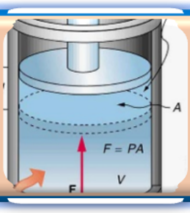


UNIT-16 FIRST LAW OF THERMODYNAMICS

Heat and Work

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

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

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:

Heat 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:

$$Q = mc\,\Delta T$$

Where Q *is the heat transferred*
m *is the mass of the substance*

c is the specific heat capacity of the substance,
ΔT is the temperature change

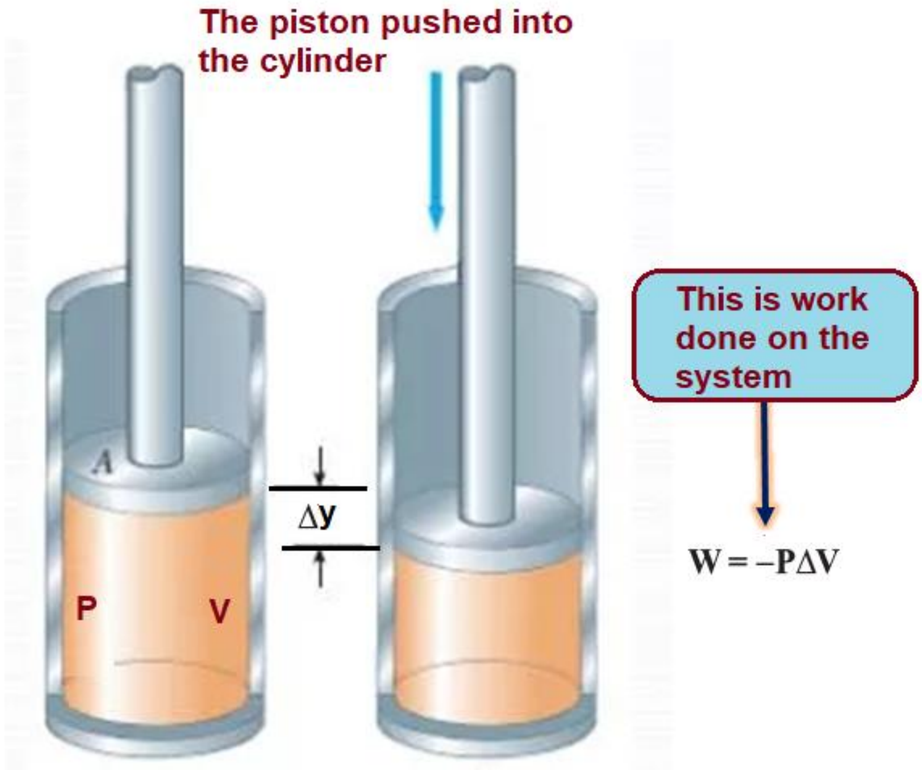
The above equation, known as the heat transfer equation, highlights the dependence of heat transfer on the material properties and the extent of temperature change.

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.

HEAT AND WORK IN THERMODYNAMICS

Consider a thermodynamic system of a gas contained in a cylinder fitted with a moveable frictionless piston of cross-sectional area **A**. The pressure exerted on the piston the gas undergoes compression, and the piston moves downward, (a force acting over a distance). The work done can be expressed as the product of the force and distance



$$W (= \text{ force }) \cdot (\text{displacement })$$
$$W = F\, S \cos\theta$$
$$W = F\, \Delta y \cos 180$$
$$W = (P\, A)\, \Delta y\, (-1)$$
$$W = -\, P\, (A\, \Delta y)$$
$$W = -\, P\, \Delta V$$

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

INTERNAL ENERGY

DEFINITION

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.***

Internal energy is denoted by U

UNIT

The SI unit if internal energy is Joule (J)

KEY POINTS

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, and 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

INTERNAL ENERGY AND TEMPERATURE

The internal energy of an object is intrinsically related to its temperature. When a container containing gas molecules is heated up, the molecules begin to move around faster, increasing their kinetic energy, If the object is a solid, where the molecules are tightly packed, when heated the molecules begin to vibrate more. Molecules in liquids and solids have both kinetic and potential energy because they are close together and bound by intermolecular forces, however, ideal gas molecules are assumed to have no intermolecular forces. This means there is no potential energy, only kinetic energy

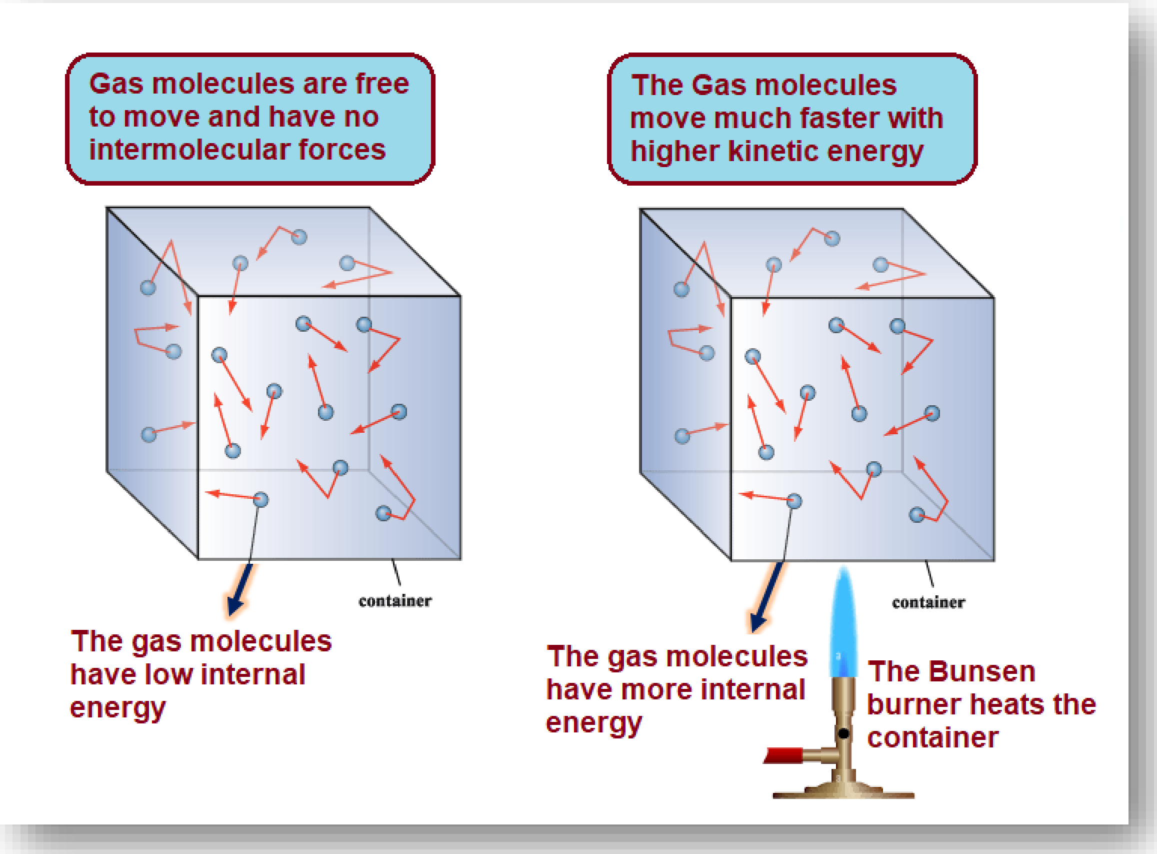
The (change in) internal energy of an ideal gas is equal to:

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

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

$$\Delta U \; \alpha \; \Delta T$$

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



THERMODYNAMICS

DEFINITION

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

Or

The study of the relationship between heat, work, other forms of energy, and methods used for their transformation is called thermodynamics

SYSTEM AND SURROUNDING

System and surroundings are relative terms. A body or collection of matter that has distinct and well defined boundaries is called system and the remaining part of the universe is surroundings.

EXAMPLES

1 In the case of the refrigerator, it is a system and the rest of it is surrounding

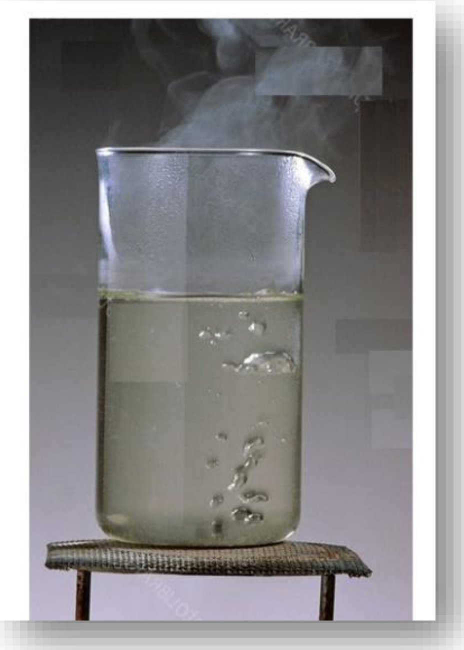
2 A boy holding a steaming cup of coffee. By the thermodynamic definition, the liquid coffee is the system and out of it is all are surroundings.

THERMODYNAMICS SYSTEM

There are three kinds of systems as

1 OPEN SYSTEM

In an open system, both heat and matter can be exchanged with the surroundings.



EXAMPLE

Consider a boiling water beaker without a lid. Heat escapes into the air, and at the same time, water molecules also escape into the air.

CLOSED SYSTEM

A closed system is a natural physical system that does not allow the transfer of matter in or out of the system but only heat energy can be exchanged with the surrounding

EXAMPLE

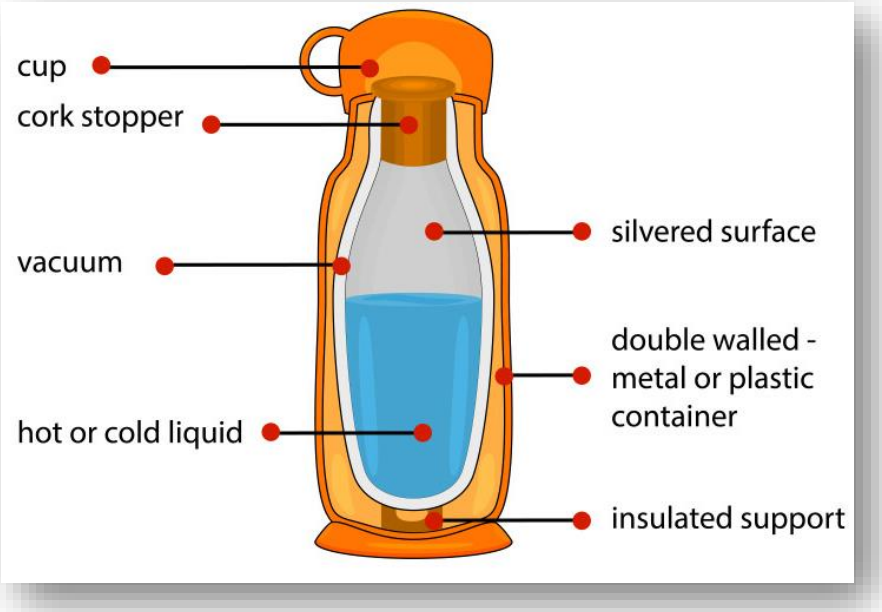
A pressure cooker is a closed system. It allows heat to be transferred from the stove to the water. Steam is not allowed to escape.

ISOLATED SYSTEM

A system that cannot exchange matter or energy with the surroundings, is known as an isolated system

EXAMPLE

A thermos flask is an isolated system.

