Topic: Concept of temperature and heat; isothermal and adiabatic changes; Reversible and irreversible processes; Zeroth law and First law of thermodynamics

Concepts of Temperature and Heat

Temperature :

Temperature is a physical quantity that expresses **the degree of *hotness or coldness*** of a body. It is a measure of ***the average kinetic energy of the particles*** in a substance. The higher the temperature, the faster the particles move.

Temperature Scales:

There are different temperature scales used to measure temperature:

- **Celsius (°C)**: Based on the freezing and boiling points of water.
- **Fahrenheit (°F)**: Mainly used in the United States.
- **Kelvin (K)**: The SI unit of temperature, used in scientific calculations. The Kelvin scale starts at absolute zero (0 K), the theoretically lowest temperature where all molecular motion stops.

The relationship between these temperature scales is given by:

$$T(C) = T(K) - 273.15$$

$$T(F) = \frac{9}{5}T(C) + 32$$

Problem 1:

Convert $25°C$ to Fahrenheit and Kelvin.

Solution

To convert from Celsius to Fahrenheit, use the formula:

$$T(F) = \frac{9}{5}T(C) + 32$$

Substituting $T(C) = 25$:

$$T(F) = \frac{9}{5} \times 25 + 32 = 45 + 32 = 77°F$$

To convert from Celsius to Kelvin, use the formula:

$$T(K) = T(C) + 273.15$$

Substituting $T(C) = 25$:

$$T(K) = 25 + 273.15 = 298.15\,K$$

Thus, the conversion results are:

$$T(F) = 77°F \quad \text{and} \quad T(K) = 298.15\,K$$

Zeroth Law of Thermodynamics

The zeroth law of thermodynamics states:

> ***If two systems, A and B, are each in thermal equilibrium with a third system, C, then A and B are in thermal equilibrium with each other.***

Example

When a thermometer and some other object are placed in contact with each other, they eventually reach thermal equilibrium. The reading of the thermometer is then taken to be the temperature of the other object. This process provides consistent and useful temperature measurements because of the zeroth law of thermodynamics.

Heat

Definition: Heat is the **form of energy** that is transferred between systems or objects with different temperatures. It flows from a body at a higher temperature to one at a lower temperature.

Units: The SI unit of heat is the *Joule (J)*. In the context of heat transfer, *calories (cal)* or *kilocalories (kcal)* may also be used. The conversion is:

$$1\,\text{cal} = 4.184\,\text{J}$$

$$1\,\text{kcal} = 4184\,\text{J}$$

Heat Capacity and Specific Heat

Heat Capacity

Heat capacity is **the amount of heat energy** required to change the temperature of an object or system by **one degree Celsius** (or one Kelvin). The formula for heat capacity is:

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

where:

- C = Heat capacity (in joules per degree Celsius or Kelvin),
- Q = Heat energy absorbed or released (in joules),
- ΔT = Change in temperature (in °C or K).

Specific Heat Capacity

Specific heat capacity is the heat required to raise the temperature of ****1 unit of mass**** of a substance by **one degree Celsius** (or one Kelvin). The formula for specific heat is:

$$c = \frac{Q}{m\Delta T}$$

where:

- c = Specific heat capacity (in joules per kilogram per degree Celsius or Kelvin),
- m = Mass of the substance (in kilograms),
- ΔT = Change in temperature (in °C or K),
- Q = Heat energy absorbed or released (in joules).

Heat of Fusion and Heat of Vaporization

Heat of Fusion L_F

The heat of fusion (L_F) is the amount of heat energy required to **convert **1 unit of mass**** of a **solid into a liquid at its melting point**, without changing its temperature. The formula for heat required during melting (fusion) is:

$$Q = mL_F$$

where:

- L_F = Heat of fusion (in joules per kilogram),
- m = Mass of the substance (in kilograms).

For water, the heat of fusion is:

$$L_F = 334 \times 10^3 \text{ J/kg}$$

Heat of Vaporization L_V

The heat of vaporization (L_V) is the amount of heat energy required to **convert **1 unit of mass**** of **a liquid into a gas at its boiling point**, without changing its temperature. The formula for heat required during vaporization is:

$$Q = mL_V$$

where:

- L_V = Heat of vaporization (in joules per kilogram),
- m = Mass of the substance (in kilograms).

For water, the heat of vaporization is:

$$L_V = 2260 \times 10^3 \text{ J/kg}$$

Problem 1

How much heat must be absorbed by ice of mass $m = 720\,\text{g}$ at $10°C$ to take it to the liquid state at $15°C$?

