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CHAPTER 10

THERMAL PROPERTIES OF MATTER

Complete Study Material with Concepts, Formulas, Derivations & Examples

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1. INTRODUCTION

We all have common sense notions of heat and temperature. Temperature is a measure of 'hotness' of a body. A kettle with boiling water is hotter than a box containing ice. In physics, we need to define the notion of heat, temperature, etc., more carefully.

Key Concepts:

- **Temperature:** A relative measure or indication of hotness or coldness
- **Heat:** Form of energy transferred between systems due to temperature difference

- **Thermal Equilibrium:** When two systems reach the same temperature and no heat transfer occurs
- **Hot and Cold:** Relative terms, like tall and short

📝 **Important Observations:**

- Glass of ice-cold water left on table eventually warms up
- Cup of hot tea on table cools down
- Heat transfer continues until body and surroundings reach same temperature
- For ice-cold water: Heat flows from environment to glass
- For hot tea: Heat flows from cup to environment

2. TEMPERATURE AND HEAT

Definitions:

Heat: Form of energy transferred between two (or more) systems or a system and its surroundings by virtue of temperature difference

SI Unit of Heat: Joule (J)

SI Unit of Temperature: Kelvin (K)

Common Unit: Degree Celsius (°C)

2.1 Effects of Heat on Objects

When an object is heated, many changes may take place:

- Its **temperature may rise**
- It may **expand**
- It may **change state** (solid → liquid → gas)

3. MEASUREMENT OF TEMPERATURE

3.1 Thermometer

A measure of temperature is obtained using a **thermometer**. Many physical properties of materials change sufficiently with temperature and are used as the basis for constructing thermometers.

> 📝 **Common Property Used:**
>
> Variation of the volume of a liquid with temperature. For example:
>
> - **Mercury thermometers:** Volume of mercury varies linearly with temperature
> - **Alcohol thermometers:** Volume of alcohol varies linearly with temperature

3.2 Temperature Scales

Thermometers are calibrated so that a numerical value may be assigned to a given temperature. For defining any standard scale, two fixed reference points are needed:

> **Fixed Points:**
>
> - **Ice Point (Freezing Point):** Temperature at which pure water freezes under standard pressure

- **Steam Point (Boiling Point):** Temperature at which pure water boils under standard pressure

Scale	Ice Point	Steam Point	Number of Divisions
Fahrenheit (°F)	32 °F	212 °F	180
Celsius (°C)	0 °C	100 °C	100

Conversion Formula:

$$(t_F - 32)/180 = t_C/100$$

Or simplified:

$$t_F = (9/5)t_C + 32$$

4. IDEAL-GAS EQUATION AND ABSOLUTE TEMPERATURE

4.1 Gas Thermometer

Liquid-in-glass thermometers show different readings for temperatures other than the fixed points because of differing expansion properties. A thermometer that uses a gas, however, **gives the same readings regardless of which gas is used**.

Gas Laws:

- **Boyle's Law:** At constant temperature, PV = constant
- **Charles' Law:** At constant pressure, V/T = constant

- All low-density gases exhibit same expansion behaviour

4.2 Ideal Gas Equation

Ideal Gas Equation:

$$PV = \mu RT$$

Where:

- P = Pressure
- V = Volume
- μ = Number of moles
- R = Universal gas constant = 8.31 J mol^{-1} K^{-1}
- T = Absolute temperature in Kelvin

4.3 Absolute Zero and Kelvin Scale

At constant volume, pressure is directly proportional to temperature: $P \propto T$

When we extrapolate the straight line P vs T to the point where P = 0, we get the **absolute minimum temperature**.

Absolute Zero:

Value: -273.15 °C = 0 K

Kelvin Scale (Absolute Scale):

$$T = t_C + 273.15$$

Where:

- T = Temperature in Kelvin (K)

- t_C = Temperature in Celsius (°C)

Note: Size of unit in Kelvin and Celsius scales is the same

Point	Kelvin (K)	Celsius (°C)	Fahrenheit (°F)
Absolute Zero	0 K	-273.15 °C	-459.69 °F
Ice Point	273.15 K	0 °C	32 °F
Steam Point	373.15 K	100 °C	212 °F

5. THERMAL EXPANSION

It is our common experience that most substances **expand on heating and contract on cooling**. A change in the temperature of a body causes change in its dimensions.

