8.2 Thermal energy transfer

Understandings, applications, and skills:

Conduction, convection, and thermal radiation

- Discussion of conduction and convection will be qualitative only

Guidance

- *Discussion of conduction is limited to intermolecular and electron collisions.*
- *Discussion of convection is limited to simple gas or liquid transfer via density difference.*

The role of the physicist is to observe our physical surroundings, take measurements and think of ways to explain what we see. Up to this point in the course we have been dealing with the motion of bodies. We can describe bodies in terms of their mass and volume, and if we know their speed and the forces that act on them, we can calculate where they will be at any given time. We even know what happens if two bodies hit each other. However, this is not enough to describe all the differences between objects. For example, by simply holding different objects, we can feel that some are hot and some are cold.

In this chapter we will develop a model to explain these differences, but first of all we need to know what is inside matter.

The particle model of matter

Ancient Greek philosophers spent a lot of time thinking about what would happen if they took a piece of cheese and kept cutting it in half.

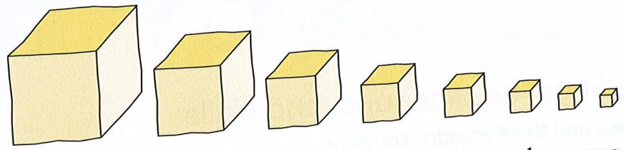

Figure 3.1 Can we keep cutting the cheese for ever?

They didn't think it was possible to keep halving it for ever, so they suggested that there must exist a smallest part – this they called the *atom*.

Atoms are too small to see (about 10^{-10} m in diameter) but we can think of them as very small perfectly elastic balls. This means that when they collide, both momentum and kinetic energy are conserved.

Elements and compounds

We might ask: 'If everything is made of atoms, why isn't everything the same?' The answer is that there are many different types of atom.

There are 117 different types of atom, and a material made of just one type of atom is called an *element*. There are, however, many more than 117 different types of material. The other types of matter are made of atoms that have joined together to form molecules. Materials made from molecules that contain more than one type of atom are called *compounds*.

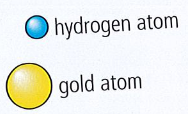

Figure 3.2 Gold is made of gold atoms and hydrogen is made of hydrogen atoms.

How do we know atoms have different masses?

The answer to that question is thanks originally to the chemists, John Dalton in particular. Chemists make compounds from elements by mixing them in very precise proportions. This is quite complicated but we can consider a simplified version as shown in Figure 3.3. An atom of A joins with an atom of B to form molecule AB.

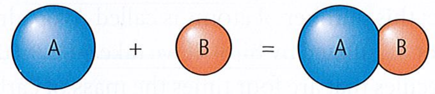

Figure 3.3 Atoms join to make a molecule.

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This is a good example of how models are used in physics. Here we are modelling something that we can't see, the atom, using a familiar object, a perfectly elastic ball.

We first try by mixing the same masses of A and B but find that when the reaction has finished there is some B left over; we must have had too many atoms of B. If we reduce the amount of B until all the A reacts with all the B to form AB we know that there must have been the same number of atoms of A as there were of B as shown in Figure 3.4. We can therefore conclude that the mass of an A atom is larger than an atom of B. In fact the ratio of

$$\frac{\text{mass of atom A}}{\text{mass of atom B}} = \frac{\text{total mass A}}{\text{total mass B}}.$$

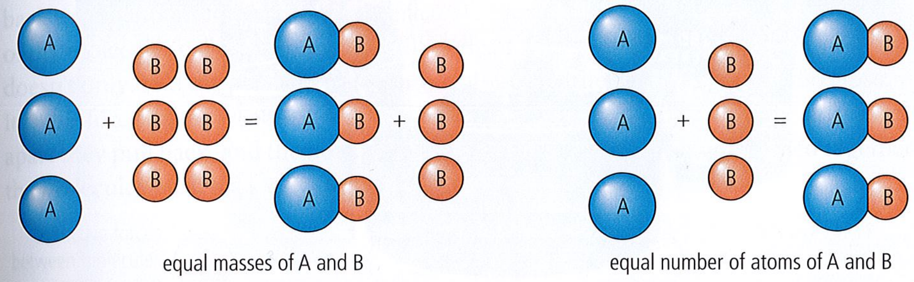

Figure 3.4 To combine completely there must be equal numbers of atoms.

By finding out the ratio of masses in many different reactions the atomic mass of the elements relative to each other was measured. Originally everything was compared to oxygen since it reacts with so many other atoms. Later, when physicists started to measure the mass of individual atoms the standard atom was changed to carbon-12. This is taken to have an atomic mass of exactly 12 unified mass units (u), the size of 1 u is therefore equal to $\frac{1}{12}$ of the mass of a carbon-12 atom, which is approximately the mass of the smallest atom, hydrogen.

The volume of 1 mole of any gas at normal atmospheric pressure (101.3 kPa) and a temperature of 0°C is 22.4 litres (L).

Avogadro's hypothesis

The simplified version of chemistry given here does not give the full picture: for one thing, atoms don't always join in pairs. Maybe one A joins with two Bs. Without knowing the ratio of how many atoms of B join with one atom of A we can't calculate the relative masses of the individual atoms. Amedeo Avogadro solved this problem by suggesting that equal volumes of all gases at the same temperature and pressure will contain the same number of molecules. So if one atom of A and one atom of B join to give one molecule AB then the number of molecules of AB is equal to the number of atoms of A or B and the volume of AB = $\frac{1}{2}$(A + B); but if one atom of A joins with two atoms of B then the volume of B atoms is twice the volume of the A atoms, and the volume of AB_2 = the volume of A. This is illustrated in Figure 3.5.

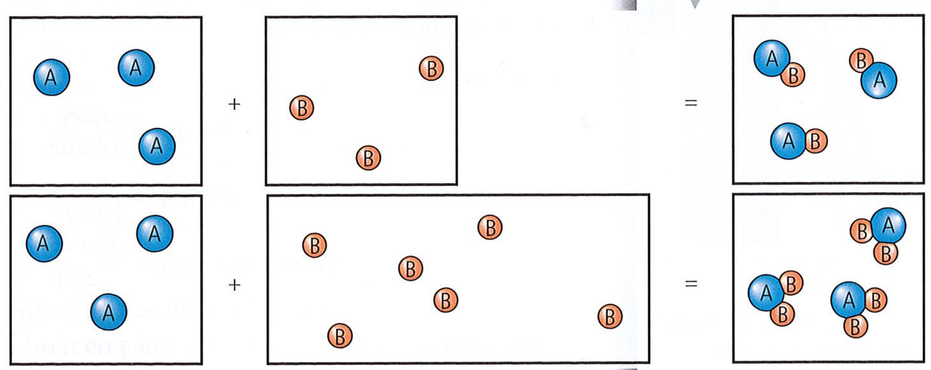

Figure 3.5 Equal volumes of gases contain the same number of molecules.

calculation of
; a grain of sand is
ed to volume of
veals that there are
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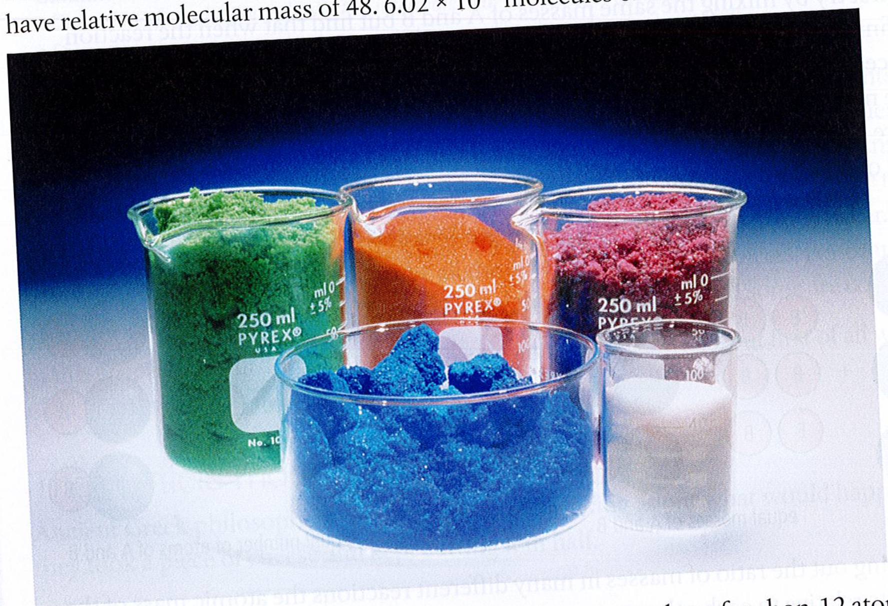
of sand in the
desert.

