CLASS IX - CBSE

Physics Study Material

Name:

1 Introduction to Thermodynamics

In the previous chapter, we studied the **thermal properties of matter**. In this chapter, we focus on the **laws that govern thermal energy** and on the processes in which **work is converted into heat and heat into work**.

Everyday Observations

- In winter, when we rub our palms together, the work done against friction produces heat and our hands feel warmer.
- In a steam engine, the heat of the steam is used to do mechanical work on the pistons, which ultimately move the wheels of a train.

These simple examples suggest a deep connection between heat and mechanical work.

To understand this connection, physics needs clear definitions of **heat**, **temperature** and **work**. Historically, the correct idea of heat took a long time to develop.

Historical Idea of Heat

Earlier, heat was thought to be a kind of invisible fluid called *caloric* that filled the pores of a substance. When a hot body was placed in contact with a cold body, caloric was supposed to flow from one to the other until their "caloric levels" (temperatures) became equal, somewhat like water flowing between two connected tanks until their levels are the same.

Definition of Thermodynamics

Thermodynamics is the branch of physics that deals with **heat**, **temperature** and the **interconversion of heat and other forms of energy**. It treats matter in terms of bulk (macroscopic) properties and does not require knowledge of the molecular constitution of matter.

A thermodynamic description of a system uses a small number of **macroscopic variables**, such as

pressure, volume, temperature, mass and composition,

which can be measured directly in the laboratory. It does not track the positions and velocities of individual molecules, unlike a microscopic or kinetic theory description.

Thermodynamics vs Mechanics

- **Mechanics** studies the motion of particles or bodies under forces and torques, focusing on quantities like position, velocity and kinetic energy.
- **Thermodynamics** is concerned with the *internal macroscopic state* of a body, described by variables like pressure, volume and temperature.

For example, when a bullet is fired, its mechanical state (especially kinetic energy) changes. When it comes to rest inside a block of wood, its kinetic energy is converted into heat, raising the temperature of the bullet and the surrounding wood.

2 Thermal Equilibrium

In everyday life, we observe that when a hot object is placed in contact with a cold object, the hot one cools down and the cold one warms up. After some time, both reach the *same* temperature and then no further change in temperature is observed. At that stage, we say that the two objects are in **thermal equilibrium**.

2.1 Thermal contact and heat flow

Thermal Contact and Direction of Heat Flow

Two systems are said to be in **thermal contact** if they are allowed to exchange heat with each other but not necessarily matter. As long as their temperatures are different, there is a net transfer of energy (heat) from the **higher temperature** system to the **lower temperature** system.

- Heat flows *spontaneously* from a hotter body to a colder body.
- This heat flow continues until the temperatures of the two bodies become equal.

Once the temperatures become equal, the macroscopic state no longer changes with time and there is **no net heat flow** between them, even though microscopic molecular motion continues.

2.2 Definition of thermal equilibrium

Thermal equilibrium between two systems may now be defined as follows:

> Two systems are in **thermal equilibrium** if, when they are in thermal contact, their

temperatures remain equal and there is **no net transfer of heat** between them.

Similarly, a system is in thermal equilibrium with its surroundings if its temperature is the same as that of the surroundings and does not change with time.

2.3 Illustrative examples

1. **Hot tea in a room:** A cup of hot tea initially has a temperature higher than the surrounding air. Heat flows from the tea to the air, so the tea cools down and the air warms up slightly. After some time, the tea and the air reach the same temperature (room temperature). At this stage, the tea and the air are approximately in thermal equilibrium and there is no further net heat flow between them.

2. **Water in a refrigerator:** When a bottle of water at room temperature is placed in a refrigerator, the water is initially warmer than the cold air inside. Heat flows from the water to the cold air and to the cooling coils. Eventually, the water cools to the refrigerator temperature. Then the water and the air inside the fridge are in thermal equilibrium.

3. **Metal block and thermometer:** Consider a metal block whose temperature we want to measure. A thermometer is put in contact with the block. Initially, the thermometer reading changes as heat is exchanged between the block and the thermometer. When the reading becomes steady and stops changing, the thermometer and the block have reached the same temperature. They are now in thermal equilibrium, and the thermometer reading correctly represents the temperature of the block.

These examples show that thermal equilibrium is always associated with:

- Equal temperatures of the systems in contact.
- Absence of any further macroscopic change in temperature with time.

2.4 Thermal equilibrium and macroscopic description

In thermodynamics, a system in **equilibrium** (more precisely, in *thermodynamic equilibrium*) has macroscopic variables such as pressure, volume and temperature that do not change with time. Thermal equilibrium is specifically concerned with the temperature part of this description.

When we say that a system has a well-defined temperature, we are implicitly assuming that it is in thermal equilibrium, either with itself (all parts at the same temperature) or with its surroundings. Only in such a state can we assign a single temperature to the entire system and use a thermometer to measure it reliably.

