Unit 12

Heat

This unit online, with its interactives:
science-playground.com/s/10a12

On a sunny winter morning in Bhaktapur, the rooftops fill with washing. By noon the clothes are dry, though the air never got anywhere near the 100 °C at which water boils. Where did the water go, and why do the clothes on a breezy roof dry before the ones in a sheltered courtyard?

The same morning, the steel rails of a railway in the Terai are a few millimeters shorter than they will be by mid-afternoon. A kettle's steam can scald a hand far worse than a splash of the boiling water it came from. And a lump of ice dropped into a glass of juice cools it much better than the same mass of ice-cold water would. All four are about what heat does to matter when it isn't simply making it hotter: it lets molecules escape, it pushes them apart, and it breaks them free of each other. This unit is about those three things, and about the one equation that keeps track of every joule.



Figure 12.1 *Washing day at Kumari Chowk in Kathmandu. Spread out in the sun, the clothes will be dry in a few hours: the water in them turns into vapor and goes into the air, molecule by molecule, far below its boiling point.*

Photo: erikjohansson from Göteborg, Sweden, CC BY 2.0

12.1 Evaporation

Put a spoonful of water in a steel plate and leave it on a windowsill. After a few hours the plate is dry. The water hasn't soaked into anything and hasn't boiled; it has gone into the air as an invisible gas, water vapor. This is **evaporation**.

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

Water evaporates from a puddle at 15 °C, from wet clothes at 25 °C and from your skin at 34 °C. Spirit, petrol and nail-polish remover evaporate even faster. In every case the vapor leaves only from the surface, the boundary between the liquid and the air.

Why a liquid can evaporate below its boiling point

The molecules of a liquid are close together and pull on each other, which is what holds the liquid together. They are also always moving, sliding past each other and colliding millions of millions of times a second. In those collisions some molecules gain speed and others lose it, so at any moment the molecules have a whole **range of**

speeds: most are near the average, a few are slow, and a few are very fast. The temperature tells us only the **average** kinetic energy of the molecules.

Now think of a molecule right at the surface. Its neighbors are all below it and beside it, and they pull it back down. If it happens to be one of the very fast ones, and happens to be moving upward, it can break free of that pull and fly off into the air. That is evaporation, one molecule at a time.

The cooling effect of evaporation

Here is the important part. The molecules that escape are the **fastest** ones. Take the fastest out of any group, and the average speed of those left behind goes down. A lower average kinetic energy means a **lower temperature**. So:

Evaporation cools the liquid that is left behind, and anything the liquid touches.

Once the liquid is cooler than its surroundings, heat flows into it from them: from the air, from the plate, from your skin. That heat keeps the evaporation going, and the surroundings cool too. Every kilogram of water that evaporates carries away a lot of energy: about 2,400,000 J at the temperature of your skin (more than the 2,260,000 J needed at 100 °C: in cooler water the molecules have less energy of their own, so each needs more to break free). You will meet this energy again in 12.3 as latent heat.

The cooling effect around us

Sweating. When you are hot, your skin releases sweat, and as it evaporates it takes heat from your body. Heavy work in hot weather can make a person sweat a liter an hour; evaporating that much carries away about 2.4 million joules, more than 650 joules every second, several times what the body makes at rest. Sweat that drips off without evaporating cools you hardly at all, which is why a hot, humid day feels so much worse than a hot, dry one.

Dogs and buffaloes. A dog can't sweat much, so it pants: air rushing over its wet tongue and mouth evaporates water from them. A buffalo wallows in mud and water on a hot afternoon, and the water evaporating from its skin cools it.

A wet cloth on the forehead. A cloth soaked in water and laid on the forehead of someone with a high fever cools them by evaporation, as well as by the water warming up.

The *matka*. Water seeps through the unglazed clay of a pot and evaporates from its outside, so the water inside stays cooler than the room. It works best in a dry, airy place.