📝 **Examples of Thermal Expansion:**

- Metallic lid on sealed bottle loosens when put in hot water
- Mercury in thermometer rises when placed in warm water
- Balloon partially inflated in cool room expands in warm water
- Fully inflated balloon shrinks when immersed in cold water

5.1 Types of Thermal Expansion

Three Types:

- **Linear Expansion:** Expansion in length
- **Area Expansion:** Expansion in area
- **Volume Expansion:** Expansion in volume

5.2 Coefficient of Linear Expansion (α_l)

Definition:

$$\Delta l/l = \alpha_l \Delta T$$

Where:

- Δl = Change in length
- l = Original length
- ΔT = Change in temperature
- α_l = Coefficient of linear expansion

SI Unit: K^{-1} or $(°C)^{-1}$

Dimensions: $[\Theta^{-1}]$ where Θ represents temperature

Material	α_l (10^{-5} K^{-1})
Aluminium	2.5
Brass	1.8
Iron	1.2
Copper	1.7
Silver	1.9

Material	α_l (10^{-5} K^{-1})
Gold	1.4
Glass (pyrex)	0.32
Lead	0.29

📝 **Observations:**

- Metals expand more and have relatively high values of α_l
- Copper expands about 5 times more than glass for same temperature rise
- This explains why blacksmiths heat iron rings before fitting on wooden wheels

5.3 Coefficient of Area Expansion (α_A)

Definition:

$$\Delta A/A = \alpha_A \Delta T = 2\alpha_l \Delta T$$

Where:

- ΔA = Change in area
- A = Original area

5.4 Coefficient of Volume Expansion (α_V)

Definition:

$$\Delta V/V = \alpha_V \Delta T$$

Relationship:

$$\alpha_V = 3\alpha_l$$

For solids that expand uniformly in all directions

Material	α_V (K^{-1})
Aluminium	7×10^{-5}
Brass	6×10^{-5}
Iron	3.55×10^{-5}
Paraffin	58.8×10^{-5}
Glass (ordinary)	2.5×10^{-5}
Glass (pyrex)	1×10^{-5}
Hard rubber	2.4×10^{-4}
Invar	2×10^{-6}
Mercury	18.2×10^{-5}
Water	20.7×10^{-5}
Alcohol (ethanol)	110×10^{-5}

📝 **Key Points:**

- Thermal expansion of solids and liquids is rather small
- Pyrex glass and invar have particularly low values of α_V
- Alcohol expands more than mercury for same temperature rise

5.5 Anomalous Behavior of Water

⚠️ **Special Case:**

Water contracts on heating between 0 °C and 4 °C!

- Volume decreases as water is cooled from room temperature until 4 °C
- Below 4 °C, volume increases (density decreases)
- **Water has maximum density at 4 °C**

Environmental Importance:

This property has important environmental effects:

- Bodies of water (lakes, ponds) freeze at the top first
- As lake cools toward 4 °C, surface water becomes denser and sinks
- Warmer, less dense water near bottom rises
- Below 4 °C, colder water becomes less dense and remains at surface
- Surface freezes while bottom remains at 4 °C
- This allows aquatic life to survive in winter

5.6 Thermal Expansion of Gases

Gases, at ordinary temperature, expand more than solids and liquids.

For Ideal Gas at Constant Pressure:

From ideal gas equation: $PV = \mu RT$

At constant pressure: $P\Delta V = \mu R\Delta T$

$$\alpha_V = \Delta V/(V\Delta T) = 1/T$$

At 0 °C: $\alpha_V = 3.7 \times 10^{-3}\ K^{-1}$

Much larger than solids and liquids!

At room temperature: $\alpha_V \approx 3300 \times 10^{-6}\ K^{-1}$

5.7 Thermal Stress

What happens if we prevent the thermal expansion of a rod by fixing its ends rigidly?

Thermal Stress:

When thermal expansion is prevented, the rod acquires a **compressive strain** due to external forces provided by rigid support. The corresponding stress is called **thermal stress**.

📘 **Example - Steel Rail:**

Steel rail: Length = 5 m, Area = 40 cm^2, ΔT = 10 °C

α_l(steel) = 1.2×10^{-5} K^{-1}

Y(steel) = 2×10^{11} N m^{-2}

Solution:

Compressive strain = $\Delta l/l = \alpha_l \Delta T = 1.2 \times 10^{-5} \times 10 = 1.2 \times 10^{-4}$

Thermal stress = Y × strain = $2 \times 10^{11} \times 1.2 \times 10^{-4}$

= 2.4×10^7 N m^{-2}

External force = Stress × Area = $2.4 \times 10^7 \times 40 \times 10^{-4}$

≈ 10^5 N

This large force can easily bend the rails!