The mole and Avogadro's constant

It can be shown that 12 g of carbon-12 contains 6.02 × 10^{23} atoms. This amount of material is called a mole; this number of atoms is called Avogadro's constant (N_A) (named after him but not calculated by him). If we take 6.02 × 10^{23} molecules of a substance that has molecules that are four times the mass of carbon-12 atoms it would have relative molecular mass of 48. 6.02 × 10^{23} molecules of this substance would



es of different compounds
have different masses.

therefore have a mass four times the mass of the same number of carbon-12 atoms, 48 g. So, to calculate the mass of a mole of any substance we simply express its relative molecular mass in grams. This gives us a convenient way of measuring the amount of substance in terms of the number of molecules rather than its mass.

Worked example

If a mole of carbon has a mass of 12 g, how many atoms of carbon are there in 2 g?

Solution

One mole contains 6.022 × 10^{23} atoms.

2 g is $\frac{1}{6}$ of a mole so contains $\frac{1}{6}$ × 6.022 × 10^{23} atoms = 1.004 × 10^{23} atoms

Worked example

The density of iron is 7874 kg m^{-3} and the mass of a mole of iron is 55.85 g. What is the volume of 1 mole of iron?

Solution

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

$$\text{volume} = \frac{\text{mass}}{\text{density}}$$

$$\text{volume of 1 mole} = \frac{0.05585}{7874}\ \text{m}^3$$

$$= 7.093 \times 10^{-6}\ \text{m}^3$$

$$= 7.093\ \text{cm}^3$$

e careful with the
units. Do all calculations
using m^3.

Exercises

1 The mass of 1 mole of copper is 63.54 g and its density 8920 kg m^{-3}

(a) What is the volume of one mole of copper?

(b) How many atoms does one mole of copper contain?

(c) How much volume does one atom of copper occupy?

2 If the density of aluminium is 2700 kg m^{-3} and the volume of 1 mole is 10 cm^3, what is the mass of one mole of aluminium?

The three states of matter

Solid

A solid has a fixed shape and volume so the molecules must have a fixed position.

This means that if we try to pull the molecules apart there will be a force pulling them back together. This force is called the *intermolecular force*. This force is due to a property of the molecules called *charge* which will be dealt with properly in Chapter 6. This force doesn't only hold the particles together but also stops the particles getting too close. It is this force that is responsible for the tension in a string: as the molecules are pulled apart they pull back, and the normal reaction – when the surfaces are pushed together the molecules push back.

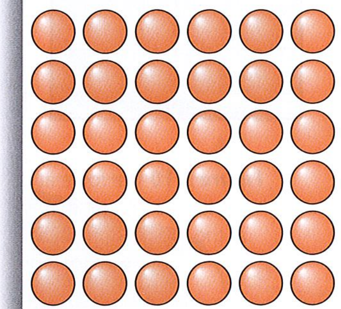


Figure 3.6 Molecules in a solid.

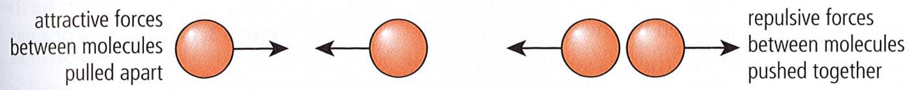


Figure 3.7 Intermolecular force.

Molecules of a solid are not free to move about but they can vibrate.

The molecules of a liquid are often drawn further apart than molecules of a solid but this isn't always the case; water is an example of the opposite. When ice turns to water it contracts which is why water is more dense than ice and ice floats on water.

Liquid

A liquid does not have a fixed shape but does have a fixed volume so the molecules are able to move about but still have an intermolecular force between then. This force is quite large when you try to push the molecules together (a liquid is very difficult to compress) but not so strong when pulling molecules apart (if you throw a bucket of water in the air it doesn't stay together).

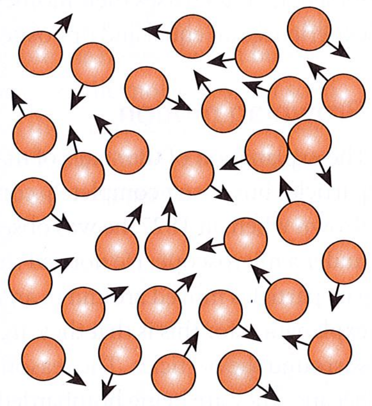


Figure 3.8 Molecules in a liquid.

When a liquid is put into a container it presses against the sides of the container. This is because of the intermolecular forces between the liquid and the container. At the bottom of the container the molecules are forced together by the weight of liquid above. This results in a bigger force per unit area on the sides of the container; this can be demonstrated by drilling holes in the side of the container and watching the water squirt out as shown in Figure 3.9. The pressure under a solid block also depends on the height of the block but this pressure only acts downwards on the ground not outwards or upwards as it does in a fluid.

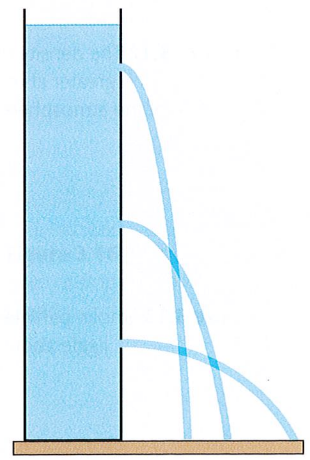


Figure 3.9 Higher pressure at the bottom of the container forces the water out further.

So the force per unit area, or *pressure*, increases with depth. If you have ever been diving you may have felt this as the water pushed against your ears. The increase in pressure with depth is also the reason why submerged objects experience a buoyant force.

If you consider the submerged cube shown in Figure 3.10, the bottom surface is deeper than the top so experiences a greater force, resulting in a resultant upward force. Note that this is only the case if the water is in a gravitational field e.g. on the Earth, as it relies on the weight of the water pushing down on the water below.

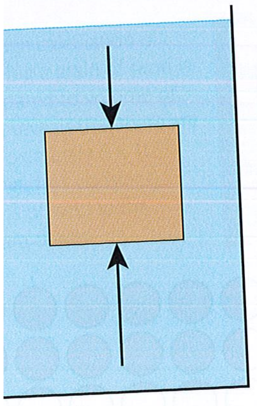


Figure 3.10 The force at the bottom is greater than at the top resulting in a buoyant force.

Gas

A gas does not have fixed shape or volume, it simply fills whatever container it is put into. The molecules of a gas are completely free to move about without any forces between the molecules except when they are colliding.