2.5 Preview: Connection with the Zeroth Law

The intuitive idea of thermal equilibrium is made precise by the **Zeroth Law of Thermodynamics**, which will be discussed in the next section. In simple words, it states that:

> If system A is in thermal equilibrium with system C, and system B is also in thermal equilibrium with system C, then A and B are in thermal equilibrium with each other.

This law provides the logical foundation for the concept of temperature and for using thermometers to compare temperatures of different systems.

3 Thermal Equilibrium and Types of Walls

In general, whether or not a system is in a state of equilibrium depends on the **surroundings** and on the **nature of the wall** that separates the system from the surroundings. A system is in equilibrium only when its macroscopic variables (such as pressure, volume and temperature) do not change with time.

3.1 Two gases separated by an adiabatic wall

Consider two gases, A and B, occupying two different containers. For a given mass of gas, we know that pressure and volume can be chosen as two independent variables. Let the states of the gases be

$$(P_A, V_A) \quad \text{and} \quad (P_B, V_B)$$

respectively.

Adiabatic Wall

An **adiabatic wall** (thermally insulating wall) is a wall that does *not* allow heat (energy as heat) to pass from one system to another. It may be movable (so work can sometimes be done), but there is no transfer of heat across it.

Suppose the two containers are placed next to each other but separated by an adiabatic wall. The systems

are also insulated from the rest of the surroundings by similar adiabatic walls. In this situation:

- No heat can flow between gas A and gas B.
- Each gas can remain in its own equilibrium state independently.
- Any possible pair of values (P_A, V_A) can be in equilibrium with any possible pair (P_B, V_B).

Since there is no thermal contact, there is no requirement that the temperatures of A and B be related in any particular way.

3.2 Replacing the adiabatic wall with a diathermic wall

Now suppose that the adiabatic wall between A and B is replaced by a **diathermic wall** (also called a conducting wall).

Diathermic Wall

A **diathermic wall** is a wall that allows heat to flow between the two systems in contact. It prevents mixing of matter but permits energy transfer in the form of heat.

Once the diathermic wall is introduced:

- If gases A and B are initially at different temperatures, heat flows spontaneously from the higher-temperature gas to the lower-temperature gas.
- Because of this heat exchange, the macroscopic variables of each gas change:

$$(P_A, V_A) \to (P'_A, V'_A), \qquad (P_B, V_B) \to (P'_B, V'_B).$$

- This process continues until no further macroscopic change occurs in either gas.

Thermal Equilibrium Between Two Systems

When the macroscopic variables of both systems stop changing and there is **no net flow of heat** between them, the two systems are said to be in **thermal equilibrium**. At this stage, their temperatures are equal.

In the final situation with the diathermic wall:

- Gas A is in a new equilibrium state (P'_A, V'_A).
- Gas B is in a new equilibrium state (P'_B, V'_B).
- There is no more net heat transfer between A and B.

Therefore, we say that **system A is in thermal equilibrium with system B**. This example clearly shows how the nature of the wall (adiabatic or diathermic) determines whether the systems can exchange heat and whether they can come to thermal equilibrium.

4 Zeroth Law of Thermodynamics

The concept of **thermal equilibrium** leads naturally to an important principle called the **Zeroth Law of Thermodynamics**. This law gives a precise logical basis for the idea of temperature and for the use of thermometers.

4.1 Thought experiment with three systems

Consider three systems, A, B and C:

- Systems A and C are in thermal contact with each other through a conducting (diathermic) wall.
- Systems B and C are also in thermal contact through another conducting wall.
- Systems A and B are separated from each other by an **adiabatic** (insulating) wall.

In this arrangement:

- Heat can flow between A and C, and between B and C.
- No heat can flow directly between A and B because of the adiabatic wall.

The macroscopic variables (such as pressure, volume and temperature) of A and B will change with time due to heat exchange with C, until both A and B each come to **thermal equilibrium** with C. After some time:

A is in thermal equilibrium with C, B is in thermal e

Now, in the second step of the experiment:

- Replace the adiabatic wall between A and B by a conducting wall.
- At the same time, insulate system C from both A and B using an adiabatic wall.

A key experimental observation is:

- After this change, there is *no further change* in the macroscopic states of A and B.
- No net heat flows between A and B, even though they are now in thermal contact.

This means that A and B are already in **thermal equilibrium** with each other.

4.2 Statement of the Zeroth Law

This observation is summarised in the form of the Zeroth Law:

Zeroth Law of Thermodynamics

If two systems are each in thermal equilibrium with a third system, then they are in thermal equilibrium with each other.

In symbols, if

A is in thermal equilibrium with C and B is in thermal equilibrium with C,

then

$$A \text{ is in thermal equilibrium with } B.$$

4.3 Importance of the Zeroth Law

The Zeroth Law looks simple, but it has very important consequences:

- It allows us to introduce **temperature** as a property that is the same for all systems in mutual thermal equilibrium.
- It justifies the use of a third system (a **thermometer**) to compare temperatures:
  - If a thermometer is in thermal equilibrium with system A, it shows the temperature of A.
  - If the same thermometer is then in thermal equilibrium with system B and gives the same reading, we conclude that A and B are at the same temperature and hence in thermal equilibrium with each other.