6. SPECIFIC HEAT CAPACITY

6.1 Experimental Observations

Simple experiments show that the quantity of heat required to raise temperature of a substance depends on:

- **Mass of substance (m):** Double the mass → Double the heat required
- **Change in temperature (ΔT):** Double the temperature rise → Double the heat required
- **Nature of substance:** Different substances require different amounts of heat for same mass and temperature change

6.2 Heat Capacity

Heat Capacity (S):

$$S = \Delta Q/\Delta T$$

Where:

- ΔQ = Amount of heat supplied
- ΔT = Change in temperature

SI Unit: J K^{-1}

6.3 Specific Heat Capacity

Specific Heat Capacity (s):

$$s = S/m = (1/m)(\Delta Q/\Delta T)$$

Or rearranging:

$$\Delta Q = ms\Delta T$$

Where:

- m = Mass of substance
- s = Specific heat capacity
- ΔT = Change in temperature

Definition: Amount of heat per unit mass required to change temperature by one unit

SI Unit: J kg^{-1} K^{-1}

Substance	Specific Heat Capacity (J kg^{-1} K^{-1})
Water	4186.0
Ice	2060

Substance	Specific Heat Capacity ($J\ kg^{-1}\ K^{-1}$)
Aluminium	900.0
Glass	840
Carbon	506.5
Iron	450
Copper	386.4
Silver	236.1
Lead	127.7
Mercury	140
Kerosene	2118
Edible oil	1965

🔑 **Key Observations:**

- **Water has the highest specific heat capacity** among common substances
- This is why water is used as coolant in automobile radiators
- Water is also used in hot water bags due to high heat retention
- Water warms up more slowly than land during summer (sea breeze)
- Desert earth warms up quickly during day and cools quickly at night

6.4 Molar Specific Heat Capacity

Molar Specific Heat Capacity (C):

$$C = S/\mu = (1/\mu)(\Delta Q/\Delta T)$$

Where:

- μ = Number of moles
- C = Molar specific heat capacity

SI Unit: J mol^{-1} K^{-1}

6.5 For Gases: C_p and C_v

📝 **Two Types for Gases:**

- **C_p:** Molar specific heat capacity at constant pressure
- **C_v:** Molar specific heat capacity at constant volume
- For gases: $C_p > C_v$

Gas	C_p (J mol^{-1} K^{-1})	C_v (J mol^{-1} K^{-1})
Helium (He)	20.8	12.5
Hydrogen (H_2)	28.8	20.4
Nitrogen (N_2)	29.1	20.8
Oxygen (O_2)	29.4	21.1
Carbon dioxide (CO_2)	37.0	28.5

7. CALORIMETRY

Calorimetry means measurement of heat. When a body at higher temperature is brought in contact with another body at lower temperature, heat is transferred until thermal equilibrium is reached.

> **Principle of Calorimetry:**
>
> **Heat lost by hot body = Heat gained by cold body**
>
> (Assuming no heat is lost to surroundings)

7.1 Calorimeter

A device in which heat measurement can be done. It consists of:

- Metallic vessel (usually copper or aluminium)
- Stirrer of same material
- Wooden jacket with heat insulating material (glass wool)
- Opening for mercury thermometer

📘 Example 1 - Specific Heat of Aluminium:

Problem: A sphere of 0.047 kg aluminium is placed in boiling water (100 °C), then immediately transferred to 0.14 kg copper calorimeter containing 0.25 kg water at 20 °C. Final temperature is 23 °C. Calculate specific heat capacity of aluminium.

Given:

m_{Al} = 0.047 kg, $T_{initial,Al}$ = 100 °C, T_{final} = 23 °C

m_{water} = 0.25 kg, m_{cal} = 0.14 kg, $T_{initial,water}$ = 20 °C

$s_{water} = 4.18 \times 10^3$ J kg^{-1} K^{-1}

$s_{Cu} = 0.386 \times 10^3$ J kg^{-1} K^{-1}

Solution:

Heat lost by aluminium:

$Q_{lost} = m_{Al} \times s_{Al} \times \Delta T_{Al}$

$= 0.047 \times s_{Al} \times (100 - 23)$

$= 0.047 \times s_{Al} \times 77$

Heat gained by water + calorimeter:

$Q_{gained} = (m_{water} \times s_{water} + m_{cal} \times s_{Cu}) \times \Delta T$

$= (0.25 \times 4.18 \times 10^3 + 0.14 \times 0.386 \times 10^3) \times (23 - 20)$

$= (1045 + 54.04) \times 3$

$= 3297.12$ J

By principle of calorimetry:

$0.047 \times s_{Al} \times 77 = 3297.12$

$s_{Al} = 911$ J kg^{-1} K^{-1} ≈ 0.91 kJ kg^{-1} K^{-1}

8. CHANGE OF STATE

Matter normally exists in three states: solid, liquid and gas. A transition from one state to another is called a **change of state**.