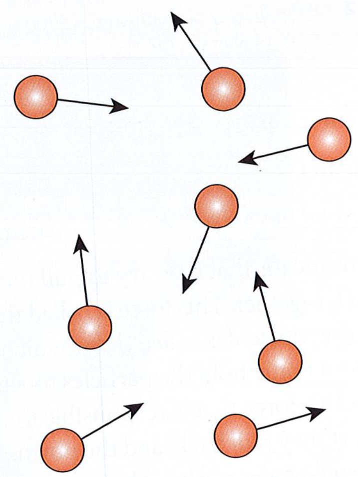


Figure 3.11 Molecules of a gas.

Since the molecules of a gas are moving they collide with the wall of the container. The change in momentum experienced by the gas molecules means that they must be subjected to an unbalanced force resulting in an equal and opposite force on the container. This results in gas pressure. If the gas is on the Earth then the effect of gravity will cause the gas nearest the ground to be compressed by the gas above. This increases the density of the gas so there are more collisions between the gas molecules and the container, resulting in a higher pressure. The difference between the pressure at the top of an object and the pressure at the bottom results in a buoyant force is illustrated in Figure 3.12.

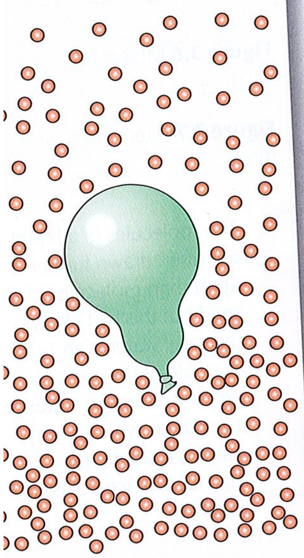


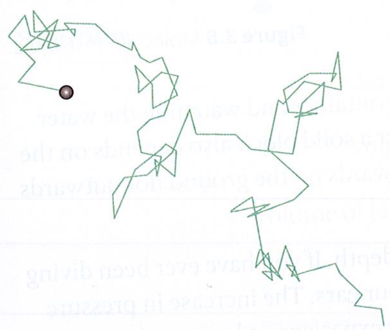
Figure 3.12 The density of air molecules is greater at the bottom of the atmosphere.

A car moving through air will collide with the air molecules. As the car hits the air molecules it increases their momentum so they must experience a force. The car experiences an equal and opposite force which we call air resistance or *drag*.

Brownian motion

The explanation of the states of matter supports the theory that matter is made of particles but it isn't completely convincing. More solid evidence was found by Robert Brown when, in 1827, he was observing a drop of water containing pollen grains under a microscope. He noticed that the pollen grains had an unusual movement. The particles moved around in an erratic zigzag pattern similar to Figure 3.13. The explanation for this is that they are being hit by the invisible molecules of water that surround the particles. The reason we don't see this random motion in larger objects is because they are being bombarded from all sides so the effect cancels out.

Figure 3.13 Smoke particles jiggle about.



Internal energy

In Chapter 2 we considered a car moving along the road at constant velocity.

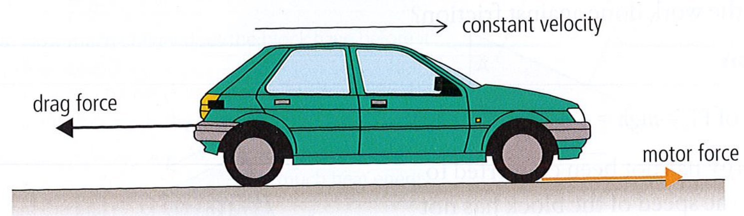


Figure 3.14 Forces on a car moving with constant velocity.

The force of the motor is applied via the friction between the tyres and the road. In this example we calculated the energy transferred from the petrol to the car so the motor is doing work but since the car is not getting any faster or going up a hill it is not gaining any kinetic or potential energy. If we consider the air to be made up of particles we can answer the question of what is happening to the energy and thus gain a better understanding of drag force.

As the car moves through the air it collides with the molecules of air as in Figure 3.15. When a collision is made the car exerts a force on the molecules so according to Newton's third law the car must experience an equal and opposite force. This is the drag force.

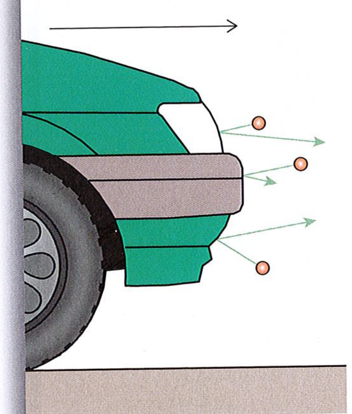


Figure 3.15 The front of the car collides with air molecules.

As the car moves forward hitting the air molecules it gives them kinetic energy; this is where all the energy is going. We call this *internal energy*, and since gas molecules have no forces between them this energy is all KE.

Another example we could consider is a block sliding down a slope at a constant speed as in Figure 3.16. As it slides down the slope it is losing PE but where is the energy going? This time it is the friction between the block and the slope that provides the answer. As the surfaces rub against each other energy is transferred to the molecules of the block and slope; the rougher the surfaces the more the molecules get bumped about. The effect of all this bumping is to increase the KE of the molecules, but solid molecules can't fly about; they can only vibrate and as they do this they move apart. This moving apart requires energy because the molecules have a force holding them together; the result is an increase in both kinetic and potential energy.

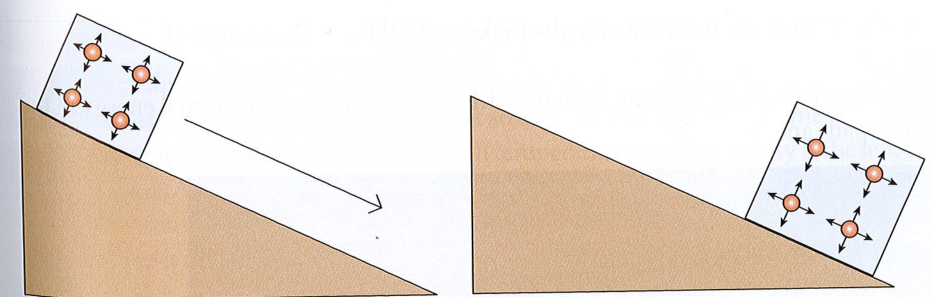


Figure 3.16 A block slides down a slope at constant speed.

Worked example

A 4 kg block slides down the slope at a constant speed of $1\,\text{m s}^{-1}$ as in Figure 3.17. What is the work done against friction?

Solution

The loss of PE $= mgh = 4 \times 10 \times 3 = 120\,\text{J}$.

This energy has not been converted to KE since the speed of the block has not increased. The energy has been given to the internal energy of the slope and block. The work done against friction (friction force × distance travelled in direction of the force) is therefore 120 J. The block is losing energy so this should be negative.

$$\text{Friction} \times 5 = -120\,\text{J}$$

$$\text{So friction} = -24\,\text{N}$$

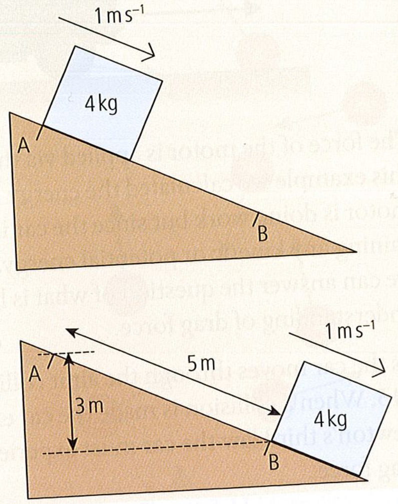


Figure 3.17 A block slides down a slope.