The Zeroth Law clearly suggests that when two systems, say A and B, are in thermal equilibrium with each other, there must exist a physical quantity that has the *same value* for both systems. This thermodynamic variable, whose value is equal for two systems in thermal equilibrium, is called the **temperature**, usually denoted by T.

Temperature as a Thermodynamic Variable

If systems A and B are each in thermal equilibrium with a third system C, then

$$T_A = T_C \quad \text{and} \quad T_B = T_C.$$

From these equalities, it follows that

$$T_A = T_B,$$

which means that A and B are at the same temperature and hence are in thermal equilibrium with each other.

Thus, temperature is the thermodynamic variable that characterises thermal equilibrium: all systems in mutual thermal equilibrium share the same temperature.

5 Heat, Internal Energy and Work

The Zeroth Law of Thermodynamics led us to the concept of **temperature**, which agrees with our everyday idea of how hot or cold a body is. Temperature is a measure of the "hotness" of a body and plays a central role in determining how heat flows.

When two bodies at different temperatures are placed in thermal contact:

- Heat flows spontaneously from the body at **higher temperature** to the body at **lower temperature**.
- This flow of heat continues until the temperatures of the two bodies become equal.

Once their temperatures are equal, there is no further net flow of heat between them and the two bodies are said to be in **thermal equilibrium**. Temperature, therefore, tells us both:

- The sense of "how hot" or "how cold" a body is.
- The *direction* in which heat will flow when two bodies are brought into contact.

In earlier discussions, we have seen how to construct temperature scales (such as Celsius and Kelvin) in order to assign numerical values of temperature to different bodies. Having established temperature as a thermodynamic variable, we now move on to clarify three closely related concepts:

- **Heat** – energy transferred between systems *because* of a temperature difference.

- **Internal energy** – the total energy associated with the microscopic constituents (molecules/atoms) of a system.
- **Work** – energy transferred to or from a system through macroscopic mechanical means (such as expansion against an external pressure, lifting a weight, etc.).

In the following sections, we will define these quantities more precisely and study how they are related by the laws of thermodynamics, especially the **First Law of Thermodynamics**.

6 Internal Energy

Every substance (gas, liquid or solid) is made of a huge number of tiny particles called molecules or atoms. These particles are always in motion and also interact with each other.

What is internal energy?

- Each molecule has **kinetic energy** because it is moving.
- Molecules also have **potential energy** because of forces between them (attraction or repulsion).

If we add up:

- the kinetic energy of all molecules, and
- the potential energy of all molecules,

we get the **internal energy** of the system.

Symbol: Internal energy is denoted by U.

Very important points:

- Internal energy is the energy *inside* the system due to the random motion and interactions of its molecules.
- It does *not* include the kinetic energy of the whole body moving as one piece. (Example: If a gas cylinder is moving in a truck, the energy of motion of the whole cylinder is not counted in U.)

Internal energy as a state variable

In thermodynamics, we treat internal energy as a quantity that depends only on the **state** of the system.

Internal Energy is a State Variable

For a given amount of gas, the internal energy U depends only on its state (for example, on its pressure, volume and temperature). It does *not* depend on how that state was reached.

This means:

- If a gas ends up in the same state (same P, V, T), then its internal energy U is the same, no matter what process or path it followed.
- It does not matter whether you compressed the gas quickly or slowly, or heated it first and then compressed it. If the final state is the same, U is the same.

Thus, **pressure** P, **volume** V, **temperature** T and **internal energy** U are all examples of **thermodynamic state variables**. They describe the condition (state) of the system at that moment.

7 How Can We Change the Internal Energy of a System?

We have seen that the internal energy U of a system (for example, a gas in a cylinder) is the total microscopic energy of its molecules. Now we ask a practical question:

In what ways can the internal energy of a system be changed?

To keep things simple, imagine a fixed mass of gas enclosed in a cylinder fitted with a movable piston.

Observation from experiments: There are two basic ways to change the state of the gas, and therefore to change its internal energy:

1. **By heating (transfer of heat)**
2. **By doing work**

7.1 Changing internal energy by heating

Place the cylinder in contact with a body (for example, a hot reservoir) at a **higher temperature** than the gas.

- Because of the temperature difference, energy flows from the hotter body to the gas. This energy transfer due to temperature difference is called **heat**.
- As heat flows into the gas, the internal energy U of the gas **increases**. (Usually its temperature also rises.)

If instead the surroundings are at a **lower temperature** than the gas:

- Heat flows from the gas to the surroundings.
- The gas **loses** internal energy.

7.2 Changing internal energy by doing work

Now consider mechanical action on the piston:

- If an external agent **pushes the piston down**, it compresses the gas.
- In this process, **work is done on the gas** by the external agent.
- As a result, the internal energy U of the gas **increases**. (Again, the temperature usually increases.)

