal energy within a system.

Thermal Energy and Work:

Internal Energy and Heat (Q): In thermodynamics, thermal energy is associated with the internal energy of a system. Heat (Q) is the transfer of thermal energy between systems due to a temperature difference.

Heat Flow:

It flows spontaneously from a region of higher temperature to one of lower temperature. The rate at which heat is transferred can be quantified using the equation:

$$\boldsymbol{Q = mc\Delta T} \;\ldots\ldots..\; 16.1$$

Where Q is the heat transferred, m is the mass of the substance, c is the specific heat capacity of the substance, and ΔT is the temperature change.

Above equation, known as the heat transfer equation, highlights the dependence of heat transfer on the material properties and the extent of temperature change.

Conversion of Work to Thermal Energy:

Work done on a system can lead to an increase in its internal energy, contributing to thermal energy. For instance, when you rub your hands together, the mechanical work done generates heat due to friction, increasing the thermal energy of your hands.

The concept of PV work can be illustrated by imagining a gas in a cylinder equipped with a movable piston as shown in figure 16.1. The pressure exerted on the piston is essentially a force per unit area. When the gas undergoes expansion, the piston moves upward, necessitating energy input (a force acting over a distance). The work done can be expressed as the product of the force and distance. In the case of a gas in a cylinder (thermal system), this simplifies to the pressure multiplied by the change in volume.

$$W = P\Delta V \;\ldots\ldots\ldots\; 16.2$$

$$W = -P\Delta V$$

The negative sign is to keep our sign convention for energy going into or out of the system.

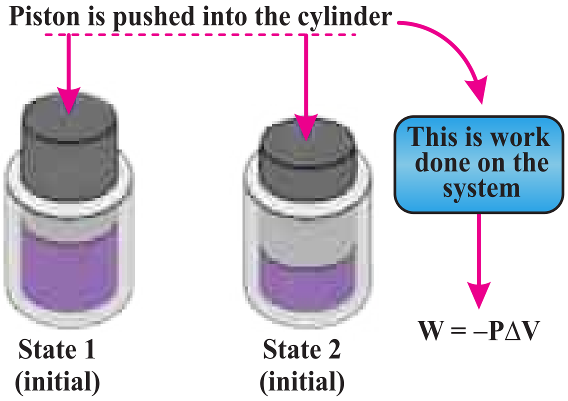


Figure 16.1 Work done by PV relation.

Self-Assessment Questions:

1. **Define heat is a form of energy.**
2. **Explain the basic concept of how work can be converted into thermal energy.**

16.2.1 Internal Energy:

Energy can generally be classified into two forms: kinetic and potential energy. The molecules of all substances contain both kinetic and potential energies. The amount of kinetic and potential energy of a substance depends on the phases of matter (solid, liquid or gas), this is known as the **internal energy**

The internal energy of a substance is defined as:

"***The sum of the random distribution of kinetic and potential energies within a system of molecules."***

The symbol for internal energy is *U*and its SI unit is **Joule** (J).

- The internal energy of a system is determined by:
  - Temperature
  - The random motion of molecules
  - The phase of matter: gases have the highest internal energy, solids have the lowest
- The internal energy of a system can increase by:
  - Doing work on it
  - Adding heat to it
- The internal energy of a system can decrease by:
  - Losing heat to its surroundings

DO YOU KNOW?

Processes such as heating, cooling, phase transitions, and chemical reactions can cause changes in the internal energy of a system

Internal Energy & Temperature:

The internal energy of an object is intrinsically related to its temperature. When a gas in a container is heated, the gas molecules move faster, increasing their kinetic energy. In a solid, where molecules are tightly packed, heating causes the molecules to vibrate more. In both liquids and solids, molecules have both kinetic and potential energy due to intermolecular forces that keep them close together. However, ideal gas molecules are assumed to have no intermolecular forces, meaning they only possess kinetic energy and no potential energy.The (change in) internal energy of an ideal gas is equal to:

$$\Delta U = \frac{3}{2} k\Delta T \ldots\ldots\ldots 16.3$$

Therefore, the change in internal energy is proportional to the change in temperature:

$$\Delta U \propto \Delta T$$

Where:ΔU = change in internal energy (J)
ΔT = change in temperature (K)

"As the container is heated up, the gas molecules move faster with higher kinetic energy and therefore higher internal energy". as shown in the figure 16.2.

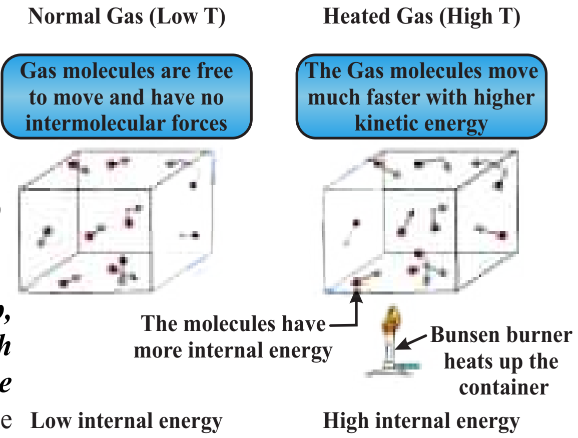


Figure 16.2 Internal Energy & Temperature

Self-Assessment Questions:

1. How the internal energy of an object is related to its system?
2. How does an increase in temperature affect the internal energy of a substance?