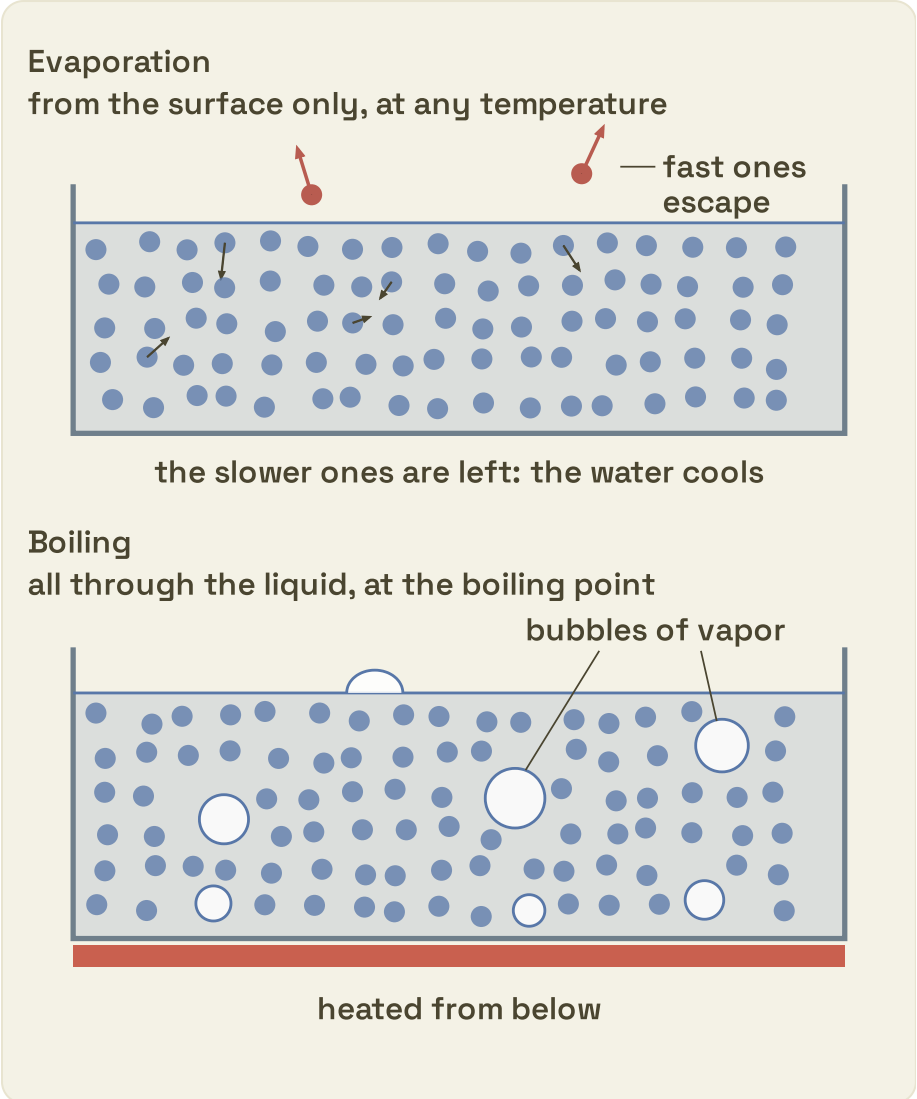


Figure 12.2 *Top: in evaporation, only molecules at the surface can escape, and only the fastest ones. Bottom: in boiling, the liquid is at its boiling point and bubbles of vapor form inside it, all through the liquid, starting where it is heated.*

Cooling the courtyard. Sprinkling water on a hot floor or a courtyard on a summer afternoon cools it, and a breeze blowing over wet ground feels cool. An evaporative ("desert") cooler blows air through wet pads for the same reason.

