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Chapter 10: Thermal Physics

Class 10 Physics · New Syllabus 2026-27 · Punjab Boards

Notes by 10th.pk
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WHAT YOU LEARN

- The difference between **heat** and **temperature**, and what internal energy means.
- How a liquid-in-glass thermometer works and what the two fixed points are.
- How to convert between the Celsius, Kelvin and Fahrenheit scales.
- How to use $Q = mc\Delta T$ and $Q = mL$ in numericals.
- Why temperature stays constant during melting and boiling, and why evaporation causes cooling.

10.1 Heat and temperature

DEFINITION

Heat is the energy that flows from a hotter body to a colder body because of the difference in their temperatures. Its SI unit is the joule (J).

DEFINITION

Temperature is the degree of hotness or coldness of a body. It decides the direction in which heat flows. Its SI unit is the kelvin (K); the degree Celsius (°C) is also used.

All matter is made of particles that are always moving. The **internal energy** of a body is the total kinetic and potential energy of all its particles. Temperature depends on the *average* kinetic energy of the particles, so heating a body usually makes its particles move faster and its temperature rise.

Heat always flows from the body at higher temperature to the body at lower temperature. The flow stops when both reach the same temperature. This state is called **thermal equilibrium**.

HEAT	TEMPERATURE
Energy in transfer between two bodies	Degree of hotness of one body
SI unit: joule (J)	SI unit: kelvin (K)
Flows because of a temperature difference	Decides the direction of heat flow
Depends on mass, material and temperature change	Depends on the average kinetic energy of the particles
Calculated, for example with $Q = mc\Delta T$	Measured with a thermometer

EVERYDAY EXAMPLE

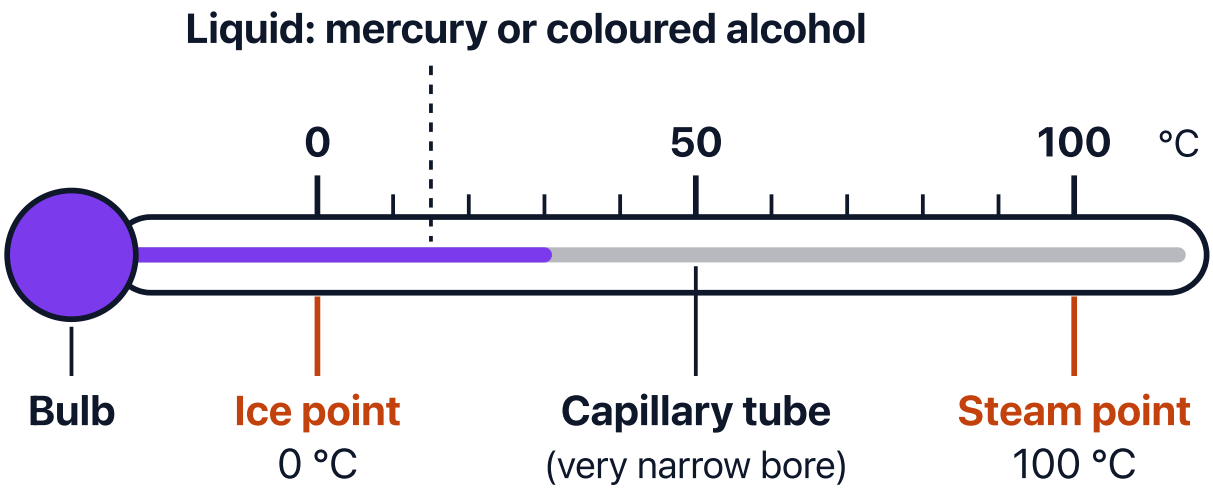
A spark flying out of a tandoor is at a very high temperature, yet it hardly hurts because it carries very little energy. A bucket of warm bath water is at a much lower temperature but holds far more internal energy, because it has so many more particles.

10.2 Thermometers

Our sense of touch is not reliable for judging temperature, so we use a thermometer.

DEFINITION

A **thermometric property** is a physical property of a substance that changes evenly with temperature, such as the volume of a liquid, the electrical resistance of a wire or the voltage of a thermocouple.



Liquid-in-glass thermometer. When the bulb is warmed, the liquid expands and moves along the narrow capillary tube; the scale is marked between the two fixed points.

How a liquid-in-glass thermometer works

- The **bulb** holds the liquid. Its thin glass wall lets heat pass in quickly.
- Liquids expand on heating. Because the **capillary tube** is very narrow, a small expansion moves the liquid thread a long way, which makes the thermometer sensitive.
- The position of the liquid thread is read on the **scale**.