Worked example

A car of mass 1000 kg is travelling at $30\,\text{m s}^{-1}$. If the brakes are applied, how much heat energy is transferred to the brakes?

Solution

When the car is moving it has kinetic energy. This must be transferred to the brakes when the car stops.

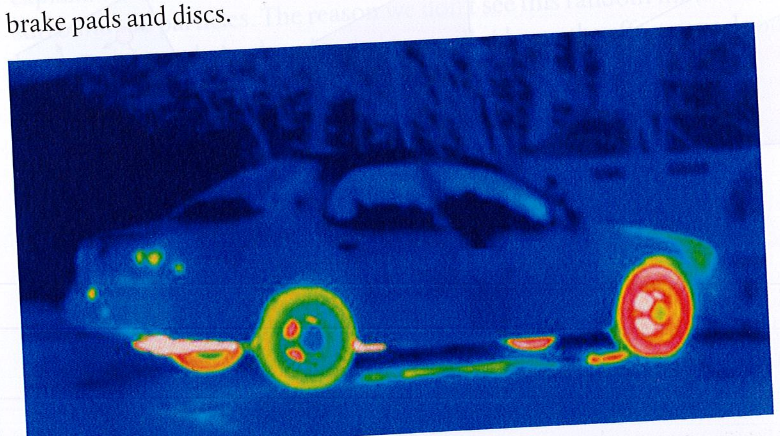
$$\text{KE} = \tfrac{1}{2}mv^2$$

$$= \tfrac{1}{2} \times 1000 \times 30^2\,\text{J}$$

$$= 450\,\text{kJ}$$

So thermal energy transferred to the brakes $= 450\,\text{kJ}$

When a car slows down using its brakes, KE will be converted to internal energy in the brake pads and discs.



This thermogram of a car shows how the wheels have become hot owing to friction between the road and the tyres, and the brakes pads

Exercises

3 A block of metal, mass 10 kg, is dropped from a height of 40 m.
  (a) How much energy does the block have before it is dropped?
  (b) How much heat energy do the block and floor gain when it hits the floor?

4 If the car in the second Worked example on page 100 was travelling at $60\,\text{m s}^{-1}$, how much heat energy would the brakes receive?

5 A 75 kg free fall parachutist falls at a constant speed of $50\,\text{m s}^{-1}$. Calculate the amount of energy given to the surroundings per second.

6 A block, starting at rest, slides down the slope as shown in Figure 3.18. Calculate the amount of work done against friction and the size of the friction force.

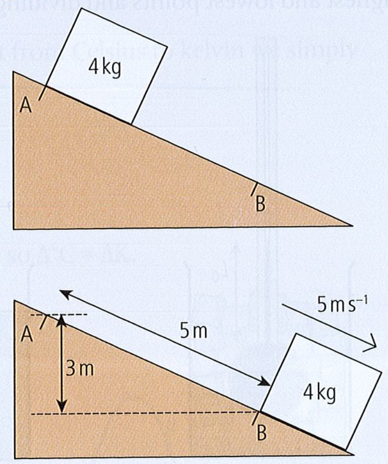


Figure 3.18.

Internal energy and the three states of matter

Internal energy is the sum of the energy of the molecules of a body. In solids and liquids there are forces between the molecules so to move them around requires work to be done (like stretching a spring). The total internal energy of solids and liquids is therefore made up of kinetic + potential energy. There is no force between molecules of a gas so changing their position does not require work to be done. The total internal energy of a gas is therefore only kinetic energy.

Temperature (T)

If we rub our hands together we are doing work since there is movement in the direction of the applied force. If work is done then energy must be transferred but we are not increasing the kinetic or potential energy of our hands; we are increasing their *internal energy*. As we do this we notice that our hands get hot. This is a sensation that we perceive through our senses and it seems from this simple experiment to be related to energy. The harder we rub the hotter our hands become. Before we can go further we need to define a quantity we can use to measure how hot or cold a body is.

Temperature is a measure of how hot or cold a body is.

When we defined a scale for length we simply took a known length and compared other lengths to it. With temperature it is not so easy. First we must find some directly measurable physical quantity that varies with temperature. One possibility is the length of a metal rod. As the internal energy of a solid temperature increases, the molecules move faster causing them to move apart. The problem is that the length doesn't change very much so it isn't easy to measure. A better alternative is to use the change in volume of a liquid. This also isn't very much but if the liquid is placed into a container with a thin tube attached as in Figure 3.19 then the change can be quite noticeable.

To define a scale we need two fixed points. In measuring length we use the two ends of a ruler; in this case we will measure the length of the liquid at two known temperatures, the boiling and freezing points of pure water. But how do we know that these events always take place at the same temperatures before we have made our thermometer? What we can do is place a tube of liquid in many containers of freezing and boiling water to see if the liquid always has the same lengths. If it does, then we can deduce that the freezing and boiling temperatures of water are always the same.

TOK We perceive how hot or cold something is with our senses but to quantify this we need a measurement.

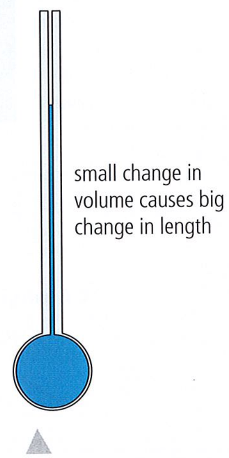


Figure 3.19 A simple thermometer.

NATURE OF SCIENCE

It is important to use pure water at normal atmospheric pressure. Otherwise the temperatures will not be quite right.

Having defined our fixed points we can make the scale by marking the tube at the highest and lowest points and dividing the range into 100 equal units.

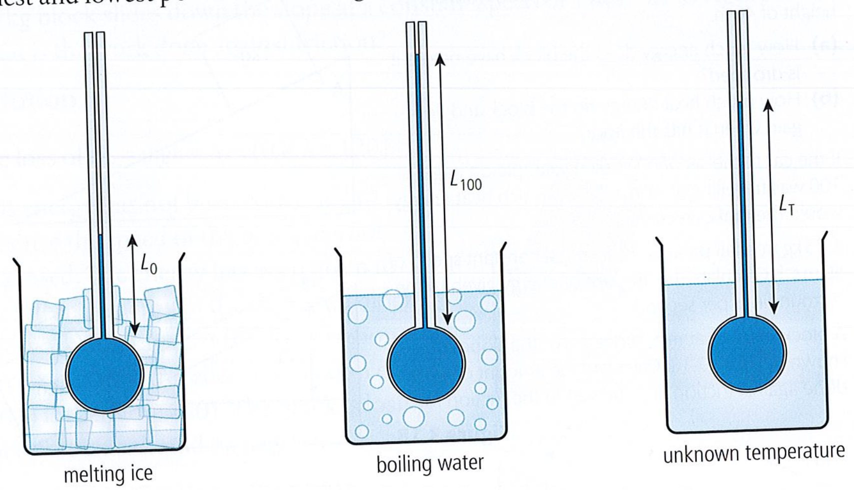


Figure 3.20 Calibrating a thermometer.

So if we place the thermometer into water at an unknown temperature resulting in length L_T, then the temperature can be found from:

$$T = \frac{L_T - L_0}{L_{100} - L_0} \times 100$$

This is how the Celsius scale is defined.

Temperature and KE

The reason that a liquid expands when it gets hot is because its molecules vibrate more and move apart. Higher temperature implies faster molecules so the temperature is directly related to the KE of the molecules. However, since the KE of the particles is not zero when ice freezes the KE is not directly proportional to the temperature in °C. The lowest possible temperature is the point at which the KE of molecules becomes zero. This happens at −273°C.