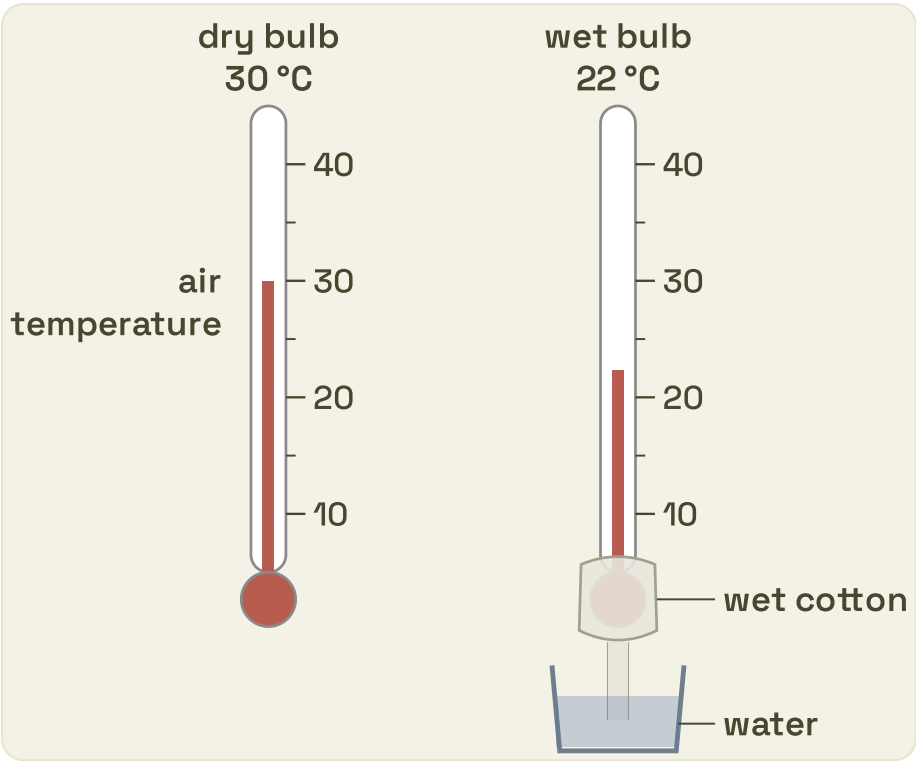


Figure 12.3 *A dry thermometer and a wet one side by side. In air at 30 °C with a relative humidity of 50%, the wet bulb settles at about 22 °C: evaporation from the cotton keeps it 8 °C cooler than the air around it.*

Check 1

A drop of spirit on your hand feels colder than a drop of water at the same temperature. Why?

a spirit is always colder than water

b spirit evaporates faster, so it takes heat from your skin faster

c spirit conducts heat better than water

d spirit has a higher boiling point than water

Under a tree. Leaves lose water through their pores (transpiration), and its evaporation keeps a tree's shade cooler than the shade of a tin roof.

CAREFUL

Evaporation needs heat; it doesn't make "cold"

A common wrong answer is that evaporation "produces cold". Cold isn't a thing that can be made or flow. Evaporation takes **heat** away: the escaping molecules carry energy with them, and heat flows from the surroundings into the liquid to replace it. So the surroundings cool.

Factors affecting the rate of evaporation

Why does washing dry in an hour on some days and stay damp all day on others? The **rate of evaporation** (how much liquid evaporates each minute) depends on six things. Each one comes straight from the picture of fast molecules escaping from the surface.

1. Temperature of the liquid. At a higher temperature the molecules move faster on average, so many more of them are fast enough to escape. Into the same dry air, water at 40 °C evaporates about six times as fast as water at 10 °C. That is why clothes dry faster in the sun.

2. Surface area. Molecules escape only from the surface, so the bigger the surface, the more escape each second. Tea in a saucer cools faster than tea in a cup; wet clothes spread out dry faster than clothes left in a heap; farmers spread paddy and chilies in a thin layer on mats to dry them.

3. Humidity of the air. The vapor in the air is made of molecules moving at random too, and some of them hit the liquid's surface and are caught again. In dry air few come back; in humid air (air already holding a lot of water vapor) almost as many come back as leave, so the liquid evaporates slowly. When the air holds all the vapor it can at that temperature (100% humidity), water evaporates no faster than vapor comes back, and wet things don't dry at all. Clothes dry slowly in the monsoon, quickly in the dry days of Mangsir.

4. Movement of the air (wind). In still air, the molecules that escape gather just above the surface, and many fall back in. A breeze sweeps them away and brings drier air, so evaporation goes faster. That is why we fan a hot face and why clothes dry fastest on a windy day.

5. Nature of the liquid. Liquids whose molecules pull on each other more weakly lose them more easily. Spirit, petrol, acetone and ether evaporate quickly (they are called **volatile** liquids); water evaporates more slowly; mustard oil hardly evaporates at all.

6. Air pressure. The lower the pressure of the air above the liquid, the more easily its molecules escape. Evaporation is faster high in the mountains, where the air pressure is low, as long as it isn't too cold.

The times in the widget are a model's estimates, not measurements: they come from a formula engineers use for water evaporating from a surface, with the towel's wet area guessed, and a real towel slows down as its last water dries, so take them as rough guides to which way each change works and by how much. The towel hangs in the shade, so, just like the wet thermometer, it cools as it dries, until the heat flowing in from the air pays for the evaporation: look at its temperature. In the sun it would be warmer and dry faster, but every factor works the same way.