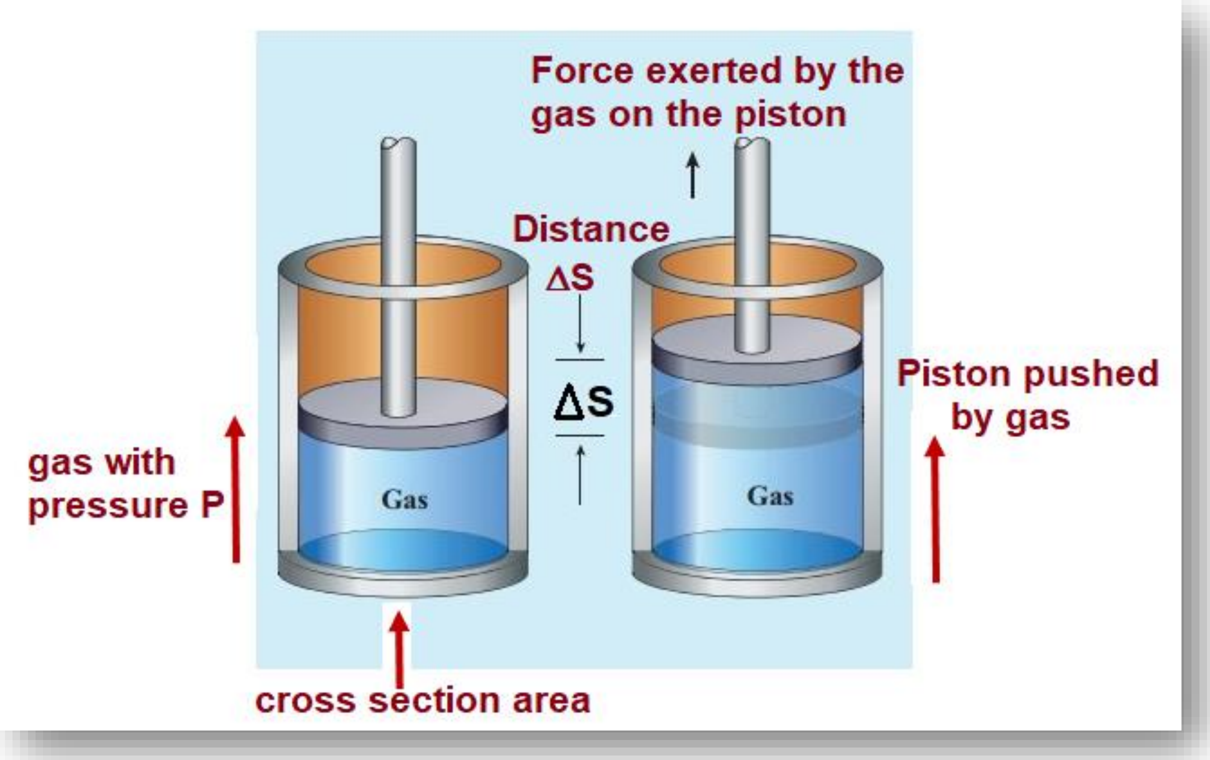


WORK DONE BY A GAS

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:

$$W = P\Delta V$$

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



DERIVATION

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

$$\boldsymbol{F} = (\boldsymbol{P})(\boldsymbol{A})$$

Where: p = pressure of the gas (Pascal)
A = cross-sectional area of the cylinder (m^2)

The definition of work done is:

$$\boldsymbol{W} = (\boldsymbol{F})(\Delta \boldsymbol{S})$$

Where:
F = force (N)
ΔS = displacement in the direction of force (m)

The displacement of the gas ΔS multiplied by the cross-sectional area A is the increase in volume ΔV of the gas:

$$\boldsymbol{W} = (\boldsymbol{P}\,\boldsymbol{A})\ (\Delta \boldsymbol{S})$$
$$\boldsymbol{W} = (\boldsymbol{P})\ (\boldsymbol{A}\,\Delta \boldsymbol{S})$$
$$\boldsymbol{W} = \boldsymbol{P}\,\Delta \boldsymbol{V} \qquad \{\Delta \boldsymbol{V} = (\boldsymbol{A}\,\Delta \boldsymbol{S})\}$$

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

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 $(\boldsymbol{W} = \boldsymbol{P}\Delta \boldsymbol{V})$.
When the gas is compressed (V decreases), work is done on the gas $(\boldsymbol{W} = -\boldsymbol{P}\Delta \boldsymbol{V})$

FIRST LAW OF THERMODYNAMICS

INRTODUCTION

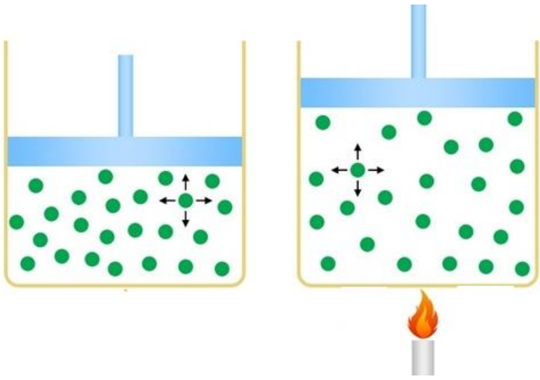
The first law of thermodynamics is based on the idea that energy can neither be created nor destroyed in any thermodynamic system. So, it is a particular form of the law of conservation of energy, which deals only with the heat energy

"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.

$$\Delta \boldsymbol{Q} = \Delta \boldsymbol{U} + \Delta \boldsymbol{W}$$

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 decreases



STATEMENT

"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

$$\Delta U = \Delta Q - W \ldots\ldots\ldots (i)$$

$\Delta U \rightarrow$ change in internal energy of the system
$\Delta Q \rightarrow$ net heat transferred into the system
$W \rightarrow$ 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 equation (i), can be written as

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

These two equations represent the mathematical form of first law of thermodynamics

SPECIAL CASES

1. ISOLATED SYSTEM

Consider an isolated system- that is, one that does not interact with its surroundings. In this case, no energy transfer by heat takes place and the value of work done by the system is zero. For such system, $\Delta Q = 0$ and $\Delta W = 0$.

$$\therefore \quad \Delta Q = \Delta W + \Delta U$$
$$0 = O + \Delta U$$
$$\text{or} \quad \Delta U = 0$$

This indicates that, the internal energy of an isolated system remains constant.