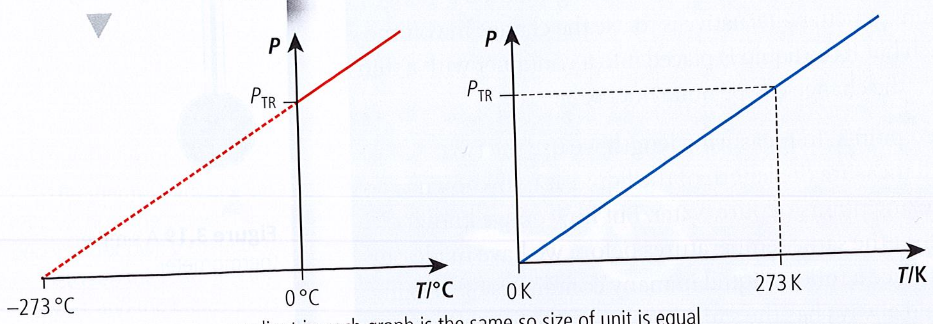
The lowest temperature is −273.15°C, but for conversion purposes we usually find it more convenient to use −273°C.

Kelvin scale

An alternative way to define a temperature scale would be to use the pressure of a constant volume of gas. As temperature increases, the KE of the molecules increases so they move faster. The faster moving molecules hit the walls of the container harder and more often, resulting in an increased pressure. As the temperature gets lower and lower the molecules slow down until at some point they stop moving completely. This is the lowest temperature possible or absolute zero. If we use this as the zero in our temperature scale then the KE is directly proportional to temperature. In defining this scale we then only need one fixed point, the other is absolute zero. This point could be the freezing point of water but the triple point is more precisely defined. This is the temperature at which water can be solid, liquid, and gas in equilibrium which in Celsius is 0.01°C. If we make this 273.16 in our new scale then a change of 1 unit will be the same as 1°C. This scale is called the kelvin scale.

Not all countries use the same units of temperature when describing the weather but the agreed-upon SI unit is the kelvin.

Figure 3.21 Pressure vs temperature in Celsius and kelvin.



Because the size of the unit is the same, to convert from Celsius to kelvin we simply add 273. So

$$10°C = 283\,K$$
$$50°C = 323\,K$$

A change from 10°C to 50°C is 50 − 10 = 40°C.

A change from 283 K to 323 K is 323 − 282 = 40 K, so $\Delta°C = \Delta K$.

Temperature and molecular speed

Now we have a temperature scale that begins at absolute zero we can say that, for an ideal gas, the average KE of the molecules is directly proportional to its temperature in kelvin. The constant of proportionality is $\frac{3}{2}k$ where k is the Boltzmann constant, $1.38 \times 10^{-23}\,m^2\,kg\,s^{-2}\,K^{-1}$.

Figure 3.22 shows the distribution of molecular velocities for different temperatures – the blue line describes molecular speed distribution of molecules in the air at about 0 °C, and the red one is at about 100 °C.

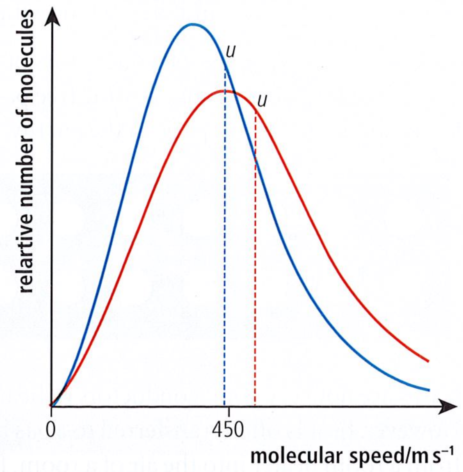


Figure 3.22 Molecular velocity distribution for a gas.

This is the *average* KE of the molecules. The different molecules of a gas travel at *different* velocities, some faster and some slower. The range of velocities can be represented by the velocity distribution curve in Figure 3.22. Because the curve is not symmetric, the mean value is to the *right* of centre.

Exercises

7 The length of a column of liquid is 30 cm at 100°C and 10 cm at 0°C. At what temperature will its length be 12 cm?

8 The average molar mass of air is 29 g mol^{-1}. Calculate:

(a) the average KE of air molecules at 20°C.
(b) the average mass of one molecule of air.
(c) the average speed of air molecules at 20°C.

Heat

We know that the temperature of a body is related to the average KE of its molecules and that the KE of the molecules can be increased by doing work, for example against friction, but the internal energy of a body can also be increased by putting it in contact with a hotter body. Energy transferred in this way is called *heat*.

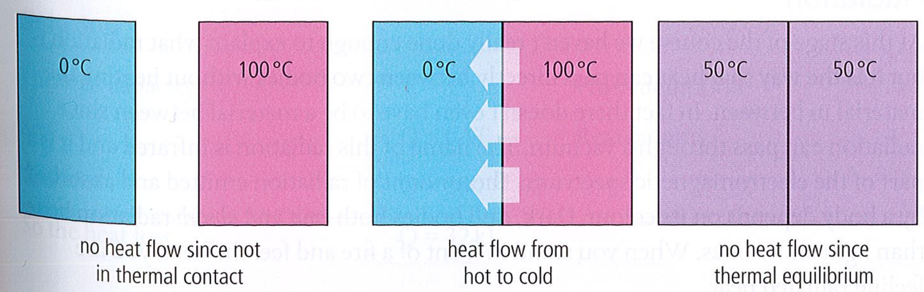


Figure 3.23 Heat travels from hot to cold until thermal equilibrium is established.

When bodies are in thermal contact heat will always flow from a high to a low temperature until the bodies are at the same temperature. Then we say they are in *thermal equilibrium*.

Heat transfer

There are three ways that heat can be transferred from one body to another; these are called conduction, convection, and radiation.

Conduction

Conduction takes place when bodies are in contact with each other. The vibrating molecules of one body collide with the molecules of the other. The fast-moving hot molecules lose energy and the slow-moving cold ones gain it.

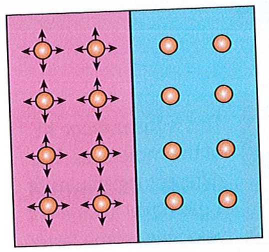


Figure 3.24 Energy passed from fast molecules to slow.

Metals are particularly good conductors of heat because not only are their atoms well connected but metals contain some free particles (electrons) that are able to move freely about, helping to pass on the energy.

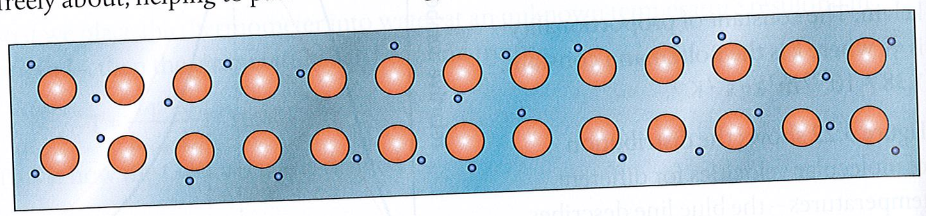


Figure 3.25 Electrons pass energy freely.

Gases are not very good conductors of heat because their molecules are far apart. However, heat is often transferred to a gas by conduction. This is how heat would pass from a room heater into the air of a room, for example.

Convection

This is the way that heat is transferred through fluids by fast-moving molecules moving from one place to another. When heat is given to air the molecules move around faster. This causes an increase in pressure in the hot air which enables it to expand, pushing aside the colder surrounding air. The hot air has now displaced more than its own weight of surrounding air so experiences an unbalanced upward force resulting in motion in that direction.

