14 THERMAL PHYSICS

14.1 Temperature

Specification reference: 5.1.1

Learning outcomes

Demonstrate knowledge, understanding, and application of:

→ thermal equilibrium

→ the absolute scale of temperature

→ temperature measurements in degrees Celsius and kelvin

→ $T(\text{K}) \approx \theta(^\circ\text{C}) + 273$.

▲ **Figure 1** *The three phases of water – solid, liquid, and gas – are in thermal equilibrium inside this triple-point cell: the ice does not melt, the water vapour does not condense*

The triple point

The **triple point** of a substance is one specific temperature and pressure where a strange thing happens. There, and nowhere else, the three **phases** of matter (solid, liquid, and gas) of that substance can exist in **thermal equilibrium**, that is, there is no net transfer of thermal energy between the phases.

For water, the triple point is at 0.01°C and 0.61 kPa, less than 1% of normal atmospheric pressure.

Temperature and thermal equilibrium

A simple way to think about temperature is as a measure of the hotness of an object on a chosen scale. The hotter an object is, the higher its temperature.

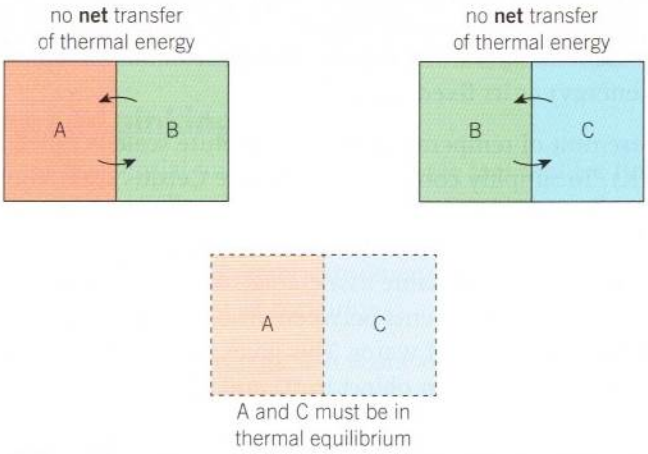
If one object is hotter than another there is a net flow of thermal energy from the hotter object into the colder one. This increases the temperature of the colder object and lowers the temperature of the hotter one. For example, when the outside air temperature is lower than your body temperature there is a net flow of energy from you to your surroundings.

When two objects are in thermal equilibrium there is no net flow of thermal energy between them. This means any objects in thermal equilibrium must be at the same temperature.

The zeroth law of thermodynamics

The zeroth (0^{th}) law of thermodynamics was proposed after three laws were already recognised. (You do not need to know the other laws for this course, although you are already familiar with conservation of energy, one aspect of the first law.) It was deemed so fundamental to the study of thermal physics that it was named the zeroth law, coming before the others.

The zeroth law states that if two objects are each in thermal equilibrium with a third, then all three are in thermal equilibrium with each other. In other words, if both A and C are in thermal equilibrium with B, then A is in thermal equilibrium with C. This means that all three objects are at the same temperature.



▲ **Figure 2** *The zeroth law introduces temperature as a physical property*

It may seem obvious, but the zeroth law means that objects have a measurable physical property that determines the direction of any transfer of thermal energy. This property is temperature.

The zeroth law forms the basis for a definition of temperature and thus for comparing temperatures, describes how thermometers work, and is important for the mathematical formulation of laws about the effect of changing temperature.

1. Describe the transfer of thermal energy from object A to object D if D is at a lower temperature than A.
2. In order to measure the temperature of an object accurately, simple liquid-in-glass thermometers must be at the same temperature as the object. Outline the reason.

Measuring temperature

In order to measure temperature a scale is needed that includes two fixed points at defined temperatures. The temperature of other objects can then be defined as a position on this scale.

Most of the world uses the **Celsius scale** proposed by the Swedish astronomer Anders Celsius in 1742. He suggested the freezing point and the boiling point of pure water (when the atmospheric pressure is 1.01×10^5 Pa) as the two fixed points, with 100 increments (or degrees) between 0°C and 100°C. By definition, any object at 100°C must be in thermal equilibrium with boiling water.

However, the Celsius scale is not perfect. Although its two fixed points seem simple to obtain, they vary significantly depending on the surrounding atmospheric pressure. For example, on top of a high mountain water boils at a lower temperature (as low as 70°C).

The **absolute temperature** scale (or **thermodynamic temperature scale**) uses the triple point of pure water and **absolute zero** (the lowest possible temperature, explored in more detail in Topic 14.3, Internal energy) as its fixed points.