2. CYCLIC PROCESS

When a system attains its initial state after undergoing a series of changes, the process is called cyclic. In such process $U_i = U_f$ & $\Delta U = O$.

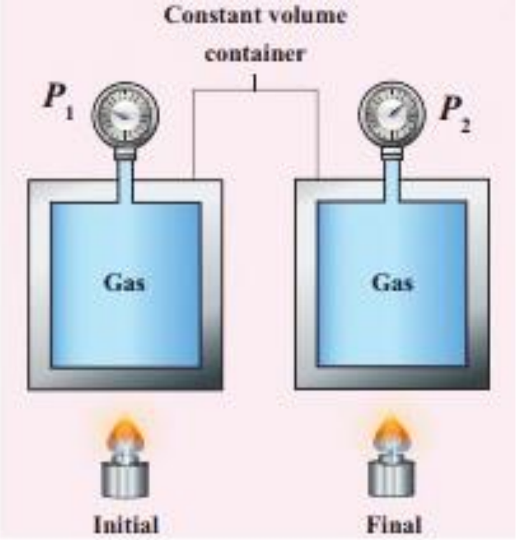
$$\therefore \quad \Delta Q = \Delta W + \Delta U$$
$$\Delta Q = \Delta W + O$$
$$\text{or} \quad \Delta Q = \Delta W$$

This indicate that in a cyclic process the heat supplied to the system equals the work done by the system.

APPLICATIONS OF FIRST LAW OF THERMODYNAMICS

1 ISOCHORIC PROCESS or ISOMETRIC PROCESS

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



Consider a gas enclosed in a cylinder fitted having a conducting base and non-conducting walls and with a fixed piston, as shown in fig

Let heat ΔQ be imparted to gas. The gas is then heated at constant volume, then since there is no change in volume, therefore, no work is done during the process ($\Delta W = 0$).

Applying first law of thermodynamics

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

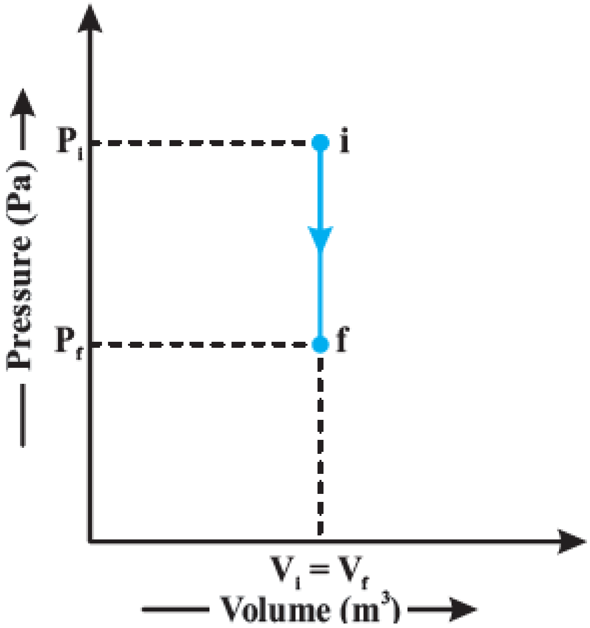
$$\Delta Q = \Delta U + 0 \qquad [\Delta W = 0]$$

$$\Delta Q = \Delta U$$

This means that in a isochoric process the entire amount of heat supplied to the gas is converted to the internal energy of the gas.

ISOCHORIC GRAPH

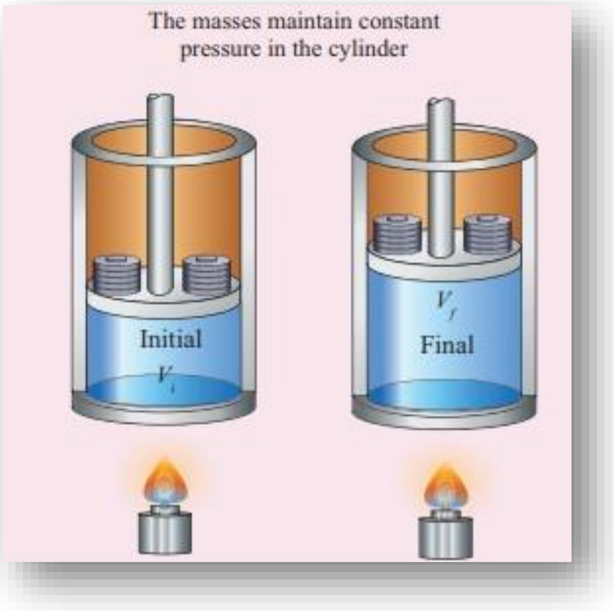
The graph between pressure and volume in isochoric process is called isochoric graph



1. ISOBARIC PROCESS

***The thermodynamics process during which the pressure is kept constant is called isobaric process**. (**iso for equal, and bar for pressure**)*

Consider a gas contained in a cylinder having a conducting base and non-conducting walls and frictionless piston of cross-section area (A) as shown in figure.



When the gas is heated at certain amount of heat energy is transferred into the system. The gas expands and move the piston outward, temperature changes from T_1 to T_2 and volume changes from V_1 to V_2. If the displacement is kept very small, the pressure of gas will not be changed much and can be considered constant.