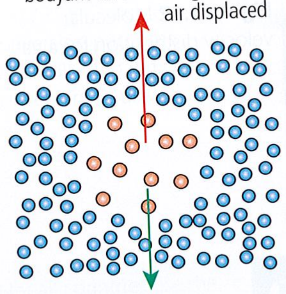


Figure 3.26 Hot air expands.

As the hot air rises it will cool and then come back down (this is also the way that a hot air balloon works). The circular motion of air is called a *convection current* and this is the way that heat is transferred around a room.

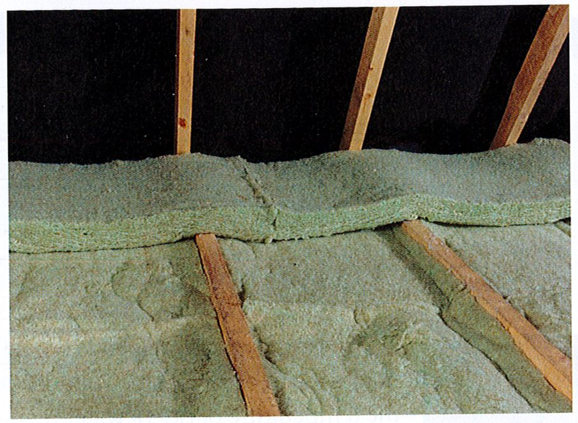
Although the first ever engine was probably a steam turbine, cylinders of expanding gas are the basis of most engines.

Radiation

At this stage of the course we haven't really done enough to explain what radiation is but it is the way that heat can pass directly between two bodies without heating the material in between. In fact there doesn't even have to be a material between since radiation can pass through a vacuum. The name of this radiation is infrared and it is a part of the electromagnetic spectrum. The amount of radiation emitted and absorbed by a body depends on its colour. Dark, dull bodies both *emit* and *absorb* radiation better than light shiny ones. When you stand in front of a fire and feel the heat, you are feeling radiated heat.

Preventing heat loss

In everyday life, as well as in the physics lab, we often concern ourselves with minimizing heat loss. Insulating materials are often made out of fibrous matter that traps pockets of air. The air is a poor conductor and when it is trapped it can't convect. Covering something with silver-coloured paper will reduce radiation.



Roof insulation in a house.

In cold countries houses are insulated to prevent heat from escaping. Are houses in hot countries insulated to stop heat entering?

Thermal capacity (C)

If heat is added to a body, its temperature rises, but the actual increase in temperature depends on the body.

The thermal capacity (C) of a body is the amount of heat needed to raise its temperature by 1°C. Unit: J°C^{-1} or J K^{-1}.

This applies not only when things are given heat, but also when they lose heat.

If the temperature of a body increases by an amount ΔT when quantity of heat Q is added, then the thermal capacity is given by the equation:

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

Worked example

If the thermal capacity of a quantity of water is 5000 J K^{-1}, how much heat is required to raise its temperature from 20°C to 100°C?

Solution

Thermal capacity	$C = \frac{Q}{\Delta T}$	From definition
So	$Q = C\Delta T$	Rearranging
Therefore	$Q = 5000 \times (100 - 20)$ J	
So the heat required	$Q = 400$ kJ	

Worked example

How much heat is lost from a block of metal of thermal capacity 800 J K^{-1} when it cools down from 60 °C to 20 °C?

Solution

Thermal capacity	$C = \frac{Q}{\Delta T}$	From definition
So	$Q = C\Delta T$	Rearranging
Therefore	$Q = 800 \times (60 - 20)$ J	
So the heat lost	$Q = 32$ kJ	

It is possible to buy a special shower head that uses less water. In some countries this is used to save energy; in others to save water.

Remember, power is energy per unit time.

Exercises

9 The thermal capacity of a 60 kg human is 210 kJ K^{-1}. How much heat is lost from a body if its temperature drops by 2 °C?

10 The temperature of a room is 10 °C. In 1hour the room is heated to 20 °C by a 1 kW electric heater.

(a) How much heat is delivered to the room?
(b) What is the thermal capacity of the room?
(c) Does all this heat go to heat the room?

Specific heat capacity (c)

The specific heat capacity of water is quite high, so it takes a lot of energy to heat up the water for a shower.

The thermal capacity depends on the size of the object and what it is made of. The *specific heat capacity* depends only on the material. Raising the temperature of 1 kg of water requires more heat than raising the temperature of 1 kg of steel by the same amount, so the specific heat capacity of water is higher than that of steel.

The specific heat capacity of a material is the amount of heat required to raise the temperature of 1 kg of the material by 1 °C. Unit: J kg^{-1} $°C^{-1}$ or J kg^{-1} K^{-1}.

If a quantity of heat Q is required to raise the temperature of a mass m of material by ΔT then the specific heat capacity (c) of that material is given by the following equation:

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

It takes 4200 J of energy to raise the temperature of water by 1 °C. This is equivalent to lifting 420 kg a height of 1 m. This makes water a good medium for transferring energy but also makes it expensive to take a shower. Oil would be cheaper to heat but not so good to wash in.

Worked example

The specific heat capacity of water is 4200 J kg^{-1} K^{-1}. How much heat will be required to heat 300 g of water from 20 °C to 60 °C?

Solution

Specific heat capacity $c = \frac{Q}{m\Delta T}$ From definition

So $Q = cm\Delta T$ Rearranging

Therefore $Q = 4200 \times 0.3 \times 40$ Note: Convert g to kg

$Q = 50.4$ kJ

Worked example

A metal block of mass 1.5 kg loses 20 kJ of heat. As this happens, its temperature drops from 60 °C to 45 °C. What is the specific heat capacity of the metal?

Solution

Specific heat capacity $c = \frac{Q}{m\Delta T}$ From definition

So $c = \frac{20\,000}{1.5(60-45)}$ Rearranging

$c = 888.9$ J kg^{-1} K^{-1}

Exercises

Use the data in Table 3.1 to solve the problems:

Substance	Specific heat capacity (J kg^{-1} K^{-1})
Water	4200
Copper	380
Aluminium	900
Steel	440

Table 3.1.

11 How much heat is required to raise the temperature of 250 g of copper from 20 °C to 160 °C?

12 The density of water is 1000 kg m^{-3}.

(a) What is the mass of 1 litre of water?
(b) How much energy will it take to raise the temperature of 1litre of water from 20 °C to 100 °C?
(c) A water heater has a power rating of 1 kW. How many seconds will this heater take to boil 1 litre of water?

13 A 500 g piece of aluminium is heated with a 500 W heater for 10 minutes.

(a) How much energy will be given to the aluminium in this time?
(b) If the temperature of the aluminium was 20 °C at the beginning, what will its temperature be after 10 minutes?

14 A car of mass 1500 kg travelling at 20 m s^{-1} brakes suddenly and comes to a stop.

(a) How much KE does the car lose?
(b) If 75% of the energy is given to the front brakes, how much energy will they receive?
(c) The brakes are made out of steel and have a total mass of 10 kg. By how much will their temperature rise?

15 The water comes out of a showerhead at a temperature of 50 °C at a rate of 8 litres per minute.

(a) If you take a shower lasting 10 minutes, how many kg of water have you used?
(b) If the water must be heated from 10 °C, how much energy is needed to heat the water?

Phase change

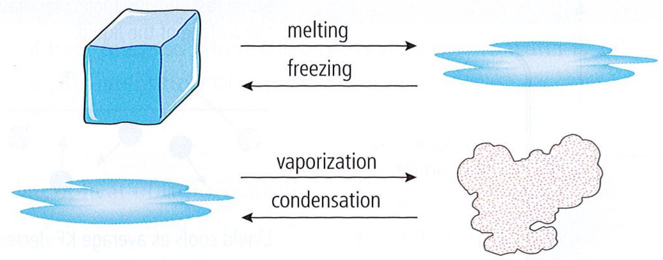


Figure 3.27 When matter changes from liquid to gas, or solid to liquid, it is changing state.