Fixed points

Lower fixed point (ice point) `0 °C`

The temperature of pure melting ice at standard atmospheric pressure.

Upper fixed point (steam point) `100 °C`

The temperature of steam above pure boiling water at standard atmospheric pressure.

On the Celsius scale the distance between the two fixed points is divided into 100 equal parts. Each part is one degree Celsius.

Mercury or alcohol?

MERCURY	ALCOHOL (COLOURED)
Shiny, so it is easy to see without dye	Colourless, so a dye is added
Freezes at about −39 °C, boils at about 357 °C	Freezes at about −114 °C, boils at about 78 °C
Suitable for high temperatures	Suitable for very cold places; cannot measure boiling water
Poisonous if the glass breaks	Safer and cheaper

Other thermometers

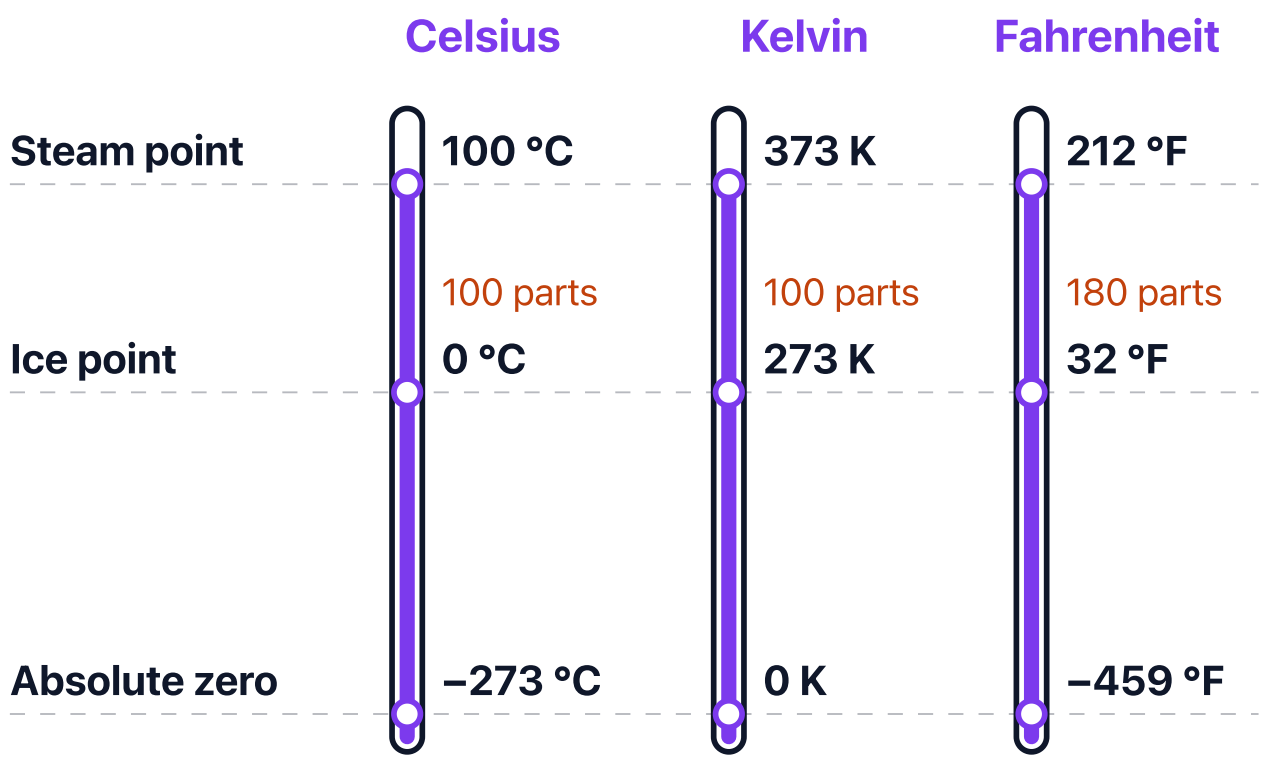
- **Clinical thermometer:** measures body temperature over a short range, about 35 °C to 42 °C. A mercury clinical thermometer has a narrow constriction above the bulb that stops the thread falling back before it is read.
- **Digital thermometer:** uses an electronic sensor whose resistance changes with temperature and shows the reading on a display.
- **Thermocouple:** two different metal wires joined at their ends. A temperature difference between the junctions produces a small voltage. It suits very high or rapidly changing temperatures, such as in a furnace.

