Cambridge

A2 Level

Physics

CODE: (9702)

Chapter 21

Chapter 21: Thermal physics

Changes of state

The kinetic model of matter can be used to describe the structure of solids, liquids and gases. You should recall that the kinetic model describes the behaviour of matter in terms of moving particles (atoms, molecules, etc.). Figure 21.2 should remind you of how we picture the three states of matter at the atomic scale:

■ ■ In a solid, the particles are close together, tightly bonded to their neighbours, and vibrating about fixed positions.

■ ■ In a gas, the particles have broken free from their neighbours; they are widely separated and are free to move around within their container

Energy changes

Energy is required to raise, melt, heat, and boil a solid, liquid, or gas. It's crucial to examine a single change of state on the atomic scale. A test tube containing octadecanoic acid is heated, cooled, and monitored using a thermometer or probe and datalogger.

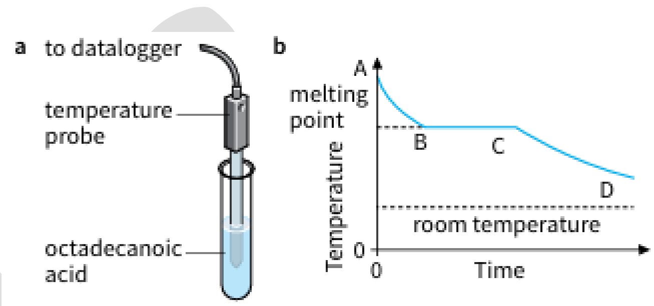

Figure 21.3 **a** Apparatus for obtaining a cooling curve, and **b** typical results.

Heating ice

The experiment can be reimagined as heating ice from deep freeze in a well-insulated container. The temperature rises, forming a container of water vapour. This is an invisible gas, unlike the steam seen in kettles. Energy is supplied at a constant rate.

We need to think about the kinetic and potential energies of the molecules. If they move around more freely and faster, their kinetic energy has increased. If they break free of their neighbours and become more disordered, their electrical potential energy has increased.

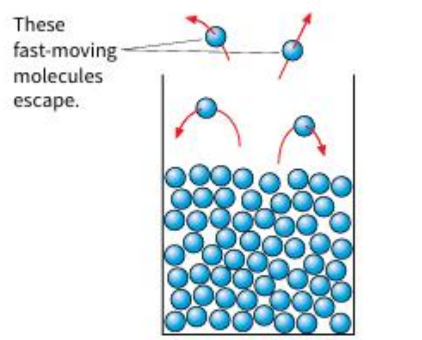

Figure 21.6 Fast-moving molecules leave the surface of a liquid – this is evaporation.

Work must be done (energy must be put in) to separate neighbouring atoms – think about the work you must do to snap a piece of plastic or to tear a sheet of paper. The graph shows that:

■ ■ the electrical potential energy of two atoms very close together is large and negative

■ ■ as the separation of the atoms increases, their potential energy also increases

■ ■ when the atoms are completely separated, their potential energy is maximum and has a value of zero.

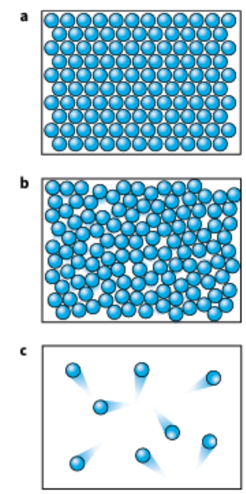

Figure 21.2 Typical arrangements of atoms in **a** a solid, **b** a liquid and **c** a gas.

Evaporation

A liquid does not have to boil to change into a gas. A puddle of rain-water dries up without having to be heated to 100 °C. When a liquid changes to a gas without boiling, we call this evaporation.

Internal energy

Matter is composed of molecules, which can have energy. For example, a stone's potential and kinetic energies are properties determined by its mass and speed. Heat can increase the energy of a stone, but it doesn't increase its gravitational potential or kinetic energies, suggesting that the energy has disappeared into the stone.

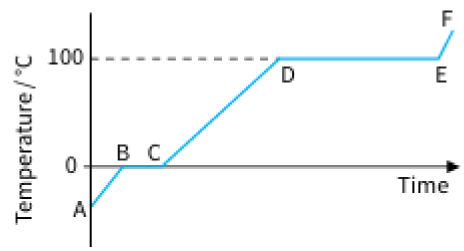

Figure 21.4 A graph of temperature against time for water, heated at a steady rate.