When water boils, this is called a *change of state* (or *change of phase*). As this happens, the temperature of the water doesn't change – it stays at 100 °C. In fact, we find that whenever the state of a material changes, the temperature stays the same. We can explain this in terms of the particle model.

An iceberg melts as it floats into warmer water.

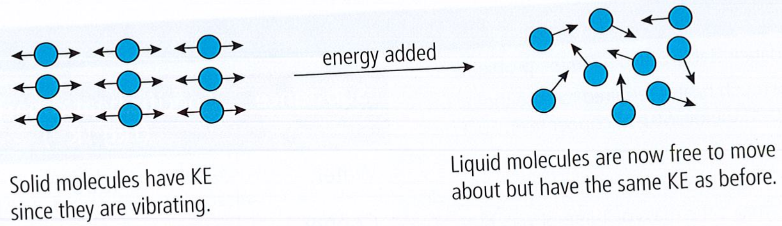


Figure 3.28 Molecules gain PE when the state changes.

When matter changes state, the energy is needed to enable the molecules to move more freely. To understand this, consider the example below.

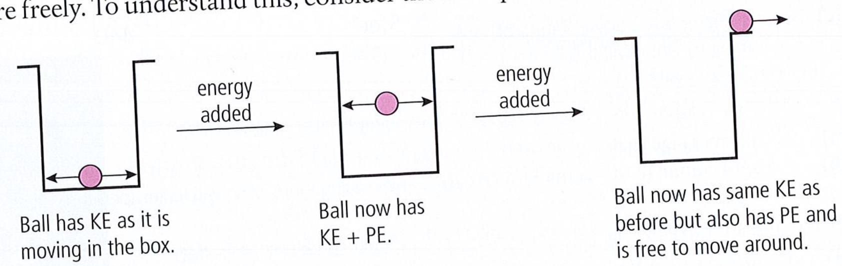


Figure 3.29 A ball-in-a-box model of change of state.

Boiling and evaporation

These are two different processes by which liquids can change to gases.

Boiling takes place throughout the liquid and always at the same temperature. *Evaporation* takes place only at the surface of the liquid and can happen at all temperatures.

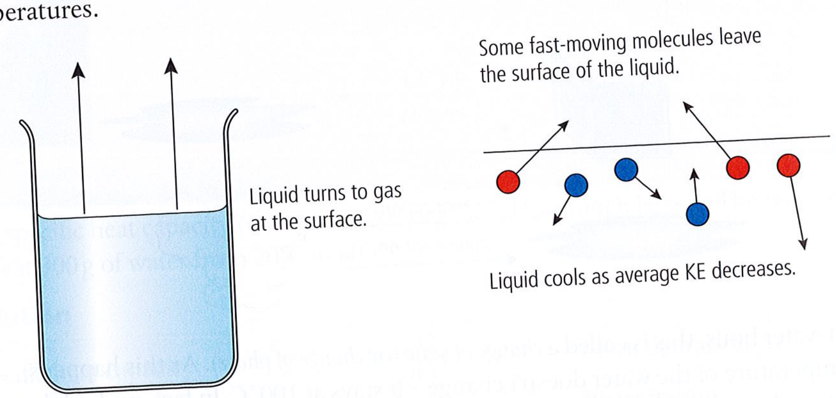


Figure 3.30 A microscopic model of evaporation.

When a liquid evaporates, the fastest-moving particles leave the surface. This means that the average kinetic energy of the remaining particles is lower, resulting in a drop in temperature.

The rate of evaporation can be increased by:

- increasing the surface area; this increases the number of molecules near the surface, giving more of them a chance to escape.
- blowing across the surface. After molecules have left the surface they form a small 'vapour cloud' above the liquid. If this is blown away, it allows further molecules to leave the surface more easily.
- raising the temperature; this increases the kinetic energy of the liquid molecules, enabling more to escape.

People sweat to increase the rate at which they lose heat. When you get hot, sweat comes out of your skin onto the surface of your body. When the sweat evaporates, it cools you down. In a sauna there is so much water vapour in the air that the sweat doesn't evaporate.

Specific latent heat (L)

The *specific latent* heat of a material is the amount of heat required to change the state of 1 kg of the material without change of temperature.

Unit: $\mathrm{J\,kg^{-1}}$

Latent means *hidden*. This name is used because when matter changes state, the heat added does not cause the temperature to rise, but seems to disappear.

If it takes an amount of energy Q to change the state of a mass m of a substance, then the specific latent heat of that substance is given by the equation:

$$L = \frac{Q}{m}$$

Solid → liquid

Specific latent heat of fusion

Liquid → gas

Specific latent heat of vaporization

Worked example

The specific latent heat of fusion of water is $3.35 \times 10^5\,\mathrm{J\,kg^{-1}}$. How much energy is required to change 500 g of ice into water?

Solution

The latent heat of fusion	$L = \frac{Q}{m}$	From definition
So	$Q = mL$	Rearranging
Therefore	$Q = 0.5 \times 3.35 \times 10^5\,\mathrm{J}$	
So the heat required	$Q = 1.675 \times 10^5\,\mathrm{J}$	

Worked example

The amount of heat released when 100 g of steam turns to water is $2.27 \times 10^5\,\mathrm{J}$. What is the specific latent heat of vaporization of water?

Solution

The specific latent heat of vaporization	$L = \frac{Q}{m}$	From definition
Therefore	$L = 2.27 \times 10^5/0.1\,\mathrm{J\,kg^{-1}}$	
So the specific latent heat of vaporization	$L = 2.27 \times 10^6\,\mathrm{J\,kg^{-1}}$	

This equation ($L = \frac{Q}{m}$) can also be used to calculate the heat lost when a substance changes from gas to liquid, or liquid to solid.

Exercises

Use the data about water in Table 3.2 to solve the following problems.

16 If the mass of water in a cloud is 1 million kg, how much energy will be released if the cloud turns from water to ice?

17 A water boiler has a power rating of 800 W. How long will it take to turn 400 g of boiling water into steam?

18 The ice covering a $1000\,\mathrm{m^2}$ lake is 2 cm thick.

(a) If the density of ice is $920\,\mathrm{kg\,m^{-3}}$, what is the mass of the ice on the lake?
(b) How much energy is required to melt the ice?
(c) If the Sun melts the ice in 5 hours, what is the power delivered to the lake?
(d) How much power does the Sun deliver per $\mathrm{m^2}$?

Latent heat of vaporization	$2.27 \times 10^6\,\mathrm{J\,kg^{-1}}$
Latent heat of fusion	$3.35 \times 10^5\,\mathrm{J\,kg^{-1}}$

Table 3.2 Latent heats of water.

Before the invention of the refrigerator, people would collect ice in the winter, and store it in well-insulated rooms so that it could be used to make ice cream in the summer. The reason it takes so long to melt is because to melt 1 kg of ice requires 3.3×10^5 J of energy; in a well-insulated room this could take many months.

In this example, we are ignoring the heat given to the kettle and the heat lost.

Graphical representation of heating

The increase of the temperature of a body can be represented by a temperature–time graph. Observing this graph can give us a lot of information about the heating process.