8.1 Common Changes of State

Types of Phase Changes:

- **Melting (Fusion):** Solid → Liquid
- **Freezing:** Liquid → Solid
- **Vaporization (Boiling):** Liquid → Gas
- **Condensation:** Gas → Liquid
- **Sublimation:** Solid → Gas (directly, without liquid phase)

8.2 Melting Point

📝 **Observations During Melting:**

- Temperature remains constant during melting
- Solid and liquid states coexist in thermal equilibrium
- Heat added does not increase temperature but changes state
- Melting point depends on pressure

Melting Point:

Definition: Temperature at which solid and liquid states of substance are in thermal equilibrium with each other

Normal Melting Point: Melting point at standard atmospheric pressure (1 atm)

8.3 Boiling Point

📝 **Observations During Boiling:**

- Temperature remains constant during boiling
- Liquid and vapour states coexist in thermal equilibrium
- Heat added converts liquid to vapour without temperature rise
- Boiling point increases with increase in pressure
- Boiling point decreases with decrease in pressure

Boiling Point:

Definition: Temperature at which liquid and vapour states of substance coexist in thermal equilibrium

Normal Boiling Point: Boiling point at standard atmospheric pressure (1 atm)

Applications:

- **Cooking on hills:** Difficult because atmospheric pressure is lower → lower boiling point of water
- **Pressure cooker:** Increased pressure → higher boiling point → faster cooking

8.4 Regelation

📝 **Regelation:**

The phenomenon of refreezing. When pressure is increased on ice, it melts at lower temperature. When pressure is removed, water refreezes.

Example: Wire passing through ice block without splitting it

> **Application:** Skating on ice - pressure under skates forms thin water layer acting as lubricant

8.5 Sublimation

Some substances pass directly from solid to vapour state without passing through liquid state. Examples:

- Dry ice (solid CO_2)
- Iodine
- Camphor
- Naphthalene

8.6 Latent Heat

The amount of heat transferred during change of state without change in temperature is called **latent heat**.

Latent Heat (L):

$$Q = mL$$

or

$$L = Q/m$$

Where:

- Q = Heat required
- m = Mass of substance
- L = Latent heat (characteristic of substance)

SI Unit: J kg^{-1}

Types of Latent Heat:

1. Latent Heat of Fusion (L_f):

Heat per unit mass required to change substance from solid to liquid at same temperature and pressure

For water: $L_f = 3.33 \times 10^5$ J kg^{-1}

This is heat needed to melt 1 kg ice at 0 °C

2. Latent Heat of Vaporization (L_v):

Heat per unit mass required to change substance from liquid to vapour at same temperature and pressure

For water: $L_v = 22.6 \times 10^5$ J kg^{-1}

This is heat needed to convert 1 kg water to steam at 100 °C

Substance	Melting Point (°C)	L_f (10^5 J kg^{-1})	Boiling Point (°C)	L_v (10^5 J kg^{-1})
Ethanol	-114	1.0	78	8.5
Gold	1063	0.645	2660	15.8
Lead	328	0.25	1744	8.67
Mercury	-39	0.12	357	2.7
Nitrogen	-210	0.26	-196	2.0
Oxygen	-219	0.14	-183	2.1

Substance	Melting Point (°C)	L_f (10^5 J kg^{-1})	Boiling Point (°C)	L_v (10^5 J kg^{-1})
Water	0	3.33	100	22.6

⚠️ **Important:**

Steam at 100 °C carries 22.6 × 10^5 J kg^{-1} more heat than water at 100 °C.

This is why burns from steam are usually more serious than those from boiling water!

📘 Example 2 - Ice Melting:

Problem: When 0.15 kg ice at 0 °C is mixed with 0.30 kg water at 50 °C, the resulting temperature is 6.7 °C. Calculate heat of fusion of ice. (s_{water} = 4186 J kg^{-1} K^{-1})

Solution:

Heat lost by water:

$Q_{lost} = m_w \times s_w \times \Delta T$

= 0.30 × 4186 × (50.0 - 6.7)

= 54376.14 J

Heat gained by ice = Heat to melt ice + Heat to warm ice water

$Q_{gained} = m_i \times L_f + m_i \times s_w \times \Delta T$

= 0.15 × L_f + 0.15 × 4186 × (6.7 - 0)

= 0.15 × L_f + 4206.93

By principle of calorimetry:

$54376.14 = 0.15 \times L_f + 4206.93$

$L_f = 3.34 \times 10^5$ J kg^{-1}

📘 Example 3 - Complete Phase Change:

Problem: Calculate heat required to convert 3 kg ice at -12 °C to steam at 100 °C at atmospheric pressure.