The reverse can also happen:

- If the gas **pushes the piston up** and makes it move, the gas is doing work on the surroundings.
- In doing this work, the gas uses up some of its internal energy, so its internal energy U **decreases**.

Key idea

Heat and work are two different modes of transferring energy between a system and its surroundings. Both can change the internal energy of the system:

- Heat: energy transfer due to temperature difference.
- Work: energy transfer by macroscopic mechanical means (like moving a piston).

In the next part, this idea will be made precise in the form of the **First Law of Thermodynamics**, which relates the change in internal energy to the heat supplied and the work done.

Heat, Work and Internal Energy

We have defined the internal energy U of a system as a state variable (a property of the state of the system). Now we clarify the roles of **heat** and **work**.

In thermodynamics:

- **Internal energy** U is the energy stored inside the system due to the random motion and interactions of its molecules.
- **Heat** Q is energy transferred between the system and surroundings *because of a temperature difference.*
- **Work** W is energy transferred between the system and surroundings by macroscopic mechanical means (for example, pushing a piston, moving a weight, electrical work, etc.).

Important: Heat and work are not properties that the system "contains". They are **modes of energy transfer**. They appear only during a process when the system interacts with its surroundings and its internal energy changes.

In everyday language, people often say that a hot body "contains a lot of heat". In thermodynamics, this is not correct. The body contains internal energy; *heat* is the energy that flows *into or out of* the body because of a temperature difference.

8 State Variables and Path Variables

The distinction above is related to an important idea in thermodynamics: the difference between **state variables** and **path variables**.

8.1 State variables (state functions)

- A **state variable** is a quantity whose value depends only on the current **state** of the system, not on how the system reached that state.
- Examples: pressure P, volume V, temperature T, internal energy U (and later, enthalpy H, entropy S, etc.).

If a gas changes from state 1 to state 2, the change in any state variable X is

$$\Delta X = X_2 - X_1,$$

and this ΔX is the same for *all* processes that take the system from state 1 to state 2.

8.2 Path variables (process-dependent quantities)

- A **path variable** (or path function) is a quantity whose value depends on the *actual path* (process) taken by the system between two states.
- In thermodynamics, **heat** Q and **work** W are path variables.

If a gas goes from state 1 to state 2 by different processes (for example, isothermal, adiabatic, or any other path on a P-V diagram), then in general

$$Q^{(A)}_{1\to2} \neq Q^{(B)}_{1\to2}, \qquad W^{(A)}_{1\to2} \neq W^{(B)}_{1\to2},$$

even though the initial and final states are the same.

Key Point

State variables (like P, V, T, U) depend only on the state of the system and are path-independent. **Path variables** (like heat Q and work W) depend on the specific process taken between two states and are path-dependent.

This is why, in the First Law of Thermodynamics, we write

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

Here ΔU is a change in a state variable (depends only on initial and final states), while Q and W are path variables that depend on how the process is carried out between those states.

9 First Law of Thermodynamics

We have seen that the internal energy U of a system can be changed in two ways:

- by supplying or removing **heat** Q,
- by doing **work** W on the system or by the system.

The First Law of Thermodynamics gives a precise relationship between these quantities.

9.1 Statement of the First Law

First Law of Thermodynamics (basic idea): The change in internal energy of a system is equal to the heat supplied to the system minus the work done by the system.

$$\Delta U = Q - W$$

Here:

- ΔU is the change in internal energy of the system.
- Q is the heat supplied *to* the system (taken positive if energy enters as heat).
- W is the work done *by* the system on the surroundings (taken positive if the system does work).

This is essentially a statement of the **law of conservation of energy** applied to thermodynamic processes.

9.2 Physical meaning

Think of the system as a bank account of internal energy U.

- When you **add heat** Q to the system, it is like depositing energy into the account.
- When the system **does work** W (for example, by pushing a piston), it is like withdrawing energy from the account.

What remains as change in the "balance" is the change in internal energy ΔU:

Change in internal energy = Energy added as heat−Ener

Sign Convention for the First Law

In this textbook convention, we use:

- $Q > 0$: heat supplied *to* the system,
- $Q < 0$: heat rejected *by* the system,
- $W > 0$: work done *by* the system (on the surroundings),
- $W < 0$: work done *on* the system (by the surroundings).

With this convention, the First Law is written as:

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

Some Useful Special Cases of the First Law

- **Adiabatic process** (no heat exchange, $Q = 0$):

  $$\Delta U = -W.$$

  If the system does work ($W > 0$), its internal energy decreases ($\Delta U < 0$).

- **Isothermal process** for an ideal gas ($\Delta U = 0$ because U depends only on temperature):

  $$0 = Q - W \quad \Rightarrow \quad Q = W.$$

  The heat absorbed by the gas is exactly equal to the work done by the gas.

- **Constant volume process** ($W = 0$, since no volume change):

  $$\Delta U = Q.$$

  All the heat supplied goes into changing internal energy.