Work done by gas

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

Now, according to 1st law of thermodynamics

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

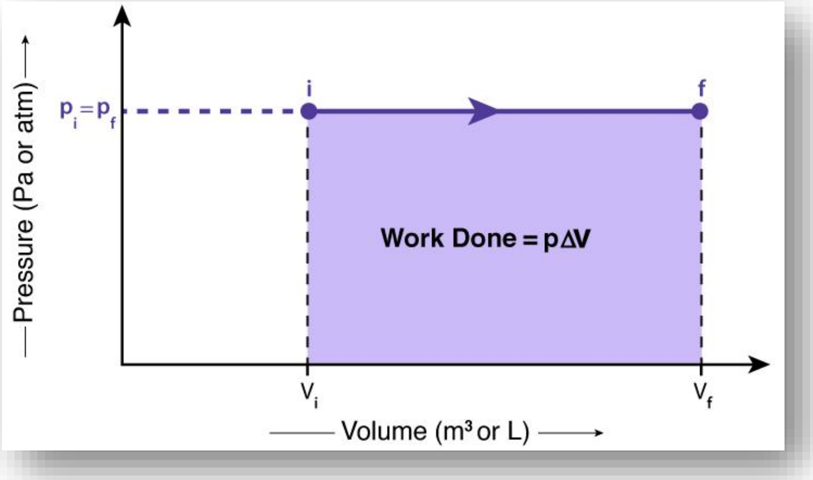
Substituting the expression of work in above equation

$$\Delta W = \Delta U + P\,\Delta V$$

In isobaric process, heat supplied to the system is used to increase the internal energy and to do work of expansion.

ISOBARIC GRAPH

The graph between pressure and volume in isbaric process is called isobar graph



3.ISOTHERMAL PROCESS

The thermodynamics process which takes place at a constant temperature is called isothermal process. (iso for equal, and thermal for temperature)

ISOTHERMAL EXPANSION

The cylinder is placed on a heat reservoir at temperature 'T'. Now the gas is allowed to expand by decreasing pressure on the piston. Due to this expansion, temperature of the gas falls down. But heat is supplied by the reservoir, so that temperature remains constant. (ΔU = O). pressure changes from P_2 to P_1 and volume changes from V_1 to V_2 but temperature remains constant

Applying first law of thermodynamics.

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

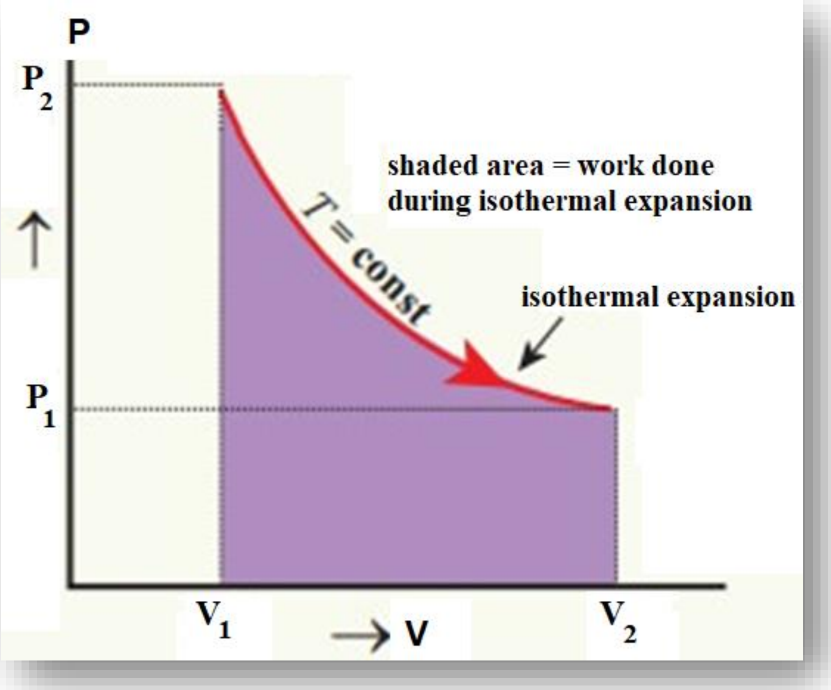
$$\Delta Q = 0 + \Delta W$$

$$\Delta Q = \Delta W$$

This indicates that the heat supplied in isothermal process is used to do work.

isothermal expansion graph

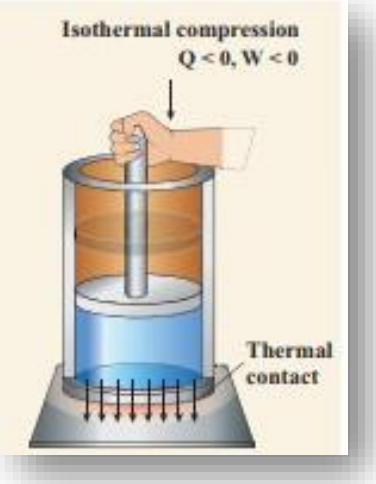
the graph between pressure and volume in isothermal process is called isotherm.



ISOTHERMAL FORM OF BOYLE'S LAW

$$PV = constant$$

ISOTHERMAL COMPRESSION



The isothermal compression takes place when the piston of the cylinder is pushed. This will increase the internal energy which will flow out of the system through thermal contact, this is shown in Figure. The temperature remains constant. (∆U = O). pressure changes from P_1 to P_2 and volume changes from V_2 to V_1 but temperature remains constant

Applying first law of thermodynamics.