They vibrate more and faster, and they move a little further apart. This energy of the molecules is known as the internal energy of the stone. The internal energy of a system (e.g. the heated stone) is defined as follows:

> The internal energy of a system is the sum of the random distribution of kinetic and potential energies of its atoms or molecules.

Molecular energy

Earlier in this chapter, where we studied the phases of matter, we saw how solids, liquids and gases could be characterised by differences in the arrangement, order and motion of their molecules. We could equally have said that, in the three phases, the molecules have different amounts of kinetic and potential energy.

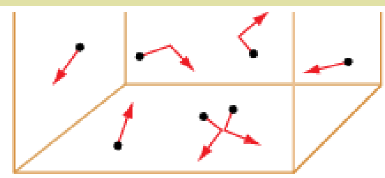

Figure 21.8 The molecules of a gas have both kinetic and potential energy.

Changing internal energy

There are two obvious ways in which we can increase the internal energy of some gas: we can heat it (Figure 21.9a), or we can do work on it by compressing it (Figure 21.9b).

Heating a gas

T h e walls of the container become hot and so its molecules vibrate more vigorously. Th e molecules of the cool gas strike the walls and bounce off faster. They have gained kinetic energy, and we say the temperature has risen.

Doing work on a gas

In this case, a wall of the container is being pushed inwards. Th e molecules of the cool gas strike a moving wall and bounce off faster. They have gained kinetic energy and again the temperature has risen. This explains why a gas gets hotter when it is compressed.

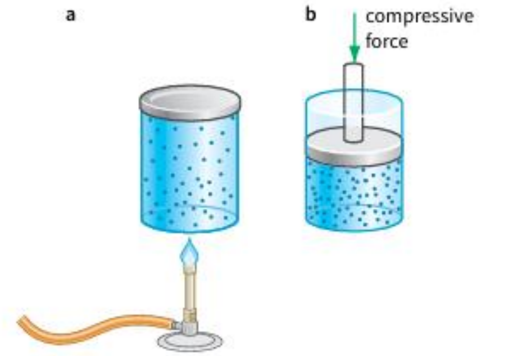

Figure 21.9 Two ways to increase the internal energy of a gas: **a** by heating it, and **b** by compressing it.

First law of thermodynamics

You will be familiar with the idea that energy is conserved; that is, energy cannot simply disappear, or appear from nowhere. Th is means that, for example, all the energy put into a gas by heating it and by doing work on it must end up in the gas; it increases the internal energy of the gas.

We can write this as an equation:

increase in internal energy = energy supplied by heating + energy supplied by doing work

In symbols:

$$\Delta U = q + w$$

This is known as the **first law of thermodynamics** and is a formal statement of the principle of conservation of energy. (It applies to all situations, not simply to a mass of gas.)

The meaning of temperature

Picture a beaker of boiling water. You want to measure its temperature, so you pick up a thermometer which is lying on the bench. Th e thermometer reads 20 °C. You place the thermometer in the water and the reading goes up … 30 °C, 40 °C, 50 °C. Th is tells you that the thermometer is getting hotter; energy is being transferred from the water to the thermometer.

This simple, everyday activity illustrates several points:

■ ■ We are used to the idea that a thermometer shows the temperature of something with which it is in contact. In fact, it tells you its own temperature. As the reading on the scale was rising, it wasn't showing the temperature of the water. It was showing that the temperature of the thermometer was rising.

■ ■ Energy is transferred from a hotter object to a cooler one. The temperature of the water was greater than the temperature of the thermometer, so energy transferred from one to the other.

■ ■ When two objects are at the same temperature, there is no transfer of energy between them. That is what happened when the thermometer reached the same temperature as the water, so it was safe to say that the reading on the thermometer was the same as the temperature of the water.

Energy flowing from a region of higher temperature to a region of lower temperature is called **thermal energy**

When two objects, in contact with each other, are at the same temperature, there will be no net transfer of thermal energy between them. We say that they are in **thermal equilibrium** with each other – see Figure 21.10.

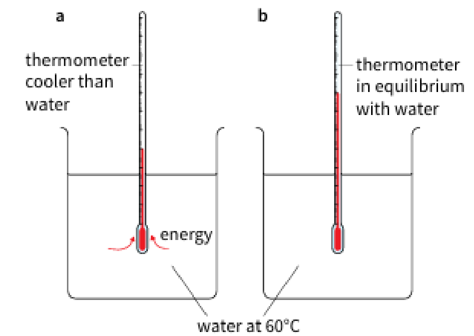

Figure 21.10 **a** Thermal energy is transferred from the hot water to the cooler thermometer because of the temperature difference between them. **b** When they are at the same temperature, there is no transfer of thermal energy and they are in thermal equilibrium.