Given: $s_{ice} = 2100$ J kg^{-1} K^{-1}, $s_{water} = 4186$ J kg^{-1} K^{-1}

$L_f = 3.35 \times 10^5$ J kg^{-1}, $L_v = 2.256 \times 10^6$ J kg^{-1}

Solution:

Step 1: Ice at -12 °C to ice at 0 °C:

$Q_1 = m \times s_{ice} \times \Delta T = 3 \times 2100 \times 12 = 75{,}600$ J

Step 2: Ice at 0 °C to water at 0 °C:

$Q_2 = m \times L_f = 3 \times 3.35 \times 10^5 = 1{,}005{,}000$ J

Step 3: Water at 0 °C to water at 100 °C:

$Q_3 = m \times s_{water} \times \Delta T = 3 \times 4186 \times 100 = 1{,}255{,}800$ J

Step 4: Water at 100 °C to steam at 100 °C:

$Q_4 = m \times L_v = 3 \times 2.256 \times 10^6 = 6{,}768{,}000$ J

Total heat required:

$Q = Q_1 + Q_2 + Q_3 + Q_4$

$= 75{,}600 + 1{,}005{,}000 + 1{,}255{,}800 + 6{,}768{,}000$

= 9.1×10^6 J

9. HEAT TRANSFER

Heat is energy transfer from one system to another arising due to temperature difference. There are **three distinct modes of heat transfer**:

Three Modes of Heat Transfer:

1. **Conduction:** Heat transfer through material without actual motion of material
2. **Convection:** Heat transfer by actual motion of matter (only in fluids)
3. **Radiation:** Heat transfer by electromagnetic waves (no medium needed)

9.1 Conduction

Conduction is the mechanism of heat transfer between two adjacent parts of a body because of their temperature difference.

📝 Example:

One end of metallic rod is put in a flame, the other end soon becomes so hot you cannot hold it. Heat transfer takes place by conduction from hot end through different parts to the other end.

Thermal Conductivity:

Consider a metallic bar of length L and uniform cross-section A with ends at temperatures T_C and T_D ($T_C > T_D$).

Fourier's Law of Heat Conduction:

$$H = KA(T_C - T_D)/L$$

Where:

- H = Rate of heat flow (heat current) in J s^{-1} or W
- K = Thermal conductivity of material
- A = Cross-sectional area
- $T_C - T_D$ = Temperature difference
- L = Length of bar

SI Unit of K: J s^{-1} m^{-1} K^{-1} or W m^{-1} K^{-1}

Material	Thermal Conductivity K (J s^{-1} m^{-1} K^{-1})
Metals (Good Conductors)	
Silver	406
Copper	385
Aluminium	205
Brass	109
Steel	50.2
Lead	34.7
Mercury	8.3
Non-metals (Poor Conductors)	
Water	0.8

Material	Thermal Conductivity K (J s^{-1} m^{-1} K^{-1})
Concrete	0.8
Glass	0.8
Ice	1.6
Insulating brick	0.15
Body fat	0.20
Wood	0.12
Glass wool	0.04
Felt	0.04
Gases (Very Poor Conductors)	
Air	0.024
Argon	0.016
Hydrogen	0.14

📝 Applications:

- **Cooking pots:** Copper coating on bottom for uniform heat distribution
- **Building insulation:** Foam or earth layer on ceiling to prevent heat transfer
- **Nuclear reactors:** Elaborate heat transfer systems to prevent overheating
- **Plastic foams:** Good insulators due to air pockets

📘 Example 4 - Steel-Copper Junction:

Problem: Steel rod (length 15 cm) and copper rod (length 10 cm) are joined. Steel area = 2 × Copper area. Furnace at 300 °C, other end at 0 °C. Find junction temperature.

K_{steel} = 50.2 J s^{-1} m^{-1} K^{-1}, K_{copper} = 385 J s^{-1} m^{-1} K^{-1}

Solution:

In steady state, heat flow through both rods is same:

$K_1A_1(300 - T)/L_1 = K_2A_2(T - 0)/L_2$

For $A_1 = 2A_2$:

$(50.2)(2A_2)(300 - T)/0.15 = (385)(A_2)(T)/0.10$

Simplifying:

670.67(300 - T) = 3850T

201,200 - 670.67T = 3850T

201,200 = 4520.67T

T = 44.4 °C

📘 Example 5 - Compound Bar:

Problem: Iron bar (L_1 = 0.1 m, A_1 = 0.02 m^2, K_1 = 79 W m^{-1} K^{-1}) and brass bar (L_2 = 0.1 m, A_2 = 0.02 m^2, K_2 = 109 W m^{-1} K^{-1}) soldered end to end. Ends maintained at 373 K and 273 K. Find (i) junction temperature, (ii) equivalent thermal conductivity, (iii) heat current.

Solution:

(i) Junction temperature T_0:

$T_0 = (K_1T_1 + K_2T_2)/(K_1 + K_2)$

$= (79 \times 373 + 109 \times 273)/(79 + 109)$

$= (29467 + 29757)/188$

= 315 K

(ii) Equivalent thermal conductivity K':

$K' = 2K_1K_2/(K_1 + K_2)$

$= 2 \times 79 \times 109/(79 + 109)$

= 91.6 W m^{-1} K^{-1}

(iii) Heat current H:

$H = K'A(T_1 - T_2)/(2L)$

$= (91.6)(0.02)(373 - 273)/(2 \times 0.1)$

= 916.1 W

9.2 Convection

Convection is a mode of heat transfer by actual motion of matter. It is possible only in fluids.