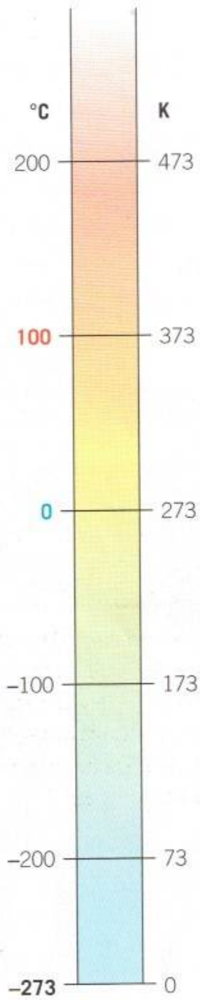
The SI base unit of temperature on the absolute scale is called the **kelvin** (K). To simplify comparison with the Celsius scale, the scientific community agreed that the increments on the absolute scale would be the same size as those on the Celsius scale, so a change in temperature of 1 K is the same as a change of 1°C. As a result, there are exactly 273.16 increments between absolute zero (now defined as 0 K) and the triple point of water. This gives the following relationship between temperature of an object in °C and in K:

$$T(\mathrm{K}) \approx \theta(^\circ\mathrm{C}) + 273$$

Temperatures in K are always positive, and the lowest temperature on the absolute scale is 0 K (see Figure 3).

Synoptic link

The kelvin is one of the seven base units studied in Topic 2.1, Quantities and units.



▲ **Figure 3** *To convert from temperature in °C to temperature in K, add 273*

Study tip

0 °C is equivalent to 273.15 K or approximately 273 K.

Summary questions

1 Describe the net transfer of thermal energy between two objects A and B at the temperatures given in Table 1. *(3 marks)*

▼ **Table 1**

	Temperature of object A / °C	Temperature of object B / °C
a	100	0
b	50	50
c	−90	−40

2 Convert the following temperatures from °C into K:
a 0°C **b** 37.0°C **c** −120.5°C. *(3 marks)*

3 Convert the following temperatures from K into °C:
a 0 K **b** 200 K **c** 350 K. *(3 marks)*

4 Explain why is it not possible for an object to have a temperature of −50 K. *(1 mark)*

5 Describe the net transfer of thermal energy and any changes in temperature when a metal block at 300 K is placed in water at 15°C. *(3 marks)*

6 The typical core temperature of a star is about 10^7 K. When astronomers discuss the core temperatures of stars they often omit the unit °C or K. Suggest whether this is sensible. *(2 marks)*

7 Suggest a reason, in terms of thermal energy transfer, why a typical liquid-in-glass thermometer at room temperature placed into a cup of hot water does not give a truly accurate reading of the initial temperature of the water even when they reach thermal equilibrium. *(3 marks)*

14.2 Solids, liquids, and gases

Specification reference: 5.1.2

Floating and sinking

Liquid water is essential for life, and the simple fact that ice floats on water (Figure 1) has allowed complex life on Earth to survive the most extreme of ice ages. It means water freezes from the top downwards, so a small amount of liquid water can remain insulated underneath the ice except in the coldest of conditions.

Water is very unusual in this regard. It is one of only a few substances that is less dense in its solid phase than its liquid phase. To understand the reasons why, we need to look carefully at the nature of water molecules.

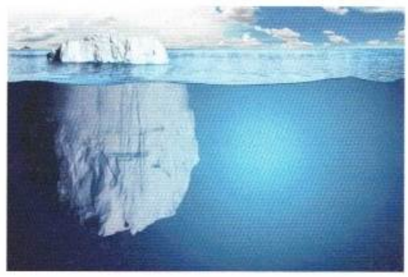
Learning outcomes

Demonstrate knowledge, understanding, and application of:

→ solids, liquids, and gases in terms of spacing, ordering, and motion of atoms or molecules

→ the simple kinetic model

→ Brownian motion.