16.3.1: Define thermodynamics and various terms associated with it.

Thermodynamic terms and basic concepts:

The study of the flow of heat or any other form of energy into or out of a system as it undergoes a physical or chemical transformation is called thermodynamics.

A thermodynamic system is a specific portion of the physical universe that is chosen for analysis. It is defined by boundaries and the surrounding. ***The boundary is the physical or imaginary separation between the system and its surroundings. The surroundings refer to everything outside the boundaries of the system.*** This includes the rest of the universe that can interact with the system. separates the system from the rest of the universe. The system can be of any size or shape and can consist of a single substance or a combination of substances as shown in figure 16.3.

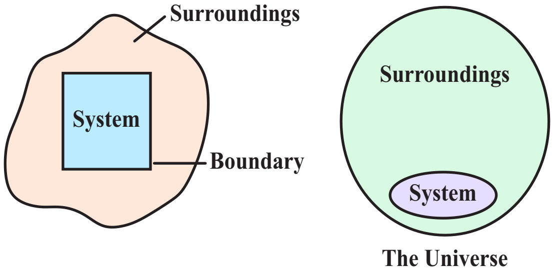


Figure 16.3 System, Boundary and Surrounding

Thermodynamic Systems:

1. **Open System:**

   In an open system, both *heat and matter can be exchanged* with the surroundings. For instance, steam (matter) can escape from the coffee, and heat can be lost to the surrounding air.

   Example: A cup of hot coffee.

2. **Closed System:**

   In a closed system, *only heat can be exchanged* with the surroundings, but matter cannot. In this example, heat can still transfer through the pot, but the water vapor (matter) is trapped inside the pot and cannot escape.

   Example: A pot with a lid.

3. **Isolated System:**

   In an isolated system, *neither heat nor matter can be exchanged with the surroundings*. A thermos flask is designed to prevent heat transfer, keeping the contents inside hot or cold for a long period without any exchange with the environment.

Systems are as shown in figure 16.4 respectively

Example: A thermos flask.

Open System **Closed System** **Isolated System**

Exchange of heat and matter **Exchange only heat** **No exchange**

Figure 16.4 Open, Close and Isolated system

16.3.2 Work Done by a Gas:

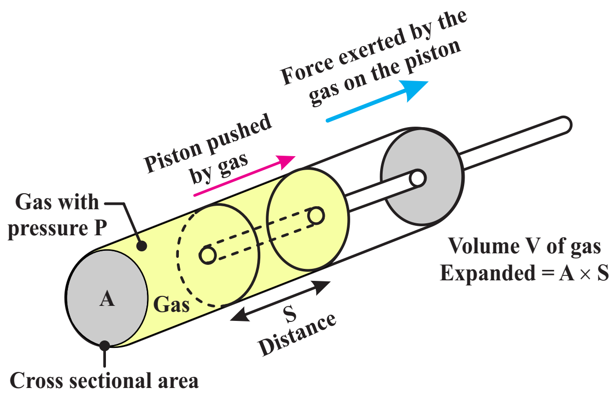
When a gas expands, it does work on its surroundings by exerting pressure on the walls of the container it's in. This is important, for example, in a steam engine where expanding steam pushes a piston to turn the engine. The work done when a volume of gas changes at constant pressure is defined as:

$$W = p\Delta V$$

Where:

- W = work done (J)
- p = external pressure (Pa)
- V = volume of gas (m^3)

For a gas inside a cylinder enclosed by a moveable piston, the force exerted by the gas pushes the piston outwards, Therefore, the gas does work on the piston



The gas expansion pushes the piston a distance S

Figure 16.5 Work Done by a Gas

Derivation:

The volume of gas is at constant pressure. This means the force F exerted by the gas on the piston is equal to:

$$F = p \times A$$

Where: p = pressure of the gas (Pa) and A = cross-sectional area of the cylinder (m^2)
The definition of work done is:

$$W = F \times S$$

Where: F = force (N) and S = displacement in the direction of force (m)
The displacement of the gas d multiplied by the cross-sectional area A is the increase in volume ΔV of the gas:

$$W = P \times A \times S \ldots\ldots 16.4$$

This gives the equation for the work done when the volume of a gas changes at constant pressure:

$$W = P\Delta V$$

Where: ΔV = increase in the volume of the gas in the piston when expanding (m^3)
This is assuming that the surrounding pressure p does not change as the gas expands, This will be true if the gas is expanding against the pressure of the atmosphere, which changes very slowly, When the gas expands (V increases), work is done by the gas. When the gas is compressed (V decreases), work is done on the gas.

Worked Example 16.1

The volume occupied by 1.00 mol of a liquid at 50 °C is 2.4×10^{-5} m^3. When the liquid is vaporized at an atmospheric pressure of 1.03×10^5Pa, the vapor has a volume of 5.9×10^{-2} m^3.The latent heat to vaporize 1.00 mol of this liquid at 50 °C at atmospheric pressure is 3.48×10^4 J. Determine change in internal energy ΔU of the system.