Key idea to remember

- The First Law does **not** tell us in which direction a process will occur; it only accounts for energy.
- It tells us that energy is always conserved: energy can be transferred as heat or work, and can change form, but the total energy of system + surroundings remains constant.

For problem solving in PUC:

1. Identify the system.
2. Decide the signs of Q and W using the convention.
3. Use $\Delta U = Q - W$ along with the equation of state (like $PV = nRT$) and the nature of the process (isothermal, adiabatic, etc.).

Alternative Form of the First Law

From the First Law in our sign convention,

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

where Q is heat supplied to the system and W is work done by the system.

We can also write this as

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

For a finite change, using the symbol ΔQ for the total heat supplied and ΔW for the total work done by the system, we may write:

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

Interpretation: The energy supplied as heat to the system is partly used to increase the internal energy and partly used to do work by the system.

10 Derivation of $\mathrm{d}W = P\,\mathrm{d}V$

Consider a gas enclosed in a cylinder fitted with a movable piston of cross-sectional area A.

Step 1: Force due to gas pressure

The gas exerts pressure P on the piston. The force on the piston is

$$F = PA.$$

Step 2: Small displacement and work done

If the piston moves outwards by a small distance $\mathrm{d}x$, the work done by the gas on the piston is

$$\mathrm{d}W = F\,\mathrm{d}x.$$

Substituting $F = PA$,

$$\mathrm{d}W = PA\,\mathrm{d}x.$$

Step 3: Relating displacement to change in volume

The small increase in volume of the gas when the piston moves by $\mathrm{d}x$ is

$$\mathrm{d}V = A\,\mathrm{d}x.$$

Therefore,

$$\mathrm{d}W = P\,\mathrm{d}V.$$

Work Done in a Quasi-static Volume Change

For a small, quasi-static change in volume,

$$\boxed{\mathrm{d}W = P\,\mathrm{d}V.}$$

For a finite process from volume V_1 to V_2,

$$W = \int_{V_1}^{V_2} P\,\mathrm{d}V.$$

If the pressure remains constant (isobaric process),

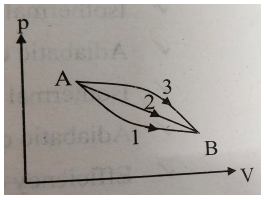
$$W = P(V_2 - V_1) = P\,\Delta V.$$

D.P.P - I (Classwork)

1. A thermodynamic system goes from state (p, V) to $(2p, V)$ and another from (p, V) to $(p, 2V)$. Work done in the two cases is

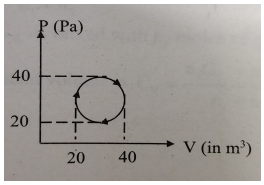
   (A) zero, zero
   (B) zero, pV
   (C) pV, zero
   (D) pV, pV

2. A given mass of a gas expands from state A to state B along three paths 1, 2 and 3 as shown. If W_1, W_2 and W_3 are the work done along paths 1, 2 and 3 respectively, then:



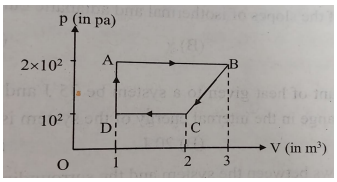
   (A) $W_1 > W_2 > W_3$
   (B) $W_1 < W_2 < W_3$
   (C) $W_1 = W_2 = W_3$
   (D) $W_2 > W_1 > W_3$

3. The heat energy absorbed during the cyclic process shown is:



   (A) $10^7\pi$ J
   (B) $10^2\pi$ J
   (C) $10^4\pi$ J
   (D) $10^{-3}\pi$ J

4. A cyclic process is shown in the figure. Work done during isobaric expansion is:



   (A) 1600 J
   (B) 100 J
   (C) 400 J
   (D) 600 J

5. The first law of thermodynamics is a special case of

   (A) Newton's law
   (B) Stefan's law
   (C) Law of heat exchange
   (D) Law of conservation of energy

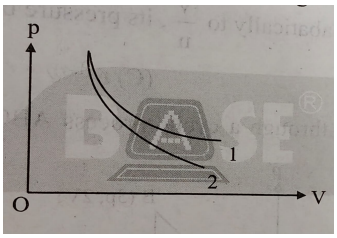
6. In a thermodynamic process, a system absorbs 2 kcal of heat and does 800 J of work. The change in internal energy is

   (A) 7600 J
   (B) 2000 J
   (C) 4800 J
   (D) -7600 J

7. One gram of water at 10^5 Pa is converted into steam at 100°C. Volume of steam is 1671 cc. If latent heat of evaporation is 540 cal, the change in internal energy is

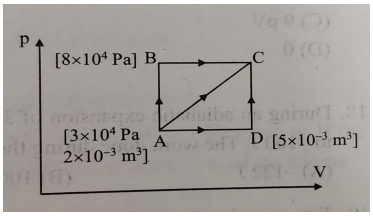
   (A) 167 cal
   (B) 540 cal
   (C) 500 cal
   (D) 581 cal