▲ **Figure 1** *An iceberg illustrates clearly how solid water is less dense than liquid water – before modern ship radar, icebergs were a significant hazard to shipping, perhaps most famously sinking the Titanic in 1912*

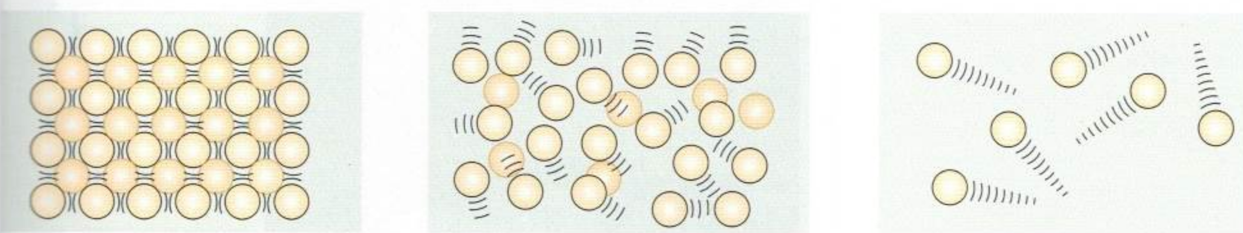
The kinetic model

The **kinetic model** describes how all substances are made up of atoms or molecules, which are arranged differently depending on the phase of the substance.

In solids the atoms or molecules are regularly arranged and packed closely together, with strong electrostatic forces of attraction between them holding them in fixed positions, but they can vibrate and so have kinetic energy (Figure 2).

In liquids the atoms or molecules are still very close together, but they have more kinetic energy than in solids, and – unlike in solids – they can change position and flow past each other.

In gases, the atoms or molecules have more kinetic energy again than those in liquids, and they are much further apart. They are free to move past each other as there are negligible electrostatic forces between them, unless they collide with each other or the container walls. They move randomly with different speeds in different directions.



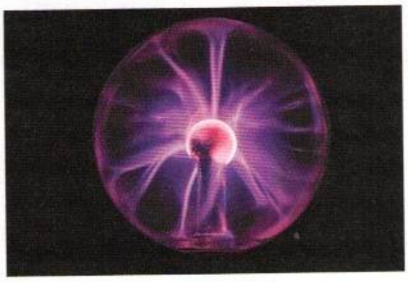
A solid is made up of particles (atoms or molecules) arranged in a regular 3-dimensional structure. There are strong forces of attraction between the particles. Although the particles can vibrate, they cannot move out of their positions in the structure.
When a solid is heated, the particles gain energy and vibrate more and more vigorously. Eventually they may break away from the solid structure and become free to move around. When this happens, the solid has turned into liquid: it has melted.

In a liquid the particles are free to move around. A liquid therefore flows easily and has no fixed shape. There are still forces of attraction between the particles.
When a liquid is heated, some of the particles gain enough energy to break away from the other particles. The particles which escape from the body of the liquid become a gas.

In a gas, the particles are far apart. There are almost no forces of attraction between them. The particles move about at high speed. Because the particles are so far apart, a gas occupies a very much larger volume than the same mass of liquid.

The molecules collide with the container. These collisions are responsible for the pressure which a gas exerts on its container.

▲ **Figure 2** *The kinetic model of three phases of matter and their differing energies*



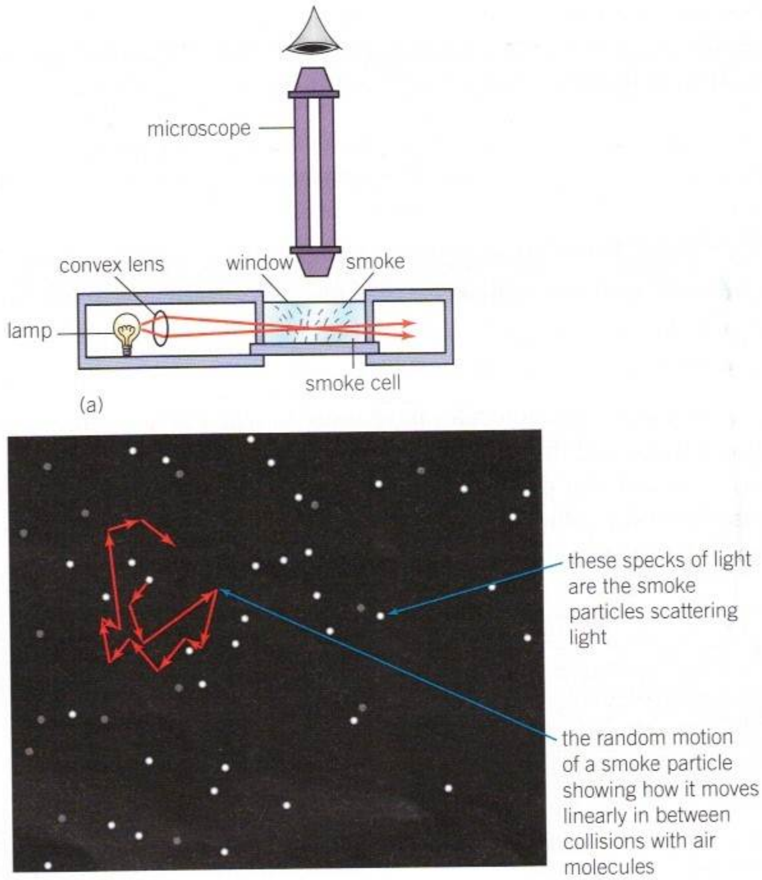
▲ Figure 3 *Matter can exist in phases other than solid, liquid, and gas – in fact plasma, formed from gas so hot that its atoms are ionised, is the most common phase of matter in the universe, and can be made on Earth by applying a high potential difference across a gas at low pressure as in this plasma ball*