EVERYDAY EXAMPLE

When a child in Lahore has a fever, the clinical thermometer may read 39 °C instead of the normal 37 °C (98.6 °F). In kelvin, normal body temperature is 37 + 273 = 310 K.

10.3 Temperature scales

Three scales are in common use. They agree on what is hot and what is cold, but they give different numbers to the same temperature.



The same three temperatures on the Celsius, Kelvin and Fahrenheit scales (not drawn to scale). Between the ice point and the steam point there are 100 divisions on the Celsius and Kelvin scales but 180 on the Fahrenheit scale.

Celsius to Kelvin

$$T = \theta + 273$$

T in K · θ in °C

Celsius to Fahrenheit

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

F in °F · C in °C

Fahrenheit to Celsius

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

Subtract 32 first, then multiply

DEFINITION

Absolute zero is the lowest possible temperature, 0 K (about −273 °C). At this temperature the particles of a substance have their minimum possible energy. The Kelvin scale starts from absolute zero, so it has no negative values.

One division on the Kelvin scale is the same size as one division on the Celsius scale, so a **change** of 1 K equals a change of 1 °C. Kelvin temperatures are written without a degree sign: 300 K, not 300 °K.

EVERYDAY EXAMPLE

On a June afternoon Jacobabad can reach 50 °C. In kelvin this is 50 + 273 = 323 K. In Fahrenheit it is (9 ÷ 5) × 50 + 32 = 90 + 32 = 122 °F.

10.4 Specific heat capacity

The same amount of heat does not warm every substance equally. In summer the sand on a beach gets too hot to walk on while the sea stays cool, although both receive the same sunshine.

DEFINITION

The **specific heat capacity** of a substance is the amount of heat needed to raise the temperature of 1 kg of it by 1 K (or 1 °C). Its SI unit is J kg^{-1} K^{-1}.

Heat gained or lost

$$Q = mc\Delta T$$

Q in J · m in kg · c in J kg^{-1} K^{-1} · ΔT in K or °C

Rearranged

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

ΔT = final temperature – initial temperature

Specific heat capacities used on this site

SUBSTANCE	C (J KG^{-1} K^{-1})
Water	4200
Ice	2100
Aluminium	900
Copper	390

Water has a very high specific heat capacity: 1 kg of water needs 4200 J to warm by 1 °C, about 11 times as much as 1 kg of copper. So water heats up slowly, cools down slowly and can store a lot of energy.

Heat capacity

DEFINITION

The **heat capacity** of a body is the heat needed to raise the temperature of the *whole body* by 1 K. Its SI unit is J K^{-1}.

Heat capacity

$$C = mc$$

C in J K^{-1} · depends on the mass of the body

Example

$$C = 2 \times 4200 = 8400 \text{ J K}^{-1}$$

for 2 kg of water

Specific heat capacity belongs to the *material*; heat capacity belongs to a particular *object*. A full bucket of water has a larger heat capacity than a cup of water, but both have the same specific heat capacity.

Uses of the high specific heat capacity of water

- **Car radiators:** water carries away a lot of heat from the engine with only a small rise in its own temperature.
- **Mild coastal climate:** the sea warms and cools slowly, so coastal cities have smaller temperature swings than inland cities.
- **Hot-water bottles:** water stays warm for a long time because it must lose a lot of heat to cool down.