Solution:

Step 1: Write down the first law of thermodynamics

$\Delta U = Q - W$

Step 2: Write the value of heating q of the system

This is the latent heat, the heat required to vaporize the liquid = **3.48×10^4 J**

Step 3: Calculate the work done W

$$W = P.\Delta V$$

$\Delta V = V_f\text{-}V_i = 5.9 \times 10^{-2} - 2.4 \times 10^{-5} = 0.058976\ m^3$

p = atmospheric pressure = 1.03×10^5 Pa

$W = (1.03 \times 10^5) \times 0.058976 = 6074.528 = 6.07 \times 10^3$ J

Since the gas is expanding, this work done is negative

$W = -6.07 \times 10^3$ J

Step 4: Substitute the values into first law of thermodynamics

$\Delta U = 3.48 \times 10^4 \quad (-6.07 \times 10^3) = 28\,730 = 29\,000$ J

$T_f = = \frac{9}{5} \times 36.88 + 32$

$T_f = 98.4°F$

Result: Result: $T_f = 98.4°$

Worked Example 16.2

When a balloon is inflated, its rubber walls push against the air around it. Calculate the work done when the balloon is blown up from 0.015 m^3 to 0.030 m^3. Atmospheric pressure = 1.0 × 10^5 Pa.

Solution:

Step 1: write the known quantities and point out quantities to be found

$V_1 = 0.015\ m^3$
$V_2 = 0.030\ m^3$
$P = 1.0 \times 10^5$ Pa.
W= ?

Step 2: Write down the equation for the work done by a gas

$$W = P\Delta V$$

Step 3: Now First, calculate ΔV:

$$\Delta V = V_f - V_i$$

$\Delta V = 0.030\ m^3 - 0.015\ m^3 = 0.015\ m^3$

Step 4: Now, substitute the values into the formula:

$$W = -P\Delta V$$

W = (1.0 × 105) × 0.015 = 1500 J

Result: The work done is **1500 J**

16.3.3: The first law of thermodynamics, internal energy and the heating of the system& conservation of energy:

First law of thermodynamics:

The first law of thermodynamics is based on the idea that energy can neither be created nor destroyed in any thermodynamic system. So, it is a particular form of the law of conservation of energy, which deals only with the heat energy.

"This law states that every thermodynamic system possesses a state variable (U) called the internal energy".

In any thermodynamic process, when heat energy (ΔQ) is added to a system, this energy appears as an increase in the internal energy (ΔU) stored in the system plus the work done (ΔW) by the system on its surroundings.

$$\Delta Q = \Delta U + \Delta W \ldots\ldots\ldots. 16.5$$

ΔQ is taken positive when heat enters the system and negative when it leaves the system. ΔU is taken positive when the temperature of the system rises, while it is negative when the temperature of the system decreases.

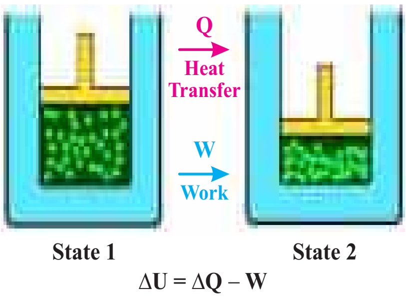


Figure 16.6 states of system

The first law of thermodynamics states that,

"The change in internal energy of a system equals the net heat transfer into the system minus the net work done by the system"

In equation form, the first law of thermodynamics is,

$$\Delta U = \Delta Q - W \;\ldots\ldots 16.6$$

ΔU →*change in internal energy* of the system

ΔQ → *net heat transferred into the system*

W →*net work done by the system*

The sum of all the kinetic (translational, rotational, and vibrational) and potential energies associated with the random motion of the atoms of a substance or system is the internal energy (U) of the substance.

According to the first law of thermodynamics Eq. No 16.6

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

It means that the change in the internal energy of a system is equal to the energy flowing in as heat minus the energy flowing out as work. In other words, the change in the internal energy of the system is measured by the energy retained by the system. The energy absorbed by the system changes the translational, vibrational, and rotatio nal kinetic energy of the molecules. It also changes the potential energy of the molecules due to intermolecular forces.

upon the initial and final states of the system and not on the path taken between these two states. If the state of a thermodynamic system is changed from state (A) to state (B) then the change in internal energy of the system is

$$\Delta U = U_B - U_A \;\ldots\ldots\ldots 16.7$$

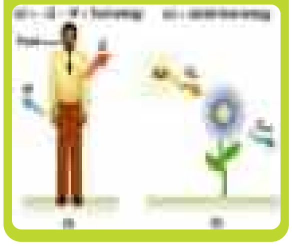
Putting this value in equation (16.6)

we have,

$$U_B - U_A = \Delta Q - \Delta W \;\ldots\ldots 16.8$$

DO YOU KNOW?