Check 2

Which of these would make wet clothes dry **more slowly**?

a hanging them in a breeze
b spreading them out
c drying them indoors on a humid monsoon day
d hanging them in the sun

DID YOU KNOW?

Salt from the sea, by evaporation

Most of the world's salt from the sea is made by letting the sun and wind do the work. Sea water is let into wide, shallow ponds, the largest possible surface, in hot, dry, windy places. The water evaporates, the salt can't, and it is left behind as white crystals to be raked up. Nepal has no sea: its salt is brought in, mostly from India, and iodized.



Figure 12.4 *Salt pans at Marakkanam on the coast of Tamil Nadu, India. Sea water evaporates from the wide, shallow ponds, and the salt left behind is raked up.*

Photo: Dey.sandip, CC BY 3.0

Evaporation and boiling

Evaporation and boiling both turn a liquid into vapor, but they are not the same process. When water is heated in a pan, it evaporates from its surface all the time, faster as it gets hotter. At its boiling point (100 °C at sea level, about 95 °C in Bhaktapur, where the air pressure is lower) something new happens: the molecules inside the liquid are energetic enough to form bubbles of vapor whose pressure can push back the air pressure on the water. Bubbles form all through the liquid, mostly where it is heated, rise and burst. That is **boiling**.

Boiling is the rapid change of a liquid into vapor, with bubbles forming all through it, at a fixed temperature (its boiling point) for a given pressure.

While water boils, its temperature stays at the boiling point however strongly it is heated: the extra heat only makes it boil away faster.

Evaporation	Boiling
Takes place only at the surface of the liquid.	Takes place all through the liquid, with bubbles.
Happens at any temperature.	Happens only at the boiling point (for a given pressure).
Slow and quiet.	Fast and vigorous.
Cools the liquid left behind.	The temperature stays at the boiling point while heat is supplied.
The heat comes from the liquid itself and its surroundings.	Needs heat supplied from outside, such as a flame.
Faster with a larger surface, wind and dry air.	Doesn't depend on surface area, wind or humidity.

Check 3

Which statement is true of **evaporation** but not of **boiling**?

a it turns a liquid into vapor

b it needs heat

c it happens at any temperature, only at the surface

d it happens faster when the liquid is hotter

12.2 Thermal expansion

When things are heated they usually get bigger, and when they cool they shrink. This is **thermal expansion** (and contraction). Grade 10 Science showed it with the ball and ring, and showed water's strange exception between 0 °C and 4 °C. Here we look at **why** it happens and at the different **kinds** of expansion.

Thermal expansion is the increase in the size of a substance when it is heated.

Expansion explained by the motion of molecules

In a solid, the molecules (or atoms) are held in fixed places by the forces between them, like balls joined to their neighbors by springs. They don't stay still: they vibrate about their places. When the solid is heated, they vibrate harder, with bigger swings.

Why should bigger swings make the solid bigger? Because the "springs" between molecules are lopsided. Pushing two molecules closer together than their usual spacing is very hard: they repel each other strongly as soon as they are squeezed. Pulling them apart is much easier: the attraction weakens gradually as they separate. So when a molecule swings harder, it can't go much closer to its neighbor, but it goes a good deal farther away. On average it sits a little farther from its neighbors than before. Every pair of neighbors gets a little farther apart, and the whole solid grows.

> CAREFUL
> **How to write it in an exam**
>
> Most books explain expansion in one line: "On heating, the molecules gain energy and vibrate with larger amplitude; they push each other farther apart, so the distance between them increases and the substance expands." That is the answer to write, and it is right as far as it goes. The fuller reason is the one above: the force between molecules is lopsided, so harder vibration moves their average spacing **outward**. (If the force pushed back equally both ways, the molecules would swing harder about the same middle place and the solid wouldn't expand at all.)

In a liquid, the molecules are close together but free to slide past each other. Heated, they move faster and jostle each other harder, and the liquid takes up more room. The forces holding a liquid together are weaker than in a solid, so a liquid expands more than a solid for the same warming.

In a gas, the molecules are far apart and fly about freely. Heated at a steady pressure, they move faster and hit the walls harder and more often, so the gas spreads into a

bigger space (Charles's law, from Grade 9 Optional Science). Gases expand far more than liquids, and liquids more than solids: warmed by the same amount, air expands about 15 times as much as water and about 100 times as much as steel.

Linear, superficial and cubical expansion

A heated solid grows in every direction at once. But often only one or two of its dimensions matter, so we name three kinds of expansion.