Observing Brownian motion

The idea that substances were made of particles (atoms or molecules) was discussed for centuries, but not confirmed until 1827 when Robert Brown looked through a microscope and recorded his observations of the random movements of fine pollen grains floating on water.

It was not until 1905 that Albert Einstein fully explained this **Brownian motion** in terms of collisions between the pollen grains and millions of tiny water molecules. He explained that these collisions were elastic and resulted in a transfer of momentum from the water molecules to the pollen grains, causing the grains to move in haphazard ways. This provided the first significant proof of the kinetic model – the idea that matter is made up of atoms and molecules and they have kinetic energy.

It is possible to observe Brownian motion in the laboratory using a smoke cell (Figure 4).



▲ Figure 4 *Observing the random paths of smoke particles using a smoke cell*

Particles of smoke are large enough to be visible under a microscope. These particles move around in a random way. The random motion is caused by air molecules constantly striking the smoke particles. The air molecules themselves are in random motion. The mean kinetic energy of the smoke particles is the same as the mean kinetic energy of the air molecules. However, while the air molecules typically move with a speed around 500 m s^{-1}, the more massive smoke particles move much more slowly.

1 Sketch the path of a pollen grain being bombarded by water molecules.
2 Explain what happens to the motion of a smoke particle in air if the air temperature decreases.

Density

The spacing between the particles (atoms or molecules) in a substance in different phases affects the density of the substance. In general a substance is most dense in its solid phase and least dense in its gaseous phase. Unusually, solid water is less dense than liquid water. Water freezes into a regular crystalline pattern held together by strong electrostatic forces between the molecules. In this structure the molecules are held slightly further apart than in their random arrangement in liquid water, so ice is slightly less dense.

Synoptic link

Density was introduced in Topic 4.8, Density and pressure.

Summary questions

1 List the three main phases of a substance in order of the energy of the particles (atoms or molecules) in that substance. *(1 mark)*

2 Use diagrams of how atoms or molecules are arranged in solids, liquids, and gases to explain why gases have a much lower density than solids. *(1 mark)*

3 Water of mass 2.0 kg is gradually heated. Its volume is measured at each of the temperatures given in Table 1.

a Use the data in Table 1 to determine the density of water at the of the temperatures shown. *(3 marks)*

Table 1 *Volume of 2.0 kg water at various temperatures*

Temperature / °C	5.0	20.0	40.0	60.0	90.0
Volume / 10^{-3} m^3	2.000	2.004	2.016	2.034	2.075

b Explain why the volume of water increases as its temperature increases. *(2 marks)*

c Suggest why, along with the melting of land ice, an increase in global temperature results in a rise in sea levels. *(1 mark)*

4 The mass of one water molecule is 3.0×10^{-26} kg. The density of ice is 920 kg m^{-3} and of water vapour (at boiling point) is 0.590 kg m^{-3}. Calculate the number of water molecules in:

a 1.0 m^3 of ice   b 1.0 m^3 of water vapour. *(5 marks)*

5 Use your values above to estimate the spacing between water molecules in ice and water vapour. *(5 marks)*

14.3 Internal energy

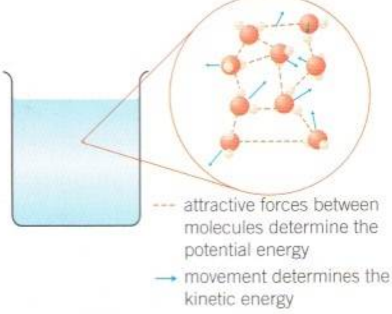
Specification reference: 5.1.2

Learning outcomes

Demonstrate knowledge, understanding, and application of:

- → internal energy as the sum of kinetic and potential energies in a system
- → absolute zero (0 K)
- → increase in internal energy with temperature
- → changes in internal energy during changes of phase
- → constancy of temperature during changes of phase.