Examples of 1st law of thermodynamics Figure No.. (a) The first law of thermodynamics applied to metabolism. Heat transferred out of the body (Q) and work done by the body (W) remove internal energy, while food intake replaces it. (Food intake may be considered as work done on the body.) (b) Plants convert part of the radiant heat transfer in sunlight to stored chemical energy, a process called photosynthesis.



Applications of the first law thermodynamics:

Applications of first law of thermodynamics are all based on the principle that "the three forms of energy: internal energy; heat; and work can be inter converted."

The system is then often called the working substance. The mathematical formulation of the law $\Delta Q = \Delta U + \Delta W$ suggests the following processes that can be used for practical applications:

- The processes in which any one of the three terms of the equation is zero.
- The processes in which any one of the state variables P,V,T of the system are held constant.

(a) Isochoric Process:

"The thermodynamics process during which the volume of the system remains constant is called isochoric process".

We consider the gas contained in a cylinder having a conducting base and non-conducting walls and with a fixed piston at the end as shown in the Figure 16.7 Let heat ΔQ be imparted to the gas. The gas is then heated at constant volume. The pressure of the gas increases from P_1 to P_2 while its temperature increases from T_1 to T_2. Since the system neither expands nor contracts, work is neither done by the system nor on the system i.e $\Delta W=0$.

Using the first law of thermodynamics equation (Eq. No 16.5)

we have:

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

$$\Delta Q = \Delta U$$

This means that in an isochoric process the entire amount of heat supplied to the gas is converted to the internal energy of the gas. The pressure and temperature of the gas will increase. On the contrary, removal of heat from a system under isochoric condition will cause an equivalent decrease in the internal energy. The system will cool down and the pressure will fall.The graph of isochoric process is called an "isochor', which is a straight line, parallel to the pressure axis.

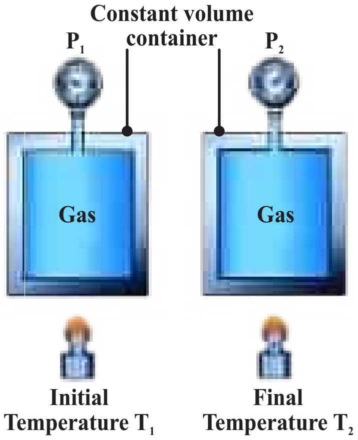


Figure 16.7 (a) Isochoric Process

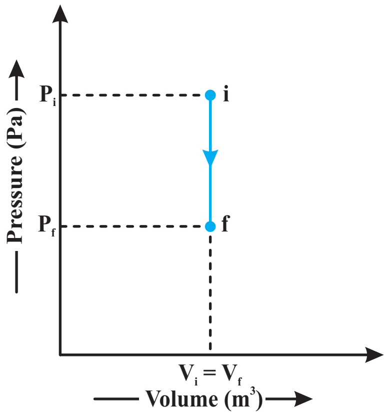


Figure 16.7 (b)

(b) Isobaric Process:

"The thermodynamics process during which the pressure is kept constant is called an isobaric process".

Isobaric expansion of a system is often used to convert heat into work. Practically all heat engines depend on the transformation of heat into work. Consider a gas contained in a cylinder having a conducting base and non-conducting walls and frictionless piston of cross-sectional area (A) as shown in figure 16.8(a).

When the gas is heated, a certain amount of heat energy is transferred into the system.

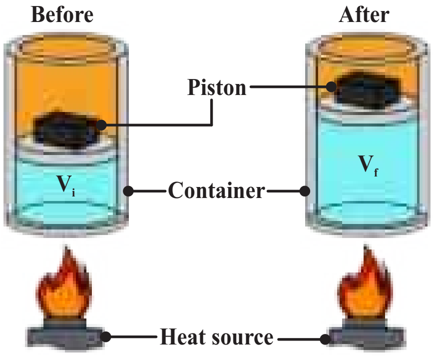


Figure 16.8(a) Isobaric Process

- The gas expands and moves the piston outward.
- Temperature changes from T_1 to T_2 and Volume changes from V_1 to V_2.
- If the displacement of the piston is kept very small, the pressure of the gas will not change much and can be considered constant.
- The work done by the gas which expands at constant pressure is:

$$\Delta W = P(V_2 - V_1)$$

$$\Delta W = P\Delta V$$

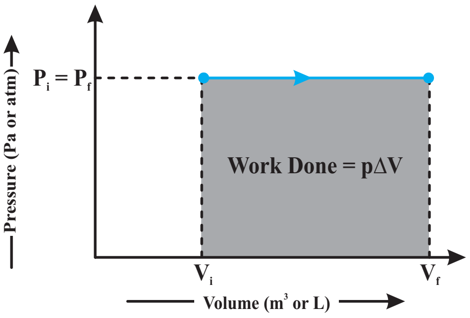


Figure 16.8 (b)
work done by change in volume of gas

Using the first law of thermodynamics equation

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

$$\Delta Q = \Delta U + P\Delta V \; \ldots\ldots 16.9$$

The work performed by the expanding or contracting gas comes from one or both sources:

- heat supplied to the gas and
- the internal energy of the gas

The graph of isobaric process is called an "isobar", which is a straight line, parallel to the volume axis.