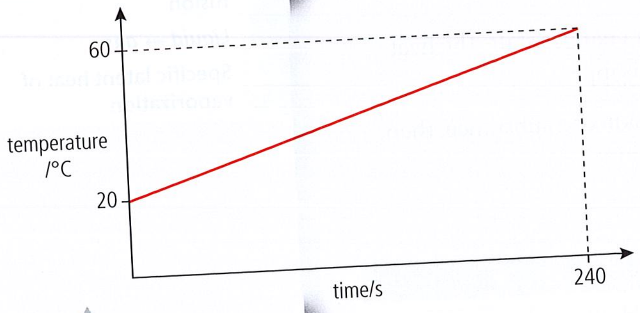


Figure 3.31 Temperature–time graph for 1 kg of water being heated in an electric kettle.

From this graph we can calculate the amount of heat given to the water per unit time (power).

$$\text{The gradient of the graph} = \frac{\text{temperature rise}}{\text{time}} = \frac{\Delta T}{\Delta t}$$

We know from the definition of specific heat capacity that

$$\text{heat added} = mc\Delta T$$

$$\text{The rate of adding heat} = P = \frac{mc\Delta T}{\Delta t}$$

So $P = mc \times \text{gradient}$

$$\text{The gradient of this line} = \frac{(60-20)}{240}\,°\text{C s}^{-1} = 0.167\,°\text{C s}^{-1}$$

So the power delivered $= 4200 \times 0.167\ \text{W} = 700\ \text{W}$

If we continue to heat this water it will begin to boil.

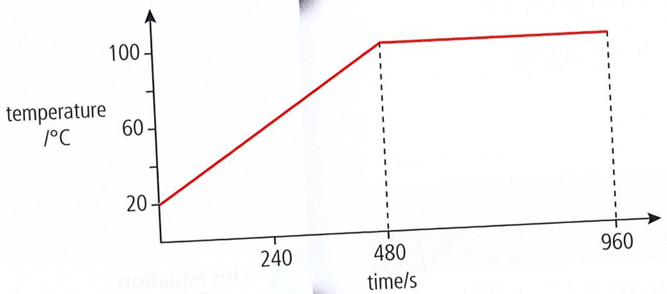


Figure 3.32 A graph of temperature vs time for boiling water. When the water is boiling, the temperature does not increase any more.

If we assume that the heater is giving heat to the water at the same rate, then we can calculate how much heat was given to the water whilst it was boiling.

Power of the heater = 700 W

Time of boiling = 480 s

Energy supplied = power × time = $700 \times 480\ \text{J} = 3.36 \times 10^5\ \text{J}$

From this we can calculate how much water must have turned to steam.

Heat added to change state = mass × latent heat of vaporization,

where latent heat of vaporization of water $= 2.27 \times 10^6\ \text{J kg}^{-1}$.

$$\text{Mass changed to steam} = \frac{3.36 \times 10^5}{2.27 \times 10^6} = 0.15\ \text{kg}$$

The amount of heat loss is proportional to the difference between the temperature of the kettle and its surroundings. For this reason, a graph of temperature against time is actually a curve, as shown in Figure 3.33. The fact that the gradient decreases tells us that the amount of heat given to the water gets less with time. This is because as it gets hotter, more and more of the heat is lost to the room.

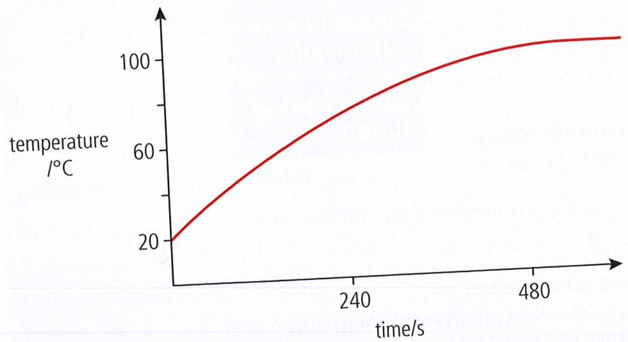


Figure 3.33 Heat loss.

Measuring thermal quantities by the method of mixtures

The method of mixtures can be used to measure the specific heat capacity and specific latent heat of substances.

Specific heat capacity of a metal

A metal sample is first heated to a known temperature. The most convenient way of doing this is to place it in boiling water for a few minutes; after this time it will be at 100 °C. The hot metal is then quickly moved to an insulated cup containing a known mass of cold water. The hot metal will cause the temperature of the cold water to rise; the rise in temperature is measured with a thermometer. Some example temperatures and masses are given in Figure 3.34.

As the specific heat capacity of water is $4180\ \text{J kg}^{-1}\ \text{K}^{-1}$, we can calculate the specific heat capacity of the metal.

ΔT for the metal = 100 − 15 = 85 °C

and ΔT for the water = 15 − 10 = 5 °C

Applying the formula $Q = mc\Delta T$ we get

$(mc\Delta T)_{metal} = 0.1 \times c \times 85 = 8.5c$

$(mc\Delta T)_{water} = 0.4 \times 4180 \times 5 = 8360$

If no heat is lost, then the heat transferred from the metal = heat transferred to the water

$8.5c = 8360$

$c_{metal} = 983\ \text{J kg}^{-1}\ \text{K}^{-1}$

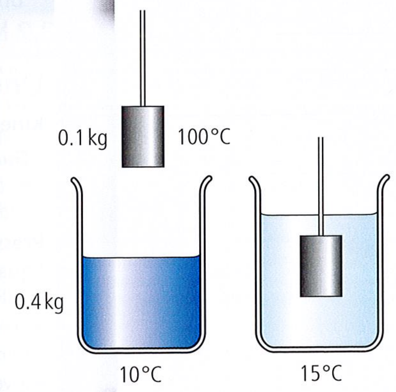


Figure 3.34 Measuring the specific heat capacity of a metal.

Measurement of the specific heat capacity of a metal by the method of mixtures

A worksheet with full details of how to carry out this experiment is available in your eBook.

Latent heat of vaporization of water

To measure the latent heat of vaporization, steam is passed into cold water. Some of the steam condenses in the water, causing the water temperature to rise.

The heat from the steam = the heat to the water.

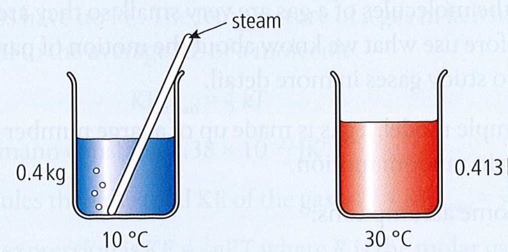


Figure 3.35 By measuring the rise in temperature, the specific latent heat can be calculated.

In Figure 3.35, 13 g of steam have condensed in the water, raising its temperature by 20 °C. The steam condenses then cools down from 100 °C to 30 °C.

Heat from steam = $ml_{steam} + mc\Delta T_{water}$

$0.013 \times L + 0.013 \times 4.18 \times 10^3 \times 70 = 0.013L + 3803.8$

Heat transferred to cold water = $mc\Delta T_{water} = 0.4 \times 4.18 \times 10^3 \times 20$

$= 33\,440\ \text{J}$

Since heat from steam = heat to water

$0.013L + 3803.8 = 33\,440$

So $L = \dfrac{33\,440 - 3803.8}{0.013}$

$L = 2.28 \times 10^6\ \text{J kg}^{-1}$

When melting sugar to make confectionary be very careful: liquid sugar takes much longer to cool down than you might think. This is because as it changes from liquid to solid it is giving out heat but doesn't change temperature. You should wait a long time before trying to pick up any of your treats with your fingers.

Heat loss

In both of these experiments, some of the heat coming from the hot source can be lost to the surroundings. To reduce heat loss, the temperatures can be adjusted, so you could start the experiment below room temperature and end the same amount above (e.g. if room temperature is 20 °C, then you can start at 10 °C and end at 30 °C).

To learn more about thermal concepts, go to the hotlinks site, search for the title or ISBN and click on Chapter 3.