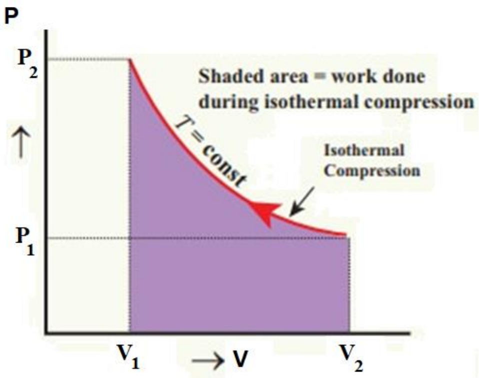
$$\Delta Q = \Delta U + \Delta W$$

$$-\Delta Q = 0 - \Delta W$$

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

isothermal Compression graph

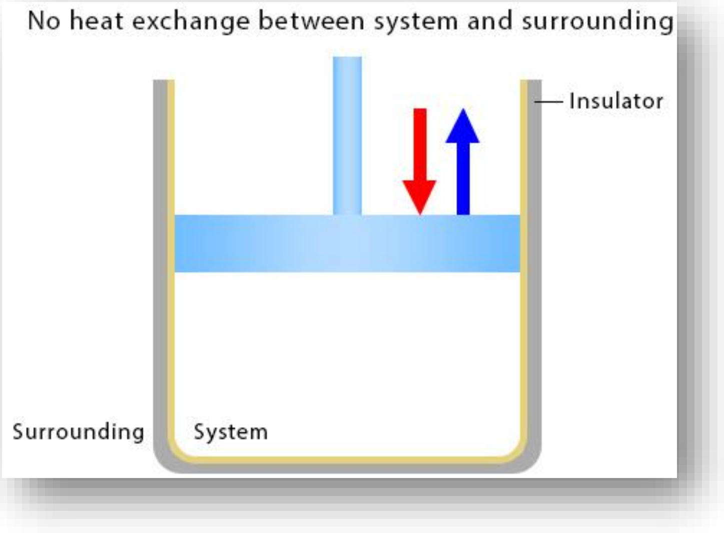
the graph between pressure and volume in isothermal process is called isotherm.



4. ADIABATIC PROCESS

An adiabatic process is one in which no heat is transferred into or out of the system. (The Greek word adiabatos means“ not possible or cannot be crossed”

Consider a gas enclosed in a cylinder fitted with a moveable frictionless piston. In this process, the cylinder is placed on an insulator so, that no heat can enter or leave the system.



ADIABATIC EXPANSION

If the gas is allowed to expand, we decrease the pressure on the piston, and the gas will be cooled, so its temperature will decrease. Therefore, the internal energy of gas will decrease, and the work done by the gas

Applying the first law of thermodynamics:

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

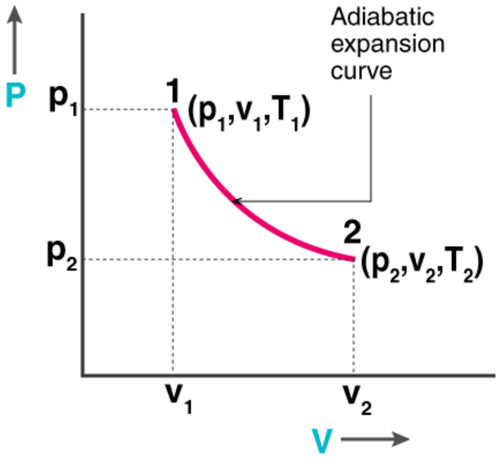
$$0 = \Delta W + \Delta U \Theta \Delta Q = O$$

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

This shows that in the adiabatic process, the system does work, this work done equals the internal energy change.

THE GRAPH BETWEEN PRESSURE AND VOLUME

The graph between pressure and volume for the adiabatic process is called adiabatic. It is steeper than isotherm.



Adiabatic process

ADIABATIC FORM OF BOYLE'S LAW

$$PV^{\gamma} = \text{constant}$$

Where $\gamma = \frac{C_P}{C_V}$

C_p = molar specific heat at constant pressure

C_v = molar specific heat at constant volume

ADIABETIC COMPRESSION

If the gas is compressed by suddenly increasing the pressure on the piston, the gas temperature will be increased. Therefore, the internal energy of the will increase According to the first law of thermodynamics.

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

$$0 = \Delta W + \Delta U \Theta \Delta Q = O$$

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

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

HEAT CAPACITY

Consider a body of which no change of state takes place. Transfer of energy, **Q**, into the body causes the internal energy to increase by **ΔU**, and the temperature to rise **by** ΔT

DEFINITION

The heat capacity, **C** of a body is defined as the amount of thermal energy required to raise the temperature of the body by 1 k or 1 ^{0}C.

$$C = \frac{Q}{\Delta T}$$

Where, Q is the thermal energy absorbed

ΔT is the change of temperature

SI UNIT

The SI unit of heat capacity is **J K^{-1}** (but J ^{0}C^{-1} is also commonly used).

Rewriting the equation, we have

$$Q = C\,\Delta T$$

$$\textit{Thermal energy} = (\textit{Heat Capacity})\ (\textit{Change in temperature})$$

SPECIFIC HEAT CAPACITY

DEFINITION

The amount of heat required to raise the temperature of the unit mass of a substance through one Kelvin is known as a specific heat capacity or specific heat.

FORMULA

Let 'ΔQ' be the heat supplied.

ΔT = change in temperature

m = mass of a substance

Then by definition:

$$C = \frac{Q}{m\,\Delta T}$$

UNIT

The SI unit of specific heat is **Jkg^{-1}k^{-1}**

MOLAR SPECIFIC HEAT

DEFINITION

The amount of heat required to raise the temperature of one mole of a substance through one Kelvin is called molar-specific heat.

OR

The product of the molecular weight of a substance and specific heat is known as molar-specific heat.

MATHEMATICAL EXPRESSION

The specific heat 'C' of a substance of mass 'm' is given by

$$C = \frac{Q}{m\Delta T} \text{ - - - - - - - - - - (1)}$$

But, $\textit{the number of moles} = \frac{\textit{mass}}{\textit{molecular weight}}.$

$$n = \frac{m}{M}$$

$$n\,M = m$$

Now equation (2) becomes

$$C = \frac{Q}{nM\Delta T}$$

Or $\quad MC = \frac{Q}{n\,\Delta T} \qquad \Theta\ C_m = MC$

$$C_m = \frac{Q}{n\Delta T} \text{ - - - - - - - - - - (2)}$$

Where C_m is molar-specific heat.