▲ **Figure 1** *The Vostok station still holds the official record for the coldest place on Earth – although satellite data in 2010 indicated a new low of −93.2°C (also in Antarctica). However, this was not confirmed by measurements on the ground*



▲ **Figure 2** *A beaker of water has an internal energy due to the kinetic and potential energies of the water molecules*

The coldest place on Earth

The lowest natural temperature ever measured on Earth is −89.2°C (184 K), recorded in 1983 at the Russian Vostok research station in Antarctica (Figure 1). This is cold enough for carbon dioxide to solidify.

Lower temperatures have been achieved artificially in laboratories. The current record, set in 1999, is 100 pK, or 1.0×10^{-10} K (much colder than the deep space between galaxies, at 2.7 K). But it will never be possible to reach 0 K, and to understand why we need to understand what happens inside a substance as it changes temperature.

Internal energy and absolute zero

The **internal energy** of a substance is defined as:

The sum of the randomly distributed kinetic and potential energies of atoms or molecules within the substance.

Consider a beaker of water at room temperature (Figure 2). The water contains a huge number of water molecules travelling at hundreds of meters per second. The internal energy of the water is the sum of all the individual kinetic energies of the water molecules in the glass and the sum of all the potential energies due to the electrostatic intermolecular forces between the molecules.

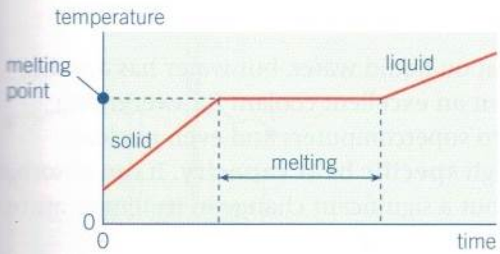
Now imagine cooling the beaker. The water will freeze and the water molecules move more slowly as the ice gets colder. Absolute zero is the lowest temperature possible. At this temperature the internal energy of a substance is a minimum. The kinetic energy of all the atoms or molecules is zero – they have stopped moving. However, the internal energy is not zero because the substance still has electrostatic potential energy stored between the particles. Even at 0 K, you cannot reduce the potential energy of the substance to zero.

Increasing the internal energy of a body

Increasing the temperature of a body will increase its internal energy. As the temperature increases, the average kinetic energy of the atoms or molecules inside the body increases. In general, the hotter a substance, the faster the atoms or molecules that make up the substance move, and the greater the internal energy of the substance.

However, it is not only increasing the temperature of a body that increases its internal energy. When a substance changes phase, for example from solid to liquid, the temperature does not change, nor does the kinetic energy of the atoms or molecules. However, their electrostatic potential energy increases significantly.

If a solid substance is heated using a heater with a constant power output, a graph showing how the temperature increases with time can be recorded (Figure 3).



▲ **Figure 3** *The change from solid to liquid or from liquid to gas increases the internal energy of the substance, even though the temperature remains the same while the substance changes phase (the horizontal line on the graph)*

When a substance reaches its melting or boiling point, while it is changing phase the energy transferred to the substance does not increase its temperature. Instead the electrostatic potential energy of the substance increases as the electrical forces between the atoms or molecules change. Only once the phase change is complete does the kinetic energy of the atoms or molecules increase further, and so the temperature rises again.

In different phases the atoms or molecules of a substance have different electrostatic potential energies:

- Gas: The electrostatic potential energy is zero because there are negligible electrical forces between atoms or molecules.
- Liquid: The electrostatic forces between atoms or molecules give the electrostatic potential energy a negative value. The negative simply means that energy must be supplied to break the atomic or molecular bonds.
- Solid: The electrostatic forces between atoms or molecules are very large, so the electrostatic potential energy has a large negative value.

The electrostatic potential energy is lowest in solids, higher in liquids, and at its highest (0 J) in gases.

Summary questions

1 Explain why it is not possible to achieve a temperature lower than 0 K. *(1 mark)*

2 Describe what happens to the energy transferred to a substance being heated when it changes phase. *(2 marks)*

3 State two ways to increase the internal energy of a substance. *(2 marks)*

4 Explain why 1.0 kg of water at 0°C has more internal energy than 1.0 kg of ice at 0°C. *(2 marks)*

5 Explain, in terms of internal energy, why a window gets slightly warmer when water vapour condenses on its surface. *(2 marks)*

Synoptic link

You will learn more about average kinetic energy and temperature in Topic 15.4, the Boltzmann constant.

▲ **Figure 4** *The metal gallium has a melting point of 30°C, low enough to melt in a hand*