(c) Isothermal Process:

"The thermodynamics process which is carried out in such a way that a system undergoes changes but its temperature remains constant is called an isothermal process".

The internal energy of an ideal gas does not depend on the volume but depends only on the temperature of the gas. Consider a gas contained in a cylinder having a conducting base and non-conducting walls and with a movable piston at the end.

The base of the cylinder is placed on a heat reservoir at temperature T_1.Let the gas be then allowed to expand slowly by decreasing the pressure on the piston. Due to this expansion, the gas tends to cool down. But heat is conducted from the heat reservoir to the gas so that the temperature of the system remains constant and is equal to the temperature of the reservoir. Pressure changes from P_1 to P_2, volume changes from V_1 to V_2 but temperature remains constant (T).

During isothermal expansion some work is done by the gas in pushing up the piston in the cylinder. Since the temperature remains constant, there is no change in the internal energy of the gas, that is,$\Delta U = 0$

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

$$\Delta Q = 0 + \Delta W$$

$$\Delta Q = \Delta W \; \ldots\ldots\ldots 16.10$$

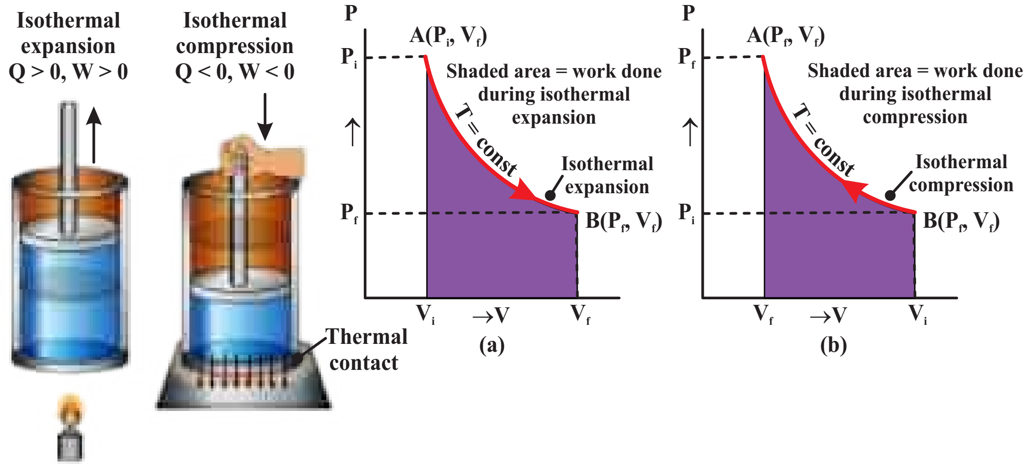


Figure 16.9 Isothermal Expansion, Compression and their graph (a, b)

This shows that:

If the gas expands and does external work, an equal amount of heat has to be supplied in order to maintain its temperature constant.

Conversely, if the gas contracts, work is being done on it and equal amount of heat has to be allowed to leave the gas. In the PV diagram the work done during the isothermal expansion is equal to the area under the graph as shown in Figure 16.9

Similarly, for an isothermal compression, the area under the PV graph is equal to the work done on the gas which turns out to be the area with a negative sign.

The whole process from an initial state to a final state is represented by a continuous curve which is called an "isotherm".

(d) Adiabatic Process:

"The thermodynamics process during which no heat enters or leaves a system is called an adiabatic process".

For every adiabatic process **$\Delta Q = 0$**, A truly adiabatic process is an ideal one which cannot be realized. However, the flow of heat may be prevented either by surrounding the system with a thick layer of heating insulating material such as cork, asbestos, or by performing the process very quickly. The flow of heat requires finite time, so any process performed quickly enough will be practically adiabatic. To perform an adiabatic process on a gas, we consider the gas to be contained in a completely insulated cylinder with a movable piston at the end.

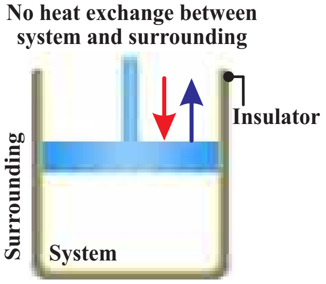


Figure 16.10 Adiabatic Process

Adiabatic Expansion:

If the gas is allowed to expand suddenly decreasing the pressure on the piston, the gas will be cooled, so that its temperature will decrease. Therefore, the internal energy of the gas will decrease.

Adiabatic Compression:

If the gas is compressed by suddenly increasing the pressure on the piston, the gas will be heated, so that temperature will increase. Therefore, the internal energy of the gas will increase.

According to the first law of thermodynamics;

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

$$0 = \Delta U + \Delta W$$

$$\Delta U = -\Delta W$$

Thus, an increase in the internal energy of the system in an adiabatic process is equal to the work done on the system.