UNIT OF MOLAR SPECIFIC HEAT

The SI unit of molar-specific heat is $Jmol^{-1}K^{-1}$.

MOLAR SPECIFIC HEATS OF GASES

Gases have two molar-specific heats:
molar specific heat at constant volume (Cv) and
molar specific heat at constant pressure (Cp).

REASON

Since the temperature of a gas can be changed in two different ways i.e.,
at constant volume
at constant pressure, therefore molar-specific heat of gases can be defined in two ways as follows.

MOLAR SPECIFIC HEAT AT CONSTANT VOLUME (Cv)

The amount of heat required to raise the temperature of one mole of a gas at constant volume through 1K is called molar-specific heat at constant volume.

FORMULA

Mathematically it is given by $\quad C_v = \frac{Q}{n\ \Delta T}$

UNIT

The SI units of molar specific heat at constant volume are $Jmol^{-1}K^{-1}$.

MOLAR SPECIFIC HEAT AT CONSTANT PRESSURE

Amount of heat required to raise the temperature of one mole of a gas at constant pressure, through one Kelvin is known as molar specific heat constant pressure.

FORMULA

Mathematically it is given by $C_p = \frac{\Delta Q}{n\,\Delta T}$

UNIT

The SI units of molar-specific heat at constant pressure are Jmol^{-1}K^{-1}.

Prove that $C_V = \frac{3}{2}R$ ***and*** $C_P = \frac{5}{2}R$ ***for monoatomic gas***

We know that

$$C_P - C_V = R$$

Dividing both sides by C_V

$$\frac{C_P}{C_V} - \frac{C_V}{C_V} = \frac{R}{C_V} \qquad \left\{\frac{C_P}{C_V} = \gamma\right\}$$

$$\gamma - 1 = \frac{R}{C_C}$$

γ for monoatomic gases is 1.67

$$1.67 - 1 = \frac{R}{C_V}$$

$$0.67 = \frac{R}{C_V}$$

$$C_V = \frac{R}{0.67}$$

$$C_V = 1.5\ R$$

$$C_V = \frac{15}{10}\ R$$

$$C_V = \frac{3}{2}\ R$$

We know that

$$C_P - C_V = R$$

Dividing both sides by C_P

$$\frac{C_P}{C_P} - \frac{C_V}{C_P} = \frac{R}{C_P}$$

$$1 - \frac{C_V}{C_P} = \frac{R}{C_P}$$

$$1 - \frac{1}{\frac{C_P}{C_V}} = \frac{R}{C_P} \qquad \left\{\frac{C_P}{C_V} = \gamma\right\}$$

γ for monoatomic gases is 1.67

$$1 - \frac{1}{1.67} = \frac{R}{C_P}$$

$$\frac{1.67 - 1}{1.67} = \frac{R}{C_P}$$

$$\frac{0.67}{1.67} = \frac{R}{C_P}$$

$$C_P = 2.5\ R$$

$$C_P = \frac{25}{10}R$$

$$C_P = \frac{5}{2}R$$

Prove that $C_V = \frac{5}{2}R$ and $C_P = \frac{7}{2}R$ for diatomic gas

We know that

$$C_P - C_V = R$$

Dividing both sides by C_V

$$\frac{C_P}{C_V} - \frac{C_V}{C_V} = \frac{R}{C_V} \qquad \left\{\frac{C_P}{C_V} = \gamma\right\}$$

$$\gamma - 1 = \frac{R}{C_C}$$

γ for diatomic gases is 1.4

$$1.4 - 1 = \frac{R}{C_V}$$

$$0.4 = \frac{R}{C_V}$$

$$C_V = \frac{R}{0.4}$$

$$C_V = \frac{1}{0.4} R$$

$$C_V = 2.5 R$$

$$C_V = \frac{25}{10} R$$

$$C_V = \frac{5}{2} R$$

We know that

$$C_P - C_V = R$$

Dividing both sides by C_P

$$\frac{C_P}{C_P} - \frac{C_V}{C_P} = \frac{R}{C_P}$$

$$1 - \frac{C_V}{C_P} = \frac{R}{C_P}$$

$$1 - \frac{1}{\frac{C_P}{C_V}} = \frac{R}{C_P} \qquad \left\{\frac{C_P}{C_V} = \gamma\right\}$$

$$1 - \frac{1}{\gamma} = \frac{R}{C_P}$$

γ for diatomic gases is 1.4

$$1 - \frac{1}{1.4} = \frac{R}{C_P}$$

$$\frac{1.4 - 1}{1.4} = \frac{R}{C_P}$$

$$\frac{0.4}{1.4} = \frac{R}{C_P}$$

$$C_P = \frac{1.4}{0.4} R$$

$$C_P = 3.5 R$$

$$C_P = \frac{35}{10} R$$

$$C_P = \frac{7}{2} R$$

$Prove\ that\ C_V = 3R$ and $C_P = 4R\ for\ polyatomic\ gas$

We know that

$$C_P - C_V = R$$

Dividing both sides by C

$$\frac{C_P}{C_V} - \frac{C_V}{C_V} = \frac{R}{C_V} \qquad \left\{\frac{C_P}{C_V} = \gamma\right\}$$

$$\gamma - 1 = \frac{R}{C_C}$$

γ for polyatomic gases is 1.33

$$1.33 - 1 = \frac{R}{C_V}$$

$$0.33 = \frac{R}{C_V}$$

$$C_V = \frac{R}{0.33}$$

$$C_V = \frac{1}{0.33}\ R$$

$$C_V = 3\ R$$

We know that

$$C_P - C_V = R$$

Dividing both sides by C_P

$$\frac{C_P}{C_P} - \frac{C_V}{C_P} = \frac{R}{C_P}$$

$$1 - \frac{C_V}{C_P} = \frac{R}{C_P}$$

$$1 - \frac{1}{\frac{C_P}{C_V}} = \frac{R}{C_P} \qquad \left\{\frac{C_P}{C_V} = \gamma\right\}$$