8. The figure shows p–V diagrams for two different gases during adiabatic processes. Curves 1 and 2 may correspond to:



   (A) helium and oxygen
   (B) oxygen and helium
   (C) hydrogen and oxygen
   (D) helium and argon

9. The p–V diagram of a thermodynamic process is shown. Along AB, 600 J of heat is added and along BC, 200 J of heat is added. The change in internal energy from A to C is:



(A) 800 J (C) 560 J
(B) 600 J (D) 400 J

10. During an adiabatic process, the pressure of an ideal gas changes by Δp and its volume by ΔV. If $\gamma = \frac{C_p}{C_v}$, then $\frac{\Delta V}{V}$ is

(A) $-\frac{\Delta p}{p}$ (C) $-\frac{\Delta p}{\gamma p}$
(B) $-\gamma\frac{\Delta p}{p}$ (D) $\frac{\Delta p}{\gamma^2 p}$

11. The temperature of 2 moles of a gas, held at constant volume, is changed from 100°C to 120°C. The change in internal energy is 80 J. The specific heat of the gas at constant volume is

(A) 2 J $mol^{-1}K^{-1}$ (C) 6 J $mol^{-1}K^{-1}$
(B) 4 J $mol^{-1}K^{-1}$ (D) 8 J $mol^{-1}K^{-1}$

12. The ratio of the slopes of isothermal and adiabatic curves on a p–V diagram is

(A) $\frac{1}{\gamma}$ (C) $\frac{1}{\gamma^2}$
(B) γ (D) γ^2

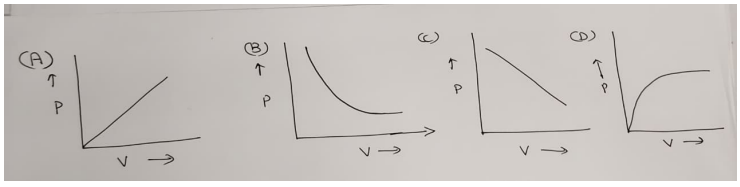
13. If the amount of heat given to a system is 35 J and the amount of work done by the system is −15 J, then the change in internal energy of the system is

(A) −50 J (C) 30 J
(B) 20 J (D) 50 J

14. No heat flows between the system and the surroundings. The thermodynamic process is

(A) isothermal (C) adiabatic
(B) isochoric (D) isobaric

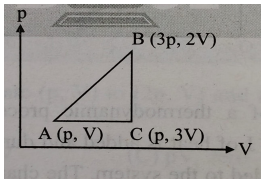
15. In a thermodynamic process (expansion), the temperature of the gas remains constant. The graph that represents the variation of pressure and volume is:



16. When an ideal gas is compressed isothermally to $\frac{V}{n}$, its pressure becomes p_1. V is the initial volume of the gas. If the gas is compressed adiabatically to $\frac{V}{n}$, its pressure becomes p_2. The ratio $\frac{p_1}{p_2}$ is

(A) n (C) $n^{(1-\gamma)}$
(B) n^{γ} (D) 1

17. An ideal monoatomic gas is taken through a cyclic process ABCA as shown in the p–V diagram. The work done per cycle is:



(A) $2pV$ (C) $9pV$
(B) $6pV$ (D) 0

18. During an adiabatic expansion of 3 moles of a gas, the change in internal energy is −100 J. The work done during the process is

(A) −122 J (C) −300 J
(B) 100 J (D) 300 J

19. Two samples A and B of a gas at the same initial state are compressed from volume V to $\frac{V}{2}$. A is compressed isothermally and B adiabatically. The final pressure of A will be _____ that of B.

(A) greater than
(B) less than
(C) same as
(D) greater than or equal to

20. An ideal gas undergoes a process according to the law $p = aV$ (where a is constant). Initial temperature is 300 K. The change in temperature when its volume doubles is

(A) 500 K (C) 1000 K
(B) 900 K (D) 1200 K

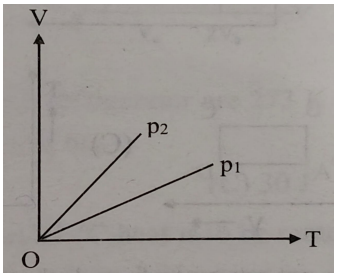
21. When a work W is done on an ideal gas consisting of N monoatomic molecules in thermal isolation, the temperature is increased by

(A) $\frac{W}{2Nk}$ (C) $\frac{2W}{3Nk}$
(B) $\frac{W}{2Nk}$ (D) $\frac{2W}{5Nk}$

22. Two identical containers A and B with frictionless pistons contain ideal gases of masses m_A and m_B respectively. They have the same temperature and volume. When gases expand isothermally to the same final volume $2V$, changes in pressure are Δp and $1.5\Delta p$ respectively. Then