Solution

We need to calculate the heat absorbed in three steps:

1. Heat required to raise the temperature of ice from $10°C$ to $0°C$,
2. Heat required to melt the ice at $0°C$ into water, then

3. Heat required to raise the temperature of the resulting water from $0°C$ to $15°C$.

Step 1: Heat required to raise the temperature of ice:

$$Q_1 = mc_{\text{ice}}\Delta T$$

Where:

$$m = 720\,\text{g} = 0.72\,\text{kg}, \quad c_{\text{ice}} = 2100\,\text{J/kg}^{\circ}C, \quad \Delta T = 10°C = 10\,\text{K}$$

$$Q_1 = 0.72 \times 2100 \times 10 = 15120\,\text{J}$$

Step 2: Heat required to melt the ice:

$$Q_2 = mL_f$$

Where:

$$L_f = 334 \times 10^3\,\text{J/kg}$$

$$Q_2 = 0.72 \times 334 \times 10^3 = 240480\,\text{J}$$

Step 3: Heat required to raise the temperature of water from $0°C$ to $15°C$:

$$Q_3 = mc_{\text{water}}\Delta T$$

Where:

$$c_{\text{water}} = 4186\,\text{J/kg}^{\circ}C, \quad \Delta T = 15°C = 15\,\text{K}$$

$$Q_3 = 0.72 \times 4186 \times 15 = 45201.6\,\text{J}$$

Total heat absorbed:

$$Q_{\text{total}} = Q_1 + Q_2 + Q_3 = 15120 + 240480 + 45201.6 = 300801.6\,\text{J}$$

Thus, the total heat absorbed is 300801.6 J.

Problem 2

A certain substance has a mass per mole of 50.0 g/mol. When 314 J is added as heat to a 30.0 g sample, the sample's temperature rises from $25.0°C$ to $45.0°C$. What are the (a) specific heat and (b) How many moles are in the sample?

Solution

We are given:

$$Q = 314\,\mathrm{J}, \quad m = 30.0\,\mathrm{g} = 0.030\,\mathrm{kg}, \quad \Delta T = (45.0 - 25.0) = 20.0°C = 20.0\,\mathrm{K}$$

(a) To find the specific heat, use the formula:

$$c = \frac{Q}{m\Delta T}$$

Substituting the values:

$$c = \frac{314}{0.030 \times 20.0} = \frac{314}{0.600} = 523.33\,\mathrm{J/kg}K$$

(b)

$$\text{moles} = \frac{30.0\,\mathrm{g}}{50.0\,\mathrm{g/mol}} = 0.600\,\mathrm{mol}$$

The number of moles is $0.600\,\mathrm{mol}$.

First Law of Thermodynamics

The **First Law of Thermodynamics**, also known as **the law of energy conservation**, states that energy can neither be created nor destroyed, but only transferred or converted from one form to another.

In the context of thermodynamics, this means that ,*The change in the internal energy of a system is equal to the amount of heat added to the system minus the work done by the system on its surroundings.*

Mathematical Expression

The First Law of Thermodynamics is mathematically expressed as:

$$\Delta U = Q - W$$

where:

- ΔU is the change in internal energy of the system,
- Q is the heat added to the system (positive if heat is added, negative if heat is removed),
- W is the work done by the system (positive if work is done by the system on the surroundings, negative if work is done on the system).

Sign Convention

- $Q > 0$ indicates heat is added to the system. - $Q < 0$ indicates heat is removed from the system. - $W > 0$ indicates the system does work on the surroundings. - $W < 0$ indicates work is done on the system.

Isothermal Changes

An *isothermal process* is a thermodynamic process in which **the temperature of the system remains constant**. In other words, the temperature does not change during the process, i.e. $\Delta T = 0$.

Key Characteristics of Isothermal Processes

- **Constant Temperature**: The temperature of the system remains unchanged throughout the process.
- **Heat Exchange**: Since the temperature is constant, the internal energy of an ideal gas remains constant. As a result, any work done by or on the system must be compensated by heat exchange with the surroundings.
- **First Law of Thermodynamics**: For an isothermal process, the first law of thermodynamics states:

  $$Q = W$$

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

Example

Consider an ideal gas in a piston. If the gas undergoes an isothermal expansion or compression, it will exchange heat with the surroundings to maintain a constant temperature while doing work.

Adiabatic Changes

An *adiabatic process* is a thermodynamic process in which **no heat is exchanged** between the system and the surroundings, i.e., $Q = 0$. In an adiabatic process, any change in the internal energy of the system is solely due to work done by or on the system.

Key Characteristics of Adiabatic Processes

- **No Heat Transfer**: There is no heat exchange with the surroundings. Any change in internal energy is due to work done by or on the system.
- **Work and Internal Energy**: Since there is no heat flow, the internal energy change is given by:

  $$\Delta U = -W$$

  where W is the work done by the system. If the system does work (expansion), the internal energy decreases, and if work is done on the system (compression), the internal energy increases.