EVERYDAY EXAMPLE

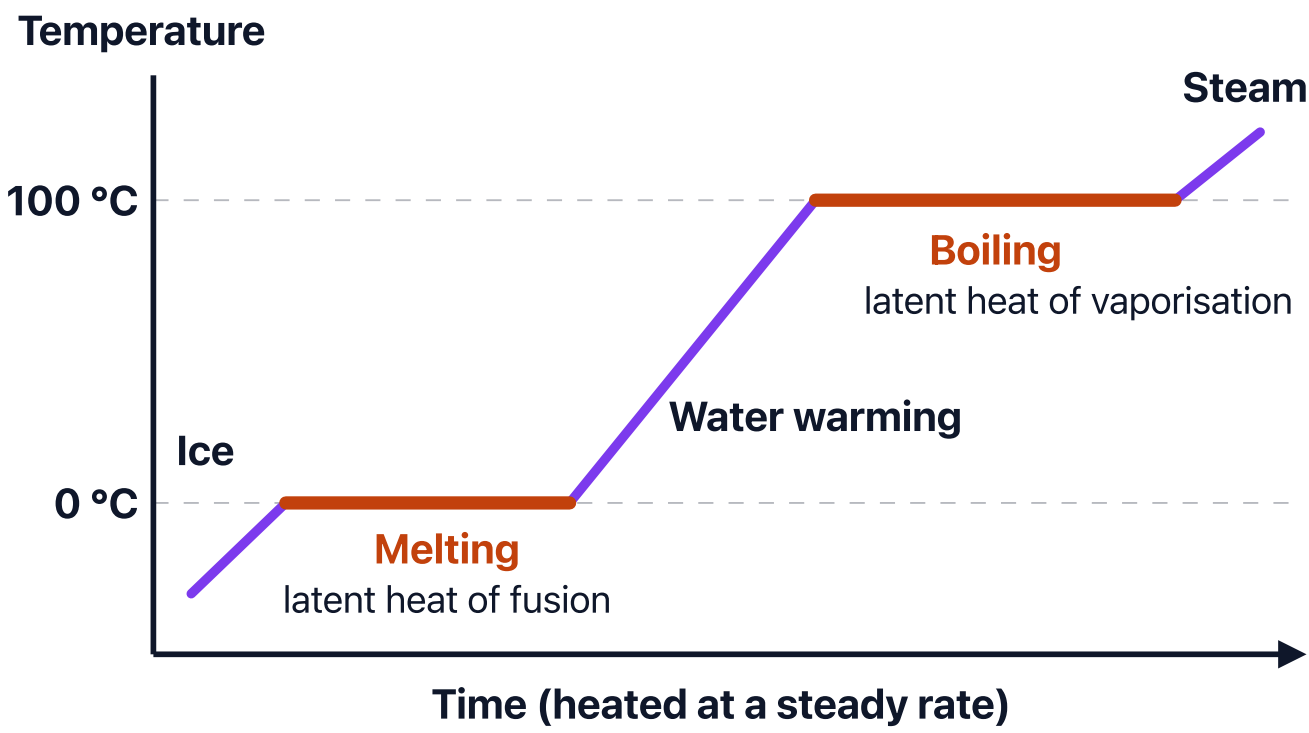
In Karachi, the land heats up faster than the sea during the day because land has a lower specific heat capacity. Warm air over the land rises and cooler air from the sea moves in to replace it. This is the sea breeze felt at Clifton in the afternoon. At night the land cools faster and the breeze reverses (land breeze).

10.5 Latent heat

When a solid melts or a liquid boils, it keeps absorbing heat but its temperature does not rise. The energy is used to pull the particles apart (it increases their potential energy), not to make them move faster.

DEFINITION

Latent heat is the heat absorbed or given out by a substance when it changes its state at constant temperature. "Latent" means hidden: a thermometer does not show it.



Heating curve for ice heated at a steady rate (time axis not to scale). The two flat parts are changes of state: the temperature stays at 0 °C until all the ice has melted, and at 100 °C until all the water has boiled.

Specific latent heat of fusion

L_f · J kg⁻¹

The heat needed to change 1 kg of a solid into liquid at its melting point, without any change in temperature. For ice: 3.36×10^5 J kg⁻¹.

Specific latent heat of vaporisation

L_v · J kg⁻¹

The heat needed to change 1 kg of a liquid into gas at its boiling point, without any change in temperature. For water: 2.26×10^6 J kg⁻¹.

Heat for a change of state

$$Q = mL$$

Q in J · m in kg · L in J kg⁻¹ · no ΔT in this formula

Melting / boiling

$$Q = mL_f, \quad Q = mL_v$$

Use L_f for melting or freezing, L_v for boiling or condensing

The same amount of heat is **given out** when the change goes the other way: steam releases latent heat when it condenses, and water releases latent heat when it freezes. Boiling 1 kg of water needs about 6.7 times as much heat as melting 1 kg of ice, which is why a burn from steam is worse than one from boiling water at the same 100 °C.

EVERYDAY EXAMPLE

Ice packed around bottles in a cooler keeps them cold for hours. To melt just 0.5 kg of ice at 0 °C, the ice must absorb $Q = mL_f = 0.5 \times 3.36 \times 10^5 = 1.68 \times 10^5$ J from the drinks, and all that time it stays at 0 °C.

10.6 Evaporation and boiling

DEFINITION

Evaporation is the change of a liquid into vapour from its *surface*, at any temperature below its boiling point.