Above equation can be written as;

$$\Delta W = -\Delta U \ldots\ldots\ldots .. 16.11$$

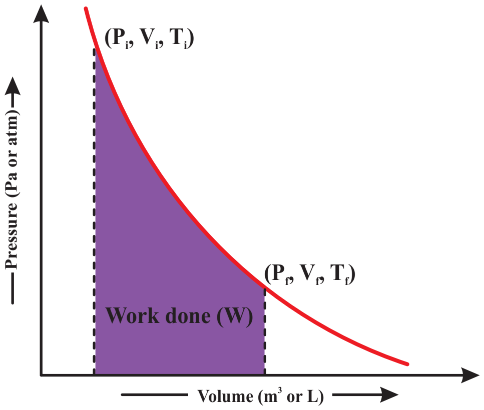


Figure 16.11 graph for adiabatic Compression

This means that if the system does the work, then in adiabatic process, the work is done at the cost of internal energy.

Self-Assessment Questions:

1. What is the first law of thermodynamics and how is it expressed mathematically?
2. How does the first law of thermodynamics relate to the conservation of energy?

Worked Example 16.3

The normal human body temperature is 36.88°C. What is this temperature on Fahrenheit scale? A gas confined in a cylinder undergoes an expansion from an initial volume of 2.0 L to a final volume of 4.0 L against a constant external pressure of 2.0 atm. Calculate the work done by the gas during this process.

Solution:

Step 1: Write the known quantities and point out quantities to be found.

Initial volume V_i= 2.0 L
Final volume V_f= 4.0 L
Constant external pressure (P) = 2.0 atm

Step 2: Write the formula and rearrange if necessary.

The work done by a gas against a constant external pressure can be calculated using the formula:

$$W = -P\Delta V$$

where: W is the work done,
P is the constant external pressure, and
ΔV is the change in volume.

Step 3: First, calculate ΔV:

$$\Delta V = V_f - V_i$$

$$\Delta V = 4.0L - 2.0L = 2.0L$$

Step 4: Now, substitute the values into the formula:

$$W = -P\Delta V$$
$$W = -(2.0atm) \times (2.0L)$$
$$W = -4.0atm \cdot L$$

Result: The negative sign indicates that work is done on the gas (compression) rather than by the gas (expansion). The final result is −4.0atm·L of work done on the gas during this expansion process.

Remember that the unit of work in the International System of Units (SI) is the joule (J), so you may want to convert the result to joules if needed (1 atm·L = 101.325 J).

Self-Assessment Questions:

1. How does an adiabatic process differ from an isothermal process in thermodynamics?
2. Why is an isothermal process important in thermodynamics, and how does it affect the internal energy of a system?
3. Why is an isobaric process significant in thermodynamics, and how does it affect the volume of a gas?

16.4.1 Heat capacity:

When heat is absorbed by a body, its temperature increases, and when heat is lost, its temperature decreases. Temperature represents the total kinetic energy of the particles comprising the body. Therefore, when heat is absorbed, it increases the kinetic energy of these particles, thereby raising the temperature. Consequently, the change in temperature ΔT is directly proportional to the amount of heat transferred q.

The formula $Q = nC\Delta T$ expresses the relationship where q is the heat transferred, n is the amount of substance (often expressed in moles), C is the molar heat capacity, and ΔT is the change in temperature. Molar heat capacity C refers to the amount of heat energy required to change the temperature of 1 mole of a substance by 1 unit. It is influenced by the substance's properties such as its nature, size, and composition.

Specific Heat:

Specific heat, often denoted by C orc, is a measure of the heat energy required to raise the temperature of a unit mass (usually one gram or one kilogram) of a substance by one degree Celsius (or one Kelvin). The formula for specific heat (c) is given by:

$$\boldsymbol{Q = mc\Delta T \ldots\ldots 16.12}$$

where:

Q is the heat energy absorbed or released,
m is the mass of the substance,
c is the specific heat, and
ΔT is the change in temperature.

Specific heat is an intensive property, meaning it does not depend on the quantity of the substance.

Molar Specific Heat:

Molar specific heat, denoted as C_m, measures how much heat energy is needed to raise the temperature of one mole of a substance by one degree Celsius or Kelvin. It's calculated using the formula:

$$Q = nC_m\Delta T \ldots\ldots 16.13$$

where:

Q is the heat transferred,
n is the number of moles,
C_m is the molar specific heat,
ΔT is the temperature change.

Molar specific heat is an intensive property and relates to specific heat C through the equation C_m= nC

In thermodynamics, both specific heat and molar specific heat are vital for quantifying how heat energy changes with temperature in substances. These values vary depending on the substance, reflecting each substance's unique ability to absorb or release heat.

There are two types of molar heat capacity:

C_P (molar heat capacity at constant pressure)
C_V (molar heat capacity at constant volume)

The relationship between CP and CV for an Ideal Gas:

From the equation $q = n\,C\,\Delta T$, we can derive:

At constant pressure $q_p = n\,C_p\Delta T$ This value corresponds to the change in enthalpy, expressed as: $q_P = n\,C_P\Delta T = \Delta H$

Similarly, at constant volume $Q_v = n\,C_v\Delta T$

This value corresponds to the change in internal energy, expressed as:

$$Q_v = n\,C_v\,\Delta T = \Delta U$$

For one mole of an ideal gas (n=1)

$$\Delta H = \Delta U + \Delta(PV) = \Delta U + \Delta(RT) = \Delta U + R\,\Delta T \ldots\ldots 16.14$$

Therefore, we have: $\Delta H = \Delta U + R\,\Delta T$

Substituting the expressions for

$$C_p\Delta T = C_v\,\Delta T + R\,\Delta T$$

$$C_p = C_p - C_v = R \ldots\ldots 16.15$$

This equation shows the relationship between the molar heat capacities at constant pressure C_P and constant volume C_V for an ideal gas, where R is the universal gas constant.