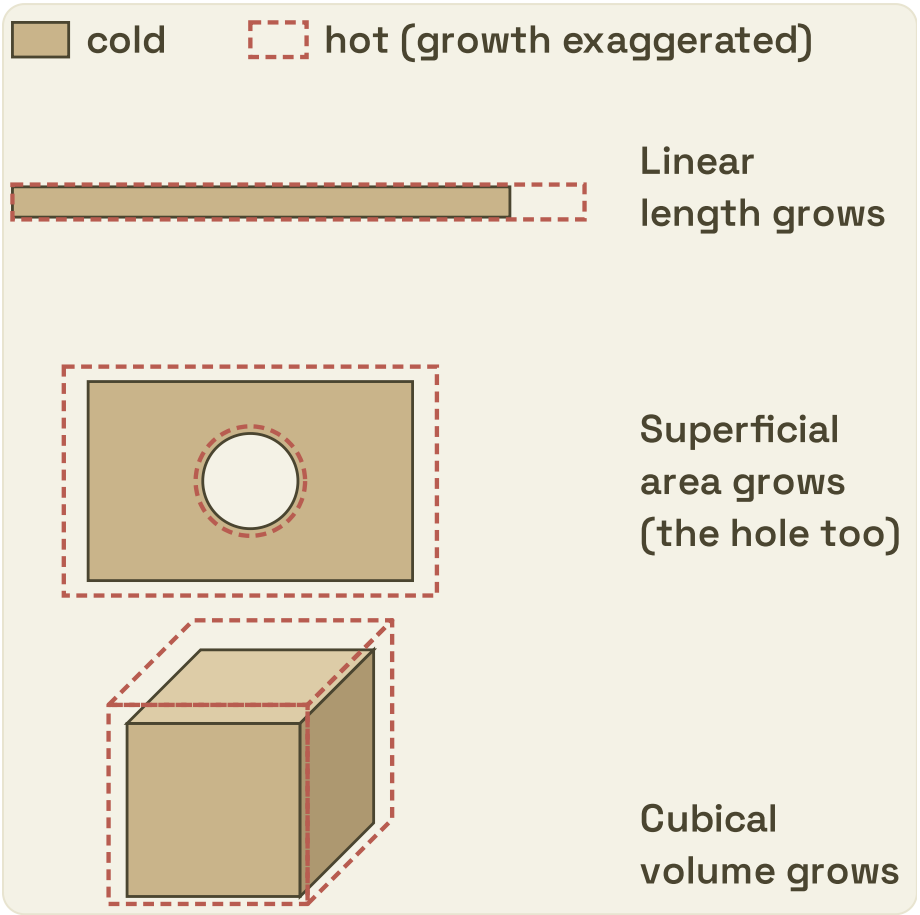


Figure 12.5 *The three kinds of thermal expansion. A rod gets longer (linear), a plate gets bigger in area (superficial), and a block gets bigger in volume (cubical). Notice that the hole in the plate grows too. The growth is drawn enormously exaggerated: heating steel by 100 °C makes it only about 0.1% bigger in each direction.*

Linear expansion is the increase in the **length** of a solid when it is heated. It matters for long, thin things: rails, wires, rods, pipes, bridges, the metal strips in a

thermostat.

Superficial expansion (also called areal expansion) is the increase in the **area** of a solid when it is heated. It matters for flat things: metal sheets and roofs, floor tiles, glass panes, a metal lid on a jar.

Cubical expansion (also called volume expansion) is the increase in the **volume** of a substance when it is heated. It matters for solid blocks, and it is the only kind of expansion a liquid or a gas can have, since they have no shape of their own.

All three are the same thing seen three ways: a heated solid grows in length, in breadth and in height. A plate that grows by 1% in length and 1% in breadth grows by about 2% in area; a block that grows by 1% each way grows by about 3% in volume.

CAREFUL

Does the hole get bigger or smaller?

Heat a metal plate with a hole in it, and the hole gets **bigger**, not smaller. Every part of the plate moves away from every other part, as if the whole plate, hole and all, were photographed and enlarged. The metal round the hole doesn't squeeze inward into it. That's why a tight metal lid comes off a jar when you run hot water over it: the lid's opening grows.