(A) $2m_A = 3m_B$ (C) $m_A = m_B$
(B) $3m_A = 2m_B$ (D) $9m_A = 4m_B$

23. The figure shows a plot of V v/s T at two different pressures p_1 and p_2. Choose the correct statement.



(A) $p_1 = p_2$ (C) $p_2 > p_1$
(B) $p_2 < p_1$ (D) none of these

24. 1 kg of water is heated from 40°C to 70°C at constant volume. Specific heat of water is 4148 J $kg^{-1}K^{-1}$. The change in internal energy is

(A) 2.44×10^5 J (C) 1.24×10^5 J
(B) 1.62×10^5 J (D) 2.62×10^5 J

25. Heat required to raise the temperature of 2 moles of an ideal gas at constant pressure from 30°C to 35°C is 70 cal. The heat required to raise the temperature of the same gas from 50°C to 55°C at constant volume is

(A) 50 cal (C) 60 cal
(B) 70 cal (D) 65 cal

26. The change in internal energy of a given mass of gas when its volume changes from V to $2V$ at constant pressure is

(A) $\frac{R}{\gamma - 1}$ (C) $\frac{\gamma pV}{1 - \gamma}$
(B) $\frac{pV}{\gamma - 1}$ (D) pV

27. A frictionless heat engine can have efficiency $\eta = 1$ if its exhaust temperature is

(A) 0°C
(B) 0 K
(C) ∞
(D) equal to its input temperature

28. An ideal Carnot engine works between temperatures T_1 and T_2 with efficiency η. If both temperatures are increased by 100 K, the new efficiency η' is

(A) $\eta' = \eta$ (C) $\eta' > \eta$
(B) $\eta' < \eta$ (D) $\eta' = 0$

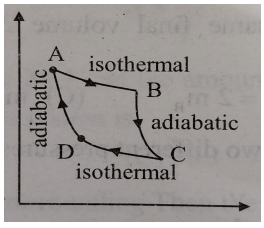
29. Three designs are proposed for an engine operating between 500 K and 300 K. For 1 kcal of heat input: Design A produces 3000 J, B produces 2000 J, and C produces 1000 J of work. The possible design is

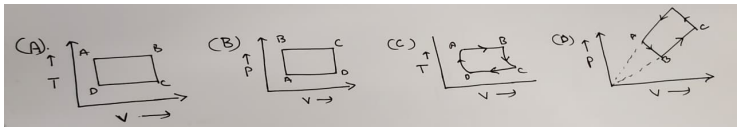
(A) A only (C) C only
(B) B only (D) All

30. A Carnot engine works first between 200°C and 0°C and then between 0°C and −200°C. The ratio of efficiencies in the two cases is

(A) 0.57 (C) 10
(B) 0.67 (D) 1

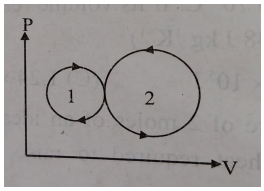
31. The p–V diagram of a Carnot cycle is shown in the figure. The process is represented as:





(A) option A
(B) option B
(C) option C
(D) option D

32. In the following indicator diagram, the net amount of work done will be:



(A) positive
(B) negative
(C) zero
(D) infinity

33. Two Carnot engines A and B have the same efficiency. A receives heat from a source at temperature 800 K and rejects it to a sink at x K. B receives heat rejected by A and rejects it to another sink at 300 K. The temperature x is nearly

(A) 200 K
(B) 480 K
(C) 350 K
(D) 515 K

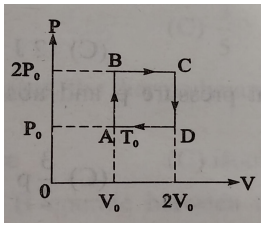
34. Two perfect monoatomic gases of n_1 and n_2 moles at temperatures T_1 and T_2 are mixed. The temperature of the mixture, if there is no loss of energy, is

(A) $\frac{n_1T_1 + n_2T_2}{2}$
(B) $\frac{T_1 + T_2}{2}$
(C) $T_1 + T_2$
(D) $\frac{n_1T_1 + n_2T_2}{n_1 + n_2}$

35. A Carnot heat engine working between temperature limits 400 K and 800 K has a work output of 1000 J per cycle. The heat supplied to the engine per cycle is

(A) 1500 J
(B) 1000 J
(C) 2460 J
(D) 2000 J

36. N moles of a monoatomic gas are taken round the reversible rectangular cycle ABCDA as shown in the diagram. The temperature at A is T_0. The thermodynamic efficiency of the cycle is:



(A) 15%
(B) 20%
(C) 25%
(D) 50%

37. The inside and outside temperatures of a refrigerator are 273 K and 303 K respectively. For every joule of work done, the heat delivered to the surroundings is

(A) 10 J
(B) 20 J
(C) 30 J
(D) 50 J

38. A refrigerator works between 0°C and 27°C. Heat is removed from the refrigerated space at the rate of 50 kcal/min. The power of the motor of the refrigerator is