Worked Example 16.4

Calculate the molar specific heat capacity of a gas when 2 moles of the gas absorb 1500 J of heat energy, and its temperature increases by 25 degrees Celsius.

Solution:

Step1: write the known quantities and point out quantities to be found

Q = 1500J, n = 2mol, and ΔT = 25°C.

Step 2: Write the formula and rearrange if necessary

The heat transfer (Q) is given by:

$Q = n\cdot Cm\cdot\Delta T$

Let's rearrange the formula to solve for Cm:

$Cm = n\cdot\Delta TQ$

Step 3: Now, substitute the values into the formula:

$Cm = 1500J/2mol\cdot 25°C$

$Cm = 30J/mol\cdot°C$

Therefore, the molar specific heat capacity of the gas is 30J/mol·°C.

Result: 30J/mol·°C.

Worked Example 16.5

A sample of water with a mass of 200 grams is heated, and its temperature rises from 25°C to 45°C. Calculate the heat transferred to the water. Given the specific heat capacity of water (c) is 4.18 J/g°C.

Solution: Write the known quantities and point out quantities to be found

$m=200g$

$c=4.18J/g°C$

Initial temperature (T1) = 25°C

Final temperature (T2) = 45°C

Calculate the temperature change (ΔT): $\Delta T=T2-T1=45°C-25°C=20°C$

Step 2: Write the formula and rearrange if necessary

heat transfer equation $Q=mc\Delta T$ to calculate the heat transferred,

Step 3: Now, substitute the values into the formula:

Substitute the values: $Q=(200g)\times(4.18J/g°C)\times(20°C)$

Calculate the result: $Q=8360J$

Result: The heat transferred to the water is 8360 joules.

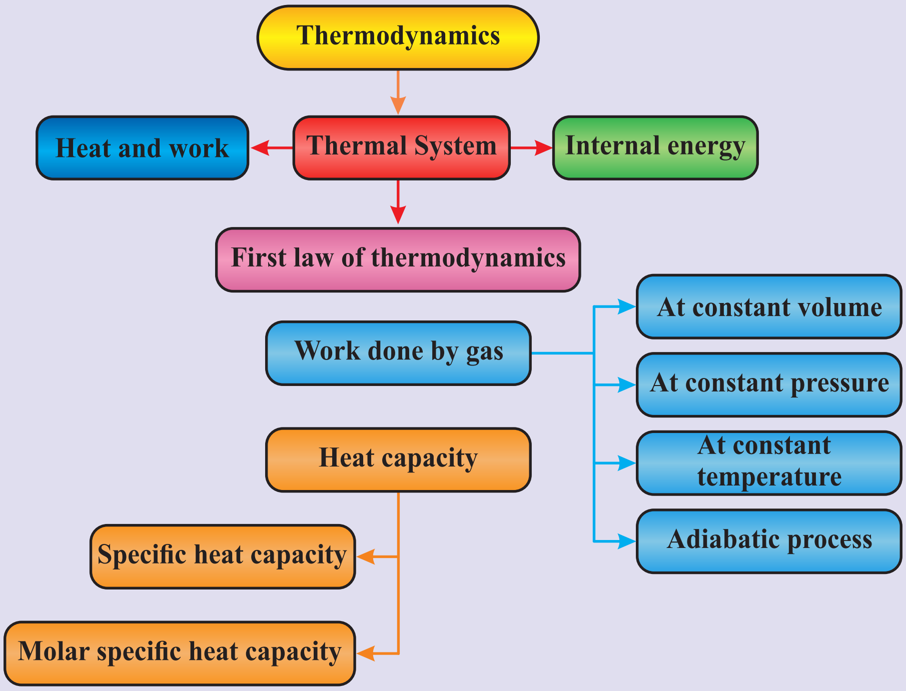
Self-Assessment Questions:

1. How does heat capacity relate to the amount of energy required to change the temperature of a substance?
2. Explain why different materials have different heat capacities.
3. How does the heat capacity of a substance change with temperature?

SUMMARY

- ✓ Heat (Q): Transfer of thermal energy between a system and its surroundings.
- ✓ Work (W): Mechanical energy transferred to or from a system through forces acting on it.
- ✓ Sign conventions: Heat added to the system and work done by the system are positive, while heat removed from the system and work done on the system are negative.
- ✓ Equation: $\Delta U = Q - W$, where ΔU is change in internal energy, Q is heat added to the system, and W is work done by the system.
- ✓ Internal energy (U): Total energy of a system, including kinetic and potential energies.
- ✓ First Law of Thermodynamics: Conservation of energy principle in thermodynamic systems.
- ✓ Energy cannot be created or destroyed, only transferred or converted from one form to another.
- ✓ Applications: Understanding energy transformations in various processes such as heating, cooling, phase changes, and chemical reactions.