Types of Convection:

1. Natural Convection:

Gravity plays important role. When fluid is heated from below:

- Hot part expands → becomes less dense
- Due to buoyancy, it rises
- Upper colder part replaces it

- This gets heated, rises, and process continues

2. Forced Convection:

Material is forced to move by pump or other physical means. Examples:

- Forced-air heating systems in homes
- Human circulatory system (heart acts as pump)
- Automobile engine cooling system

Examples of Natural Convection:

Sea and Land Breeze:

During Day:

- Ground heats up more quickly than water
- Air in contact with warm ground heats by conduction
- Warm air expands, becomes less dense, rises
- Cooler air from sea moves in → **Sea breeze**

At Night:

- Ground loses heat more quickly
- Water surface warmer than land
- Cycle reverses → **Land breeze**

Trade Winds:

Steady surface wind on Earth blowing from north-east toward equator:

- Equatorial and polar regions receive unequal solar heat
- Air at equatorial surface is hot, rises
- Air in upper atmosphere at poles is cool
- Earth's rotation modifies this convection current
- Air descends at 30°N latitude and returns to equator

9.3 Radiation

Radiation is heat transfer by electromagnetic waves. It needs no medium and travels at speed of light (3×10^8 m s^{-1}).

Key Features of Radiation:

- Does not need any medium (can occur in vacuum)
- Travels at speed of light
- All bodies emit radiant energy (solid, liquid, gas)
- Radiation emitted by virtue of temperature is called **thermal radiation**
- Amount absorbed depends on color of body

📝 Applications in Daily Life:

- **Summer clothing:** White/light colored (absorb least heat)
- **Winter clothing:** Dark colored (absorb more heat)
- **Cooking utensils:** Blackened bottoms absorb maximum heat
- **Thermos flask:** Double-walled with silvered surfaces to minimize radiation

Thermos Flask (Dewar Flask):

Device to minimize heat transfer between contents and outside:

- Double-walled glass vessel
- Inner and outer walls coated with silver
- Radiation from inner wall reflected back
- Space between walls evacuated (reduces conduction/convection)
- Flask supported on insulator (cork)

9.4 Blackbody Radiation

Thermal radiation at any temperature has continuous spectrum from small to long wavelengths. Energy content varies for different wavelengths.

Wien's Displacement Law:

$$\lambda_m T = \text{constant}$$

Wien's constant: 2.9×10^{-3} m K

Where λ_m = wavelength for which energy is maximum

📝 **Applications:**

- Color of heated iron: Dull red → Reddish yellow → White hot (as temperature increases)
- Estimating surface temperatures of celestial bodies
- Moon: λ_m = 14 μm → T = 200 K
- Sun: λ_m = 4753 Å → T = 6060 K

Stefan-Boltzmann Law:

For Perfect Radiator (Blackbody):

$$H = A\sigma T^4$$

Where:

- H = Energy radiated per unit time (W)
- A = Surface area (m^2)
- σ = Stefan-Boltzmann constant = 5.67×10^{-8} W m^{-2} K^{-4}
- T = Absolute temperature (K)

For Real Bodies (with emissivity e):

$$H = Ae\sigma T^4$$

Where e = emissivity ($0 \leq e \leq 1$)

e = 1 for perfect radiator (blackbody)

e ≈ 0.4 for tungsten lamp

Net Rate of Heat Loss/Gain:

When body at temperature T is surrounded by environment at T_s:

$$H = e\sigma A(T^4 - T_s^4)$$

Body emits as well as receives energy

📘 Example 6 - Human Body Radiation:

Problem: Surface area of person's body = 1.9 m^2, room temperature = 22 °C = 295 K, skin temperature = 28 °C = 301 K, emissivity of skin = 0.97. Calculate rate of heat loss.

Solution:

$H = e\sigma A(T^4 - T_s^4)$

$= 0.97 \times 5.67 \times 10^{-8} \times 1.9 \times [(301)^4 - (295)^4]$

$= 0.97 \times 5.67 \times 10^{-8} \times 1.9 \times [8.191 \times 10^9 - 7.566 \times 10^9]$

$= 0.97 \times 5.67 \times 10^{-8} \times 1.9 \times 6.25 \times 10^8$

≈ 66.4 W

More than half the body's resting energy production (120 W)!

📘 Example 7 - Tungsten Lamp:

Problem: Tungsten lamp at T = 3000 K, surface area = 0.3 cm^2 = 0.3×10^{-4} m^2, emissivity e = 0.4. Calculate rate of radiation.