(A) 0.346 kW
(B) 3.46 kW
(C) 34.6 kW
(D) 346 kW

39. The freezer in a refrigerator is located at the top section so that

(A) the entire chamber is cooled quickly due to convection
(B) the motor is not heated
(C) the heat gained from the environment is high
(D) the heat gained from the environment is low

D.P.P – II (K-CET)

First Law of Thermodynamics

1. First law of thermodynamics states that:

   (a) system can do work
   (b) system has temperature
   (c) system has pressure
   (d) heat is a form of energy

2. First law of thermodynamics can be explained on the basis of:

   (a) Joule's law
   (b) Boyle's law
   (c) Charles' law
   (d) Avogadro law

3. If ΔQ and ΔW represent the heat supplied to a system and the work done on the system respectively, the first law can be written as:

   (a) $\Delta Q = \Delta U + \Delta W$
   (b) $\Delta Q = \Delta W - \Delta U$
   (c) $\Delta Q = \Delta U - \Delta W$
   (d) $\Delta Q = -\Delta W - \Delta U$

4. In the figure, two processes A and B take a system from the same initial to the same final state. If ΔQ_A and ΔQ_B are the heats supplied, then:

   (a) $\Delta Q_A = \Delta Q_B$
   (b) $\Delta Q_A < \Delta Q_B$
   (c) $\Delta Q_A > \Delta Q_B$
   (d) cannot be determined

5. A system is taken from a given initial state to a final state along various paths on a p–V diagram. The quantity independent of the path is:

   (a) heat transferred Q
   (b) work done W
   (c) Q but not W
   (d) $(Q - W)$

6. A system absorbs 2 kilocalorie of heat and does 500 J of work. The change in internal energy is:

   (a) 7000 J
   (b) 7900 J
   (c) 9000 J
   (d) 10000 J

7. When the work done is 333 cal and the change in internal energy is 167 cal, the heat supplied is:

   (a) 333 cal
   (b) 500 cal
   (c) 167 cal
   (d) 166 cal

8. If 150 J of heat is added to a system and the work done by the system is 110 J, the change in internal energy is

   (a) zero
   (b) 40 J
   (c) 110 J
   (d) 150 J

9. The latent heat of vaporisation of water is 2240 J g^{-1}. If the work done in expansion of 1 g is 168 J, the increase in internal energy is:

   (a) 2240 J
   (b) 2072 J
   (c) 1680 J
   (d) 168 J

10. Two kilogram of water is converted into steam at atmospheric pressure. Volume changes from 2×10^{-3} m^3 to 3.34 m^3. The work done is approximately:

    (a) 200 kJ
    (b) 234 kJ
    (c) 340 kJ
    (d) 468 kJ

11. When 1500 J of heat is supplied at constant pressure 2.1×10^5 N m^{-2}, the increase in internal energy is:

    (a) 450 J
    (b) 525 J
    (c) 975 J
    (d) 2025 J

12. A system absorbs 100 cal of heat and performs 30 J of work. The change in internal energy is:

    (a) 14 J
    (b) 140 J
    (c) 390 J
    (d) 450 J

13. A sample of 0.1 g water at 100°C requires 54 cal to convert into steam. The change in internal energy is:

    (a) 104.3 J
    (b) 208.7 J
    (c) 42.2 J
    (d) 84.5 J

14. 1 kg of water is heated from 40°C to 70°C at constant volume. The change in internal energy is:

    (a) 2.44×10^5 J
    (b) 1.62×10^5 J
    (c) 1.24×10^5 J
    (d) 2.62×10^5 J

15. Heat supplied in path 1 is 1100 J and work done exceeds path 2 by 150 J. Heat exchanged in path 2 is:

    (a) 800 J
    (b) 750 J
    (c) 1050 J
    (d) 950 J

16. An electric heater supplies heat at 120 W. If work is done at 80 J s^{-1}, the rate of increase of internal energy is:

    (a) 30 J s^{-1}
    (b) 40 J s^{-1}
    (c) 50 J s^{-1}
    (d) 60 J s^{-1}

17. A system absorbs 2 kcal of heat and does 400 J of work. The change in internal energy is:

(a) 2 kJ
(b) 8 kJ
(c) 3.5 kJ
(d) 5.5 kJ

18. Along path iaf, $Q = 50$ cal and $W = 20$ cal. Work done along ibf is:

(a) 14 cal
(b) 6 cal
(c) 16 cal
(d) 66 cal

19. Which leads to an increase in internal energy?:

(I) $Q > 0$, $W = 0$
(II) $Q < 0$, $W = 0$
(III) $W > 0$, $Q = 0$
(IV) $W < 0$, $Q = 0$

20. $U_a = 10$ J. Along adc: $\delta Q_1 = 50$ J, $\delta W_1 = 20$ J. Along abc: $\delta Q_2 = 40$ J. Work done along abc is

(a) 10 J
(b) 12 J
(c) 36 J
(d) 6 J