Example

When a gas is rapidly compressed in a piston with no heat exchange (e.g., in an insulated piston), the temperature of the gas increases as work is done on it.

Work Associated with Volume Change

A gas may exchange energy with its surroundings through work. The amount of work W done by a gas as it expands or contracts from an initial volume V_i to a final volume V_f is given by:

$$W = \int_{V_i}^{V_f} p \, dV. \tag{1}$$

Here, p represents the pressure of the gas, which may vary with volume.

Problem 3:

1.00 kg of liquid water at 100° C is converted to steam at 100° C by boiling at standard atmospheric pressure (1.00 atm or 1.01×10^5 Pa). The volume of the water changes from an initial value of $1.00 \times 10^{-3}\,\mathrm{m}^3$ as a liquid to $1.671\,\mathrm{m}^3$ as steam.

(a) How much work is done by the system during this process?

Given Data

- Mass of water, $m = 1.00\,\text{kg}$
- Initial volume of liquid water, $V_{\text{initial}} = 1.00 \times 10^{-3}\,\text{m}^3$
- Final volume of steam, $V_{\text{final}} = 1.671\,\text{m}^3$
- Atmospheric pressure, $P = 1.01 \times 10^5\,\text{Pa}$
- The temperature remains constant at 100° C.

Solution

The work done by the system is given by the formula:

$$W = P\Delta V$$

where:

- W is the work done by the system,
- P is the pressure of the system (which is constant during the process),
- $\Delta V = V_{\text{final}} - V_{\text{initial}}$ is the change in volume of the system.

Substituting the given values:

$$\Delta V = 1.671\,\text{m}^3 - 1.00 \times 10^{-3}\,\text{m}^3$$
$$\Delta V = 1.670\,\text{m}^3$$

Now, the work done by the system is:

$$W = (1.01 \times 10^5\,\text{Pa}) \times (1.670\,\text{m}^3)$$

$$W = 1.6847 \times 10^5\,\text{J}$$

Thus, the work done by the system during the process is:

$$\boxed{W = 1.685 \times 10^5\,\text{J}}$$

(b) How much energy is transferred as heat during the process?

Given Data

- Mass of water, $m = 1.00\,\text{kg}$
- Latent heat of vaporization of water, $L_V = 2.25 \times 10^6\,\text{J/kg}$
- The temperature remains constant at 100° C.

Solution

To calculate the energy transferred as heat during the process, we can use the formula for latent heat:

$$Q = mL_V$$

where:

- Q is the energy transferred as heat,
- m is the mass of the water,
- L_V is the latent heat of vaporization of water.

Substituting the given values:

$$Q = (1.00\,\text{kg}) \times (2.25 \times 10^6\ \text{J/kg})$$

$$Q = 2.25 \times 10^6\,\text{J}$$

Thus, the energy transferred as heat during the process is:

$$\boxed{Q = 2.25 \times 10^6\,\text{J}}$$

(c) What is the change in the system's internal energy during the process?

Given Data

- Mass of water, $m = 1.00\,\text{kg}$
- Latent heat of vaporization of water, $L_V = 2.25 \times 10^6\ \text{J/kg}$
- The temperature remains constant at 100° C, so the process is isothermal.
- Energy transferred as heat, $Q = 2.25 \times 10^6$ J (calculated in part (b)).
- The work done by the system, $W = 1.685 \times 10^5$ J (calculated in part (a)).

Solution

The change in the system's internal energy is related to the heat added to the system and the work done by the system during the process by the first law of thermodynamics:

$$\Delta U = Q - W$$

where:

- ΔU is the change in internal energy,
- Q is the heat added to the system (energy transferred as heat),
- W is the work done by the system.

Substituting the given values:

$$\Delta U = (2.25 \times 10^6 \text{ J}) - (1.685 \times 10^5 \text{ J})$$

$$\Delta U = 2.0815 \times 10^6 \text{ J}$$

Thus, the change in the system's internal energy is:

$$\boxed{\Delta U = 2.0815 \times 10^6 \text{ J}}$$

Reversible and Irreversible Processes

Reversible Process: A reversible process is an ideal process that can be reversed without leaving any net change in the system or surroundings. It is infinitely slow, maintains equilibrium, and generates no entropy ($\Delta S_{\text{total}} = 0$). Examples include isothermal expansion and melting under equilibrium conditions.

Irreversible Process: An irreversible process occurs naturally and cannot be reversed without changes to the system or surroundings. It involves finite rates, disequilibrium, and entropy generation ($\Delta S_{\text{total}} > 0$). Examples include heat transfer across a temperature gradient, friction, and gas mixing.