CONCEPT MAP



EXERCISE

Section (A): Multiple Choice Questions (MCQs)

Choose the correct answer:

1. What type of process occurs when a system exchanges both heat and work with its surroundings, and there is no change in internal energy?
   (a) Isothermal process (b) Adiabatic process
   (c) Isobaric process (d) Isochoric process
2. During an isobaric process, what remains constant?
   (a) Temperature (b) Pressure (c) Volume (d) Internal energy
3. In which thermodynamic process does a system exchange heat with its surroundings but undergoes no change in temperature?
   (a) Isothermal process (b) Adiabatic process
   (c) Isobaric process (d) Isochoric process
4. What is the characteristic of an adiabatic process?
   (a) Constant pressure (b) No heat exchange with the surroundings
   (c) Constant temperature (d) Constant volume
5. In an isochoric process, what is the primary feature?
   (a) Constant temperature (b) No work done
   (c) Constant volume (d) No heat exchange
6. What is internal energy in a thermodynamic system?
   (a) The energy associated with motion
   (b) The energy associated with the system's position
   (c) The sum of kinetic and potential energy
   (d) The total energy contained within the system
7. How is the change in internal energy (ΔU) defined in terms of heat (ΔQ and work (ΔW)
   (a) $\Delta U = \Delta Q + \Delta W$ (b) $\Delta U = \Delta Q - \Delta W$
   (c) $\Delta U = \Delta Q$ times W (d) $\Delta U = \Delta Q / \Delta W$
8. What is the internal energy of an ideal gas related to?
   (a) Temperature only (b) Pressure only
   (c) Volume only (d) Temperature, pressure, and volume
9. During an adiabatic process, what happens to the internal energy of the system?
   (a) Increases (b) Decreases
   (c) Remains constant (d) Depends on the specific heat
10. What is the equation for the internal energy change of a system in an isochoric process?
    (a) $\Delta U = \Delta Q + \Delta W$ (b) $\Delta U = \Delta Q - \Delta W$
    (c) $\Delta U = \Delta Q$ times ΔW (d) $\Delta U = \Delta Q / \Delta W$

Section (B): CRQs (Short Answered Questions):

1. Explain the concept of the first law of thermodynamics in your own words.
2. How does the first law of thermodynamics relate to the conservation of energy?

3. Distinguish between the work done by a system and the heat exchanged with the surroundings in the context of the first law.
4. Give an example daily life that illustrates the principles of the first law of thermodynamics.
5. Explain the role of the system and its surroundings in the context of the first law of thermodynamics.
6. How does heat capacity relate to the amount of energy required to change the temperature of a substance?

Section (C): ERQs (Long Answered Questions):

1. Provide the mathematical expression of the first law of thermodynamics and explain each term.
2. Describe what happens to the internal energy of a system in an adiabatic process, and why.
3. What is meant by the internal energy of a system, and how does it change during various thermodynamic processes?
4. Explore the concept of positive and negative work done by a gas, depending on whether the gas is expanding or compressing.
5. Explain the concept of heat capacity and its significance in thermodynamics. Discuss how heat capacity relates to the ability of a substance to store thermal energy and its implications for temperature changes.

Section (D): Numerical:

1. A gas undergoes isothermal expansion at a constant temperature of 300 K. If the gas absorbs 500 J of heat during the process, calculate the work done by the gas. **(500 J)**
2. A piston compresses a gas adiabatically. If the initial volume is 0.02 m^3 and the final volume is 0.01 m^3, and the initial pressure is 200 kPa, determine the final pressure. Assume the gas behaves ideally. **(400 kPa)**
3. A system undergoes an isobaric process where the pressure is kept constant at 150 kPa. If the volume increases from 0.05 m^3 to 0.08 m^3,calculate the heat added to the system. **(600 J)**
4. During an isochoric process, the internal energy of a gas increases by 300 J. If no work is done, determine the heat added to the system. **(300 J)**
5. A gas undergoes a cyclic process, starting at point A with a volume of 0.02 m^3, going to B (isochoric heating), then to C (isothermal expansion), and finally back to A. If the heat added during isothermal expansion is 1000 J and the heat rejected during isochoric heating is 500 J, calculate the net work done by the system. **(500 J)**
6. A gas expands from 0.03 m^3 to 0.06 m^3 against a constant pressure of 100 kPa. Calculate the work done in both a reversible and an irreversible process, and compare the results. **(300 J and 600 J)**
7. A 50 g piece of copper at 100°C is placed in 200 g of water at 20°C. If the final temperature of the system is 30°C, calculate the specific heat capacity of copper. (Specific heat capacity of water = 4.18 J/g°C) **(0.39 J/g°C)**
8. How much heat is required to raise the temperature of 1 kg of lead from 25°C to 100°C? (Specific heat capacity of lead = 0.128 J/g°C). **(0240J).**