$$1 - \frac{1}{\gamma} = \frac{R}{C_P}$$

γ for polyatomic gases is 1.33

$$1 - \frac{1}{1.33} = \frac{R}{C_P}$$

$$\frac{1.33 - 1}{1.33} = \frac{R}{C_P}$$

$$\frac{0.33}{1.33} = \frac{R}{C_P}$$

$$C_P = \frac{1.33}{0.33}\ R$$

$$C_P = 4\ R$$

PROVE THAT C_P – C_V = R

MOLAR SPECIFIC HEAT AT CONSTANT PRESSURE

The heat transferred (Q_p) at constant pressure, the process can be written as

$$Q_P = n\,C_P\,\Delta T$$

we know that the increase in enthalpy of a system is equal to the heat absorbed by it at a constant pressure $Q_p = \Delta H$

$$\Delta H = n\,C_P\,\Delta T$$

From the first law of thermodynamics

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

$$Q = \Delta U + P\,\Delta V \dots\dots\dots\dots(i)$$

Enthalpy is defined as equal to the system's internal energy plus the product of pressure and volume.

Equation (i) can be written as

$$\Delta H = \Delta U + P\,\Delta V \qquad \{P\,\Delta V = nR\,\Delta T\}$$

$$\Delta H = \Delta U + n\,R\,\Delta T \dots\dots(ii)$$

Comparing equation (i) and (ii), we get

$$n\,C_P\,\Delta T = \Delta U + n\,R\,\Delta T$$

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

$$(1)C_P\Delta T = \Delta U + (1)\,R\,\Delta T$$

$$C_P\,\Delta T = \Delta U + R\,\Delta T$$

$$C_P\,\Delta T - R\,\Delta T = \Delta U \dots\dots\dots\dots(iii)$$

2. MOLAR SPECIFIC HEAT AT CONSTANT VOLUME

The heat is transferred at constant volume, the process can be written as

$$Q_V = n\,C_V\,\Delta T$$

from the first law of thermodynamics

$$Q = \Delta U + \Delta W \qquad \{\Delta W = 0\}$$

$$Q_V = \Delta U + 0$$

$$Q_V = \Delta U$$

$$n\,C_V\,\Delta T = \Delta U$$

For one mole of a gas

$$C_V\,\Delta T = \Delta U \dots\dots\dots\dots..(iv)$$

Comparing equation (iii) and (iv), we get

$$C_P\,\Delta T - R\,\Delta T = C_v\,\Delta T$$

$$C_P\,\Delta T - C_v\,\Delta T = R\,\Delta T$$

$$\Delta T\,(C_P - C_v) = R\,\Delta T$$

$$C_P - C_v = R$$

SHORT REASONING QUESTIONS

1 Explain the concept of the first law of thermodynamics in your own words

Ans The first law of thermodynamics is a formulation of the law of conservation of energy in the context of thermodynamic processes. The law distinguishes two principal forms of energy transfer, heat, and thermodynamic work, that modify a thermodynamic system containing a constant amount of matter. The law also defines the internal energy of a system, an extensive property for taking account of the balance of heat and work in the system. Energy cannot be created or destroyed but can be transformed from one form to another. In an isolated system, the sum of all forms of energy is constant.

2 How does the first law of thermodynamics relate to the conservation of energy?

Ans The first law of thermodynamics states that energy can neither be created nor destroyed, only altered in form. For any system, energy transfer is associated with mass crossing the control boundary, external work, or heat transfer across the boundary. These produce a change of stored energy within the control volume. The mass flow of a fluid is associated with the kinetic, potential, internal, and "flow" energies that affect the overall energy balance of the system. The exchanges of external work and heat complete the energy balance. That is why the first law of thermodynamics is referred to as the conservation of energy principle, meaning that energy can neither be created nor destroyed, but rather transformed into various forms as the fluid within the control volume changes.

3 Distinguish between the work done by a system and the heat exchanged with the surroundings in the context of the first law.

Ans Work done by a system refers to the energy transferred when the system exerts a force to displace an object or when it undergoes a volume change against an external pressure.

Heat exchanged refers to the energy transferred between the system and its surroundings due to a temperature difference. It is the thermal energy moving into or out of a system.

work done by a system and heat exchanged with the surroundings are two distinct forms of energy transfer that, along with changes in internal energy, are governed by the first law of thermodynamics. They are vital for understanding how energy is conserved in thermodynamic processes.

4 Give an example of daily life that illustrates the principles of the first law of thermodynamics

Ans An excellent example of daily life illustrating the principles of the first law of thermodynamics, can be found in cooking. Consider the process of boiling water to make tea:

1. **Energy Input**: You place a kettle filled with water on a stove and turn on the heat source. The stove burns fuel (gas, for example) or uses electricity, transforming electrical energy into thermal energy.
2. **Energy Transfer**: The thermal energy from the stove is transferred to the kettle and subsequently to the water inside it. This increase in thermal energy causes the water molecules to move faster, raising the temperature of the water.
3. **Phase Change**: As the water reaches its boiling point (100°C), it undergoes a phase change from liquid to gas (steam). Here, energy is still conserved.

5 Explain the role of the system and its surroundings in the context of the first law of thermodynamics.

Ans The first law of thermodynamics is a principle that describes the conservation of energy in a system, The system is the specific part of the universe that is under consideration. The system is defined by its boundaries, which can be fixed or movable

The surroundings encompass everything outside the defined system. This can include other substances, the environment, or any other external factors that can exchange energy or matter with the system.

6 How is heat capacity related to the amount of energy to change the temperature of a substance

Ans Heat capacity is the amount of energy required to heat a substance that is directly proportional to its specific heat capacity. This means that for a given mass and temperature change, a substance with a higher specific heat capacity will require more energy to achieve the same temperature change compared to a substance with a lower specific heat capacity.