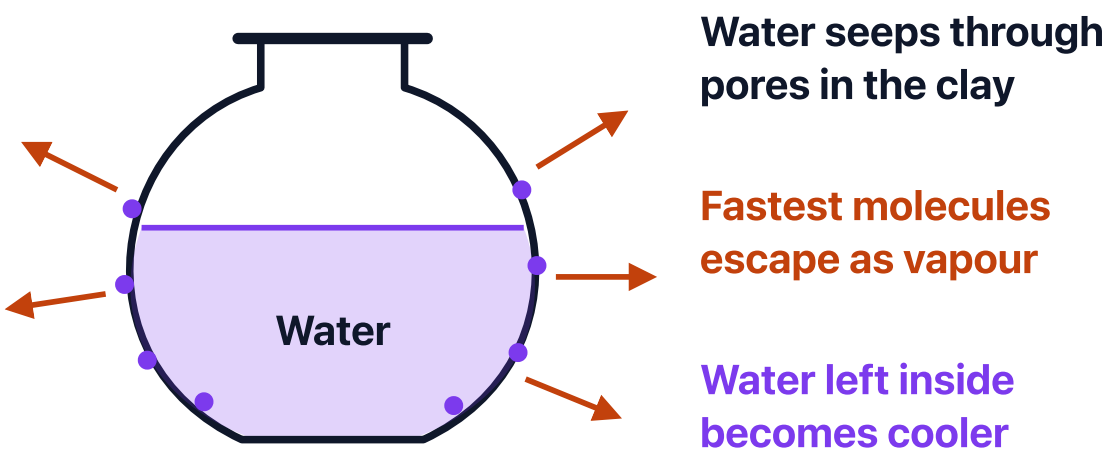
DEFINITION

Boiling is the rapid change of a liquid into vapour throughout the liquid at one fixed temperature, called its boiling point. Bubbles of vapour form inside the liquid.

EVAPORATION	BOILING
Happens at **any** temperature	Happens only at the **boiling point**
Takes place only at the surface	Takes place throughout the liquid
Slow and quiet; no bubbles	Fast; bubbles form
Takes heat from the liquid and its surroundings, so it causes cooling	Needs a source of heat; temperature stays constant

Why evaporation causes cooling

The particles of a liquid move at different speeds. Only the fastest particles near the surface have enough energy to escape. When they leave, the average kinetic energy of the particles left behind falls, so the temperature of the liquid drops.



Cooling by evaporation in a clay matka: water that seeps through the pores evaporates from the outer surface and takes heat from the water inside.

Factors that increase the rate of evaporation

Higher temperature
More particles have enough energy to escape. Wet clothes dry faster on a hot day.

Larger surface area
More particles are at the surface. Chai cools faster in a saucer than in a cup.

Wind
Moving air carries the vapour away. Clothes dry faster on a windy roof.

Lower humidity
Dry air can take up more vapour. Sweat dries slowly in the humid monsoon.

The kind of liquid also matters: spirit or perfume evaporates much faster than water at the same temperature.

EVERYDAY EXAMPLE

Water kept in a matka stays pleasantly cool even in the summer heat of Multan. The same idea works when you sweat: as sweat evaporates from your skin it takes heat from your body, and a fan speeds this up.

Worked example: using $Q = mc\Delta T$

Example

How much heat is needed to warm 500 g of water for chai from 30 °C to 100 °C? (Specific heat capacity of water = 4200 J kg^{-1} K^{-1})

GIVEN m = 500 g = 0.5 kg
c = 4200 J kg^{-1} K^{-1}
ΔT = 100 − 30 = 70 °C = 70 K

TO FIND Q = ? (J)

FORMULA $Q = mc\Delta T$

SOLUTION

$$Q = mc\Delta T$$
$$Q = 0.5 \times 4200 \times 70$$
$$Q = 147\,000 \text{ J}$$

Q = 147 000 J = 1.47 × 10^5 J (147 kJ)
Mass converted to kg first; answer written with its SI unit.

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Common mistakes (teacher's notes)

Watch out: most marks in this chapter are lost on units, not on physics.

- **Grams not changed to kilograms.** Q = mcΔT and Q = mL need mass in kg: 250 g = 0.25 kg.
- **Adding 273 to a temperature change.** A rise from 20 °C to 50 °C is ΔT = 30 °C = 30 K. Add 273 only when converting a temperature, never a temperature *difference*.
- **Using Q = mcΔT during melting or boiling.** The temperature does not change during a change of state, so ΔT = 0. Use Q = mL for that part.
- **Wrong order in °F conversions.** °C to °F: multiply by 9/5 first, then add 32. °F to °C: subtract 32 first, then multiply by 5/9.
- **Writing "°K".** The unit is K with no degree sign, and there are no negative kelvin temperatures.
- **Mixing up heat and temperature.** Heat is energy in joules; temperature is hotness in K or °C.
- **Mixing up c and C.** Specific heat capacity c is in J kg^{-1} K^{-1}; heat capacity C = mc is in J K^{-1}.

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