The thermodynamic (Kelvin) scale

There is nothing special about these two fixed points. In fact, both change if the pressure changes or if the water is impure. Th e **thermodynamic scale,** also known as the Kelvin scale, is a better scale in that one of its fixed points, **absolute zero**, has a greater significance than either of the Celsius fixed points.

The thermodynamic scale has two fixed points:

■ ■ absolute zero, which is defined as 0 K

■ ■ the triple point of water, the temperature at which ice, water and water vapour can co-exist, which is defined as 273.16 K (equal to 0.01 °C).

A **change** in temperature of 1 K is thus equal to a **change** in temperature of 1 °C.

For most practical purposes, we round off the conversion factor to 273 as shown in the conversion chart (Figure 21.11).

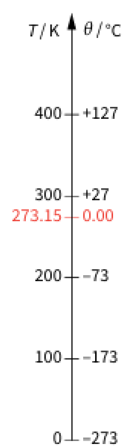

Figure 21.11 A conversion chart relating temperatures on the thermodynamic (Kelvin) and Celsius scales.

Thermometers

A thermometer is any device which can be used to measure temperature. Other properties which can be used as the basis of thermometers include:

■ ■ the resistance of an electrical resistor or thermistor

■ ■ the voltage produced by a thermocouple

■ ■ the colour of an electrically heated wire

■ ■ the volume of a fixed mass of gas at constant pressure.

A **thermocouple** is another electrical device which can be used as the sensor of a thermometer. Figure 21.12 shows the principle. Wires of two different metals, X and Y, are required. A length of metal X has a length of metal Y soldered to it at each end. This produces two **junctions**, which are the important parts of the thermocouple.

Calculating energy changes

So far, we have considered the effects of heating a substance in qualitative terms, and we have given an explanation in terms of a kinetic model of matter

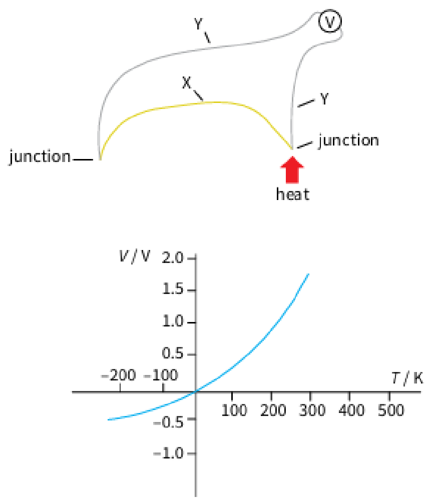

Figure 21.12 The construction of a thermocouple thermometer; the voltage produced depends on the temperature (as shown in the calibration graph) and on the metals chosen.

Feature	Resistance thermometer	Thermocouple thermometer
robustness	very robust	robust
range	thermistor: narrow range resistance wire: wide range	can be very wide
size	larger than thermocouple, has greater thermal capacity therefore slower acting	smaller than resistance thermometers, has smaller thermal capacity, so quicker acting and can measure temperature at a point
sensitivity	thermistor: high sensitivity over narrow range resistance wire: less sensitive	can be sensitive if appropriate metals chosen
linearity	thermistor: fairly linear over narrow range resistance wire: good linearity	non-linear so requires calibration
remote operation	long conducting wires allow the operator to be at a distance from the thermometer	long conducting wires allow the operator to be at a distance from the thermometer

Table 21.2 Comparing resistance and thermocouple thermometers.

Specific heat capacity

If we heat some material so that its temperature rises, the amount of energy we must supply depends on three things:

■ ■ the mass m of the material we are heating
■ ■ the temperature change Δθ we wish to achieve
■ ■ the material itself.

where c is the **specific heat** capacity of the material. Rearranging this equation gives:

$$c = \frac{E}{m\Delta\theta}$$

The specific heat capacity of a material can be defined as a word equation as follows:

$$\text{specific heat capacity} = \frac{\text{energy supplied}}{\text{mass} \times \text{temperature change}}$$

Alternatively, specific heat capacity can be defined in words as follows:

The specific heat capacity of a substance is the energy required per unit mass of the substance to raise the temperature by 1 K (or 1 °C).

Substance	c / J kg^{-1} K^{-1}
aluminium	880
copper	380
lead	126
glass	500–680
ice	2100
water	4180
seawater	3950
ethanol	2500
mercury	140

Table 21.3 Values of specific heat capacity.