Solution:

$H = Ae\sigma T^4$

$= 0.3 \times 10^{-4} \times 0.4 \times 5.67 \times 10^{-8} \times (3000)^4$

$= 0.3 \times 10^{-4} \times 0.4 \times 5.67 \times 10^{-8} \times 8.1 \times 10^{13}$

≈ 60 W

10. NEWTON'S LAW OF COOLING

Newton was the first to study systematically the relation between heat lost by a body and its temperature.

Newton's Law of Cooling:

Statement: The rate of loss of heat of a body is directly proportional to the difference of temperature (ΔT) of the body and the surroundings.

Mathematical Form:

$$-dQ/dt = k(T_2 - T_1)$$

Where:

- dQ/dt = Rate of heat loss
- T_2 = Temperature of body
- T_1 = Temperature of surroundings
- k = Positive constant depending on area and nature of surface

Note: Law holds good only for small temperature differences

10.1 Derivation of Cooling Formula

Derivation:

Consider body of mass m, specific heat capacity s, at temperature T_2

Surroundings at temperature T_1

If temperature falls by small amount dT_2 in time dt:

Heat lost: $dQ = ms\, dT_2$

Rate of heat loss: $dQ/dt = ms(dT_2/dt)$

From Newton's law:

$-ms(dT_2/dt) = k(T_2 - T_1)$

$dT_2/(T_2 - T_1) = -(k/ms)dt = -K\, dt$

where $K = k/(ms)$

Integrating:

$\ln(T_2 - T_1) = -Kt + c$

Solution:

$T_2 = T_1 + C'e^{-Kt}$

where $C' = e^c$ (integration constant)

📝 Validity:

For small temperature differences, rate of cooling due to conduction, convection, and radiation combined is proportional to temperature difference. Valid approximation for:

- Heat transfer from radiator to room
- Loss of heat through wall of room
- Cooling of cup of tea on table

📘 Example 8 - Cooling Time:

Problem: A pan filled with hot food cools from 94 °C to 86 °C in 2 minutes when room temperature is 20 °C. How long will it take to cool from 71 °C to 69 °C?

Solution:

First Case:

Average temperature = (94 + 86)/2 = 90 °C

Excess over room = 90 - 20 = 70 °C

Cooling = 8 °C in 2 minutes

From Newton's law: (Change in temperature)/Time = K × ΔT

8 °C / 2 min = K × 70 °C

Second Case:

Average temperature = (71 + 69)/2 = 70 °C

Excess over room = 70 - 20 = 50 °C

Cooling = 2 °C in time t

2 °C / t = K × 50 °C

Dividing the two equations:

(8/2) / (2/t) = 70/50

4t/2 = 7/5

2t = 7/5

t = 0.7 min = 42 seconds

11. IMPORTANT FORMULAS - QUICK REFERENCE

Temperature Conversion:

- $t_F = (9/5)t_C + 32$
- $T = t_C + 273.15$

Ideal Gas:

- $PV = \mu RT$
- $R = 8.31\ J\ mol^{-1}\ K^{-1}$

Thermal Expansion:

- $\Delta l/l = \alpha_1 \Delta T$
- $\Delta A/A = 2\alpha_1 \Delta T$
- $\Delta V/V = \alpha_V \Delta T = 3\alpha_1 \Delta T$
- For ideal gas: $\alpha_V = 1/T$

Heat Transfer:

- $\Delta Q = ms\Delta T$ (no phase change)
- $Q = mL$ (during phase change)
- s = specific heat capacity
- C = molar specific heat capacity

Latent Heat:

- L_f (water) $= 3.33 \times 10^5\ J\ kg^{-1}$
- L_v (water) $= 22.6 \times 10^5\ J\ kg^{-1}$

Conduction:

- $H = KA(T_C - T_D)/L$

Radiation:

- $H = Ae\sigma T^4$ (Stefan-Boltzmann)
- $H = e\sigma A(T^4 - T_s^4)$ (net radiation)
- $\lambda_m T$ = constant (Wien's law)
- $\sigma = 5.67 \times 10^{-8}$ W m^{-2} K^{-4}

Newton's Law of Cooling:

- $-dQ/dt = k(T_2 - T_1)$
- $T_2 = T_1 + C'e^{-Kt}$

12. MEMORY TIPS & TRICKS

🧠 Easy Ways to Remember:

1. Temperature Conversions:

"Fahrenheit Friend" = F = (9/5)C + 32

"Kelvin Climbs" = K = C + 273

2. Expansion Relation:

"Volume = 3 × Linear" → $\alpha_V = 3\alpha_l$

"Area = 2 × Linear" → $\alpha_A = 2\alpha_l$

3. Water's Special Behavior:

"Four is More" → Water has maximum density at 4 °C

4. Specific Heat Capacity:

"Water Wins" → Water has highest specific heat capacity

Q = "Most" → Q = $ms\Delta T$

5. Latent Heat:

"Fusion is 3, Vaporization is 20+" for water

$L_f = 3.33 \times 10^5$, $L_v = 22.6 \times 10^5$

6. Heat Transfer Modes:

"Can Communicate via Radio" → **C**onduction, **C**onvection, **R**adiation

7. Thermal Conductivity:

"Metals > Non-metals > Gases" for thermal conductivity

8. Stefan-Boltzmann Law:

"Power of 4" → $H \propto T^4$

9. Newton's Cooling:

"Rate ∝ Difference" → Rate of cooling ∝ Temperature difference

13. COMMON MISTAKES TO AVOID

⚠️ Watch Out For These Errors:

1. ✗ Confusing Celsius and Kelvin in formulas requiring absolute temperature
2. ✗ Using wrong expansion coefficient (linear vs volume)
3. ✗ Forgetting that $\alpha_V = 3\alpha_l$ only for isotropic materials
4. ✗ Not considering all phase changes in heat calculations
5. ✗ Mixing up latent heat of fusion and vaporization
6. ✗ Using wrong units (converting between J and cal incorrectly)
7. ✗ Forgetting emissivity (e) in radiation formula for real bodies

8. ✗ Using T instead of T^4 in Stefan-Boltzmann law
9. ✗ Applying Newton's law of cooling for large temperature differences
10. ✗ Confusing heat capacity (S) with specific heat capacity (s)
11. ✗ Not accounting for calorimeter in calorimetry problems
12. ✗ Forgetting that gas thermometers give same reading regardless of gas

14. ADDITIONAL SOLVED PROBLEMS

📘 Problem 1 - Brass Boiler:

Problem: A brass boiler has base area 0.15 m^2, thickness 1.0 cm. It boils water at rate of 6.0 kg/min when placed on gas stove. Estimate temperature of flame in contact with boiler.

$K_{brass} = 109\ J\ s^{-1}\ m^{-1}\ K^{-1}$, $L_v = 2256 \times 10^3\ J\ kg^{-1}$

Solution:

Rate of heat flow = Rate of vaporization × L_v

$H = (6.0\ kg/min) \times (1\ min/60\ s) \times 2256 \times 10^3\ J\ kg^{-1}$

$H = 0.1 \times 2256 \times 10^3 = 2.256 \times 10^5\ J\ s^{-1}$

From conduction formula:

$H = KA(T_{flame} - T_{water})/L$

$2.256 \times 10^5 = 109 \times 0.15 \times (T_{flame} - 100)/0.01$

$2.256 \times 10^5 = 1635(T_{flame} - 100)$

$T_{flame} - 100 = 138$

T_{flame} **= 238 °C**

📘 Problem 2 - Copper Block on Ice:

Problem: A copper block of mass 2.5 kg is heated in furnace to 500 °C and then placed on large ice block. What is maximum amount of ice that can melt?

$s_{copper} = 0.39\ J\ g^{-1}\ K^{-1} = 390\ J\ kg^{-1}\ K^{-1}$

$L_f(ice) = 335\ J\ g^{-1} = 3.35 \times 10^5\ J\ kg^{-1}$

Solution:

Heat given by copper block = $ms_{Cu}\Delta T$

= 2.5 × 390 × (500 - 0)

= 487,500 J

Heat required to melt m_{ice}:

$Q = m_{ice} \times L_f$

By heat balance:

$487,500 = m_{ice} \times 3.35 \times 10^5$

$m_{ice} = 487,500/(3.35 \times 10^5)$

m_{ice} **= 1.46 kg**

📘 Problem 3 - Thermacole Icebox:

Problem: A cubical thermacole icebox of side 30 cm has thickness 5.0 cm. If 4.0 kg ice is put in box, estimate amount of ice remaining after 6 hours. Outside temperature = 45 °C.

$K_{thermacole} = 0.01$ J s^{-1} m^{-1} K^{-1}

$L_f = 335 \times 10^3$ J kg^{-1}

Solution:

Surface area of cube: $A = 6 \times (0.30)^2 = 0.54$ m^2

Thickness: L = 0.05 m

Temperature difference: $\Delta T = 45 - 0 = 45$ °C

Time: $t = 6 \times 3600 = 21{,}600$ s

Rate of heat flow:

$H = KA(\Delta T)/L$

$= 0.01 \times 0.54 \times 45/0.05$

= 4.86 W

Total heat entering in 6 hours:

$Q = H \times t = 4.86 \times 21{,}600 = 104{,}976$ J

Mass of ice melted:

$m_{melted} = Q/L_f = 104{,}976/(335 \times 10^3)$

= 0.313 kg

Ice remaining:

= 4.0 - 0.313 = 3.69 kg

📚 Study Material

Comprehensive CBSE Class 11 Physics Study Material

Chapter 10: Thermal Properties of Matter
Complete with Concepts, Derivations, Formulas & Examples

For educational purposes only. Study responsibly and practice regularly.