Check 4

A steel ring has a hole in the middle. When the ring is heated, the hole:

a gets smaller, because the steel expands into it
b gets bigger, by the same share as the steel
c stays the same size
d closes up completely

Check 5

Which kind of expansion can a liquid have?

a linear only
b superficial only
c cubical only
d all three

Impacts of thermal expansion in daily life

Thermal expansion is small, but the forces behind it are enormous. A steel rail heated by 40 °C, if held so it can't grow, pushes on whatever holds it with about 100 million newtons for every square meter of its cross-section, about the weight of a small car on every square centimeter. So engineers either leave room for expansion, or put it to work.

Railway tracks. Short rails are laid with small gaps between their ends. Modern long rails are welded into one piece kilometers long and fastened down very firmly to heavy sleepers, so they can't grow; special **expansion joints**, where two rail ends slide past each other at a slant, are put in where the long rail meets the track beyond. When the heat is more than the track was built for, the rails can buckle sideways, a dangerous **sun kink**.

Bridges and flyovers. A 1 km steel or concrete bridge grows by roughly 12 mm for each 1 °C it warms: about half a meter between a cold winter night and a hot summer afternoon in the Terai. So a long bridge rests on **rollers** at one end, and has **expansion joints** across the road: interlocking steel teeth, or steel beams with rubber seals between them, whose gaps open in winter and close in summer.



Figure 12.6 *Rails buckled by the heat, on five different tracks. Expansion in the hot sun pushed the long steel rails sideways out of line. Trains must stop until a buckled track is mended, or they can come off the rails.*

Photo: The original uploader was Trainwatcher at English Wikipedia ., Public domain

Concrete roads and floors are laid in slabs with thin gaps between them, filled with a soft, tar-like material that is squeezed when the slabs expand.

Electric and telephone wires are hung with a little sag. In winter they contract and pull tighter; if they were hung tight in summer, they would snap in the cold. In summer they sag more.

Glass. Hot tea poured into a thick, cold glass can crack it: the inside expands before the outside warms up. Glass for cooking and laboratory flasks (borosilicate glass) expands only about a third as much, so it survives sudden heating.



Figure 12.7 *An expansion joint across the road on a bridge over the Danube in Austria. As the bridge grows and shrinks with the temperature, the steel beams move closer together or farther apart, and the rubber seals between them squeeze or stretch.*

Photo: GT1976, CC BY-SA 4.0

Fitting things tightly. A blacksmith fits the iron tyre onto a wooden bullock-cart wheel while the tyre is red hot and slightly too big; as it cools it shrinks and grips the wheel tightly. Metal rims, gears and bearings are shrink-fitted onto shafts in the same way.

Dental fillings. Dentists choose filling materials that expand as nearly as possible like the tooth around them: a filling that grew and shrank much more than the tooth with every hot tea and ice cream would open a gap at its edge.

Liquids. A petrol tank or a bottle filled to the brim in the cool morning overflows or bursts in the hot afternoon, because the liquid expands more than its container. Bottles and tanks are always left with a little space at the top.

Gases. A tyre or a football left in the hot sun gets harder, and an over-pumped one can burst: the air inside expands and its pressure rises.

Thermometers work by expansion: the liquid in the bulb expands more than the glass and rises up the narrow tube.

> DID YOU KNOW?
> **The tower that grows in summer**
>
> The Eiffel Tower in Paris is about 300 m of wrought iron (the antennas on top take it to 330 m). Iron grows by about 12 millionths of its length for each degree, so between a freezing winter day and a hot summer one the tower's height changes by 15 cm or more. On a sunny day it also leans slightly away from the sun, as its sunny side expands more than its shady side: in the heatwave of 1976 its top moved 18 cm.

The bimetallic strip

Two strips of different metals, such as brass and iron, riveted or welded together face to face along their whole length, make a **bimetallic strip**. Brass expands about one and a half times as much as iron for the same warming. When the strip is heated, the brass gets longer than the iron, but the two are joined, so the strip **bends**, with the brass on the outside of the curve. When it cools, it straightens again. (Cooled below the temperature at which it is straight, it bends the other way.)

That movement can switch a circuit on and off, and bimetallic strips do just that in many things you use:

- **Thermostats** in electric irons, water heaters (geysers), rice cookers and room heaters: when the right temperature is reached, the strip bends away from its contact and switches the heater off; as it cools it touches again and switches it back on. Turning the iron's knob moves the contact, which sets the temperature.
- **Fire alarms**: heat from a fire bends the strip until it touches a contact and completes the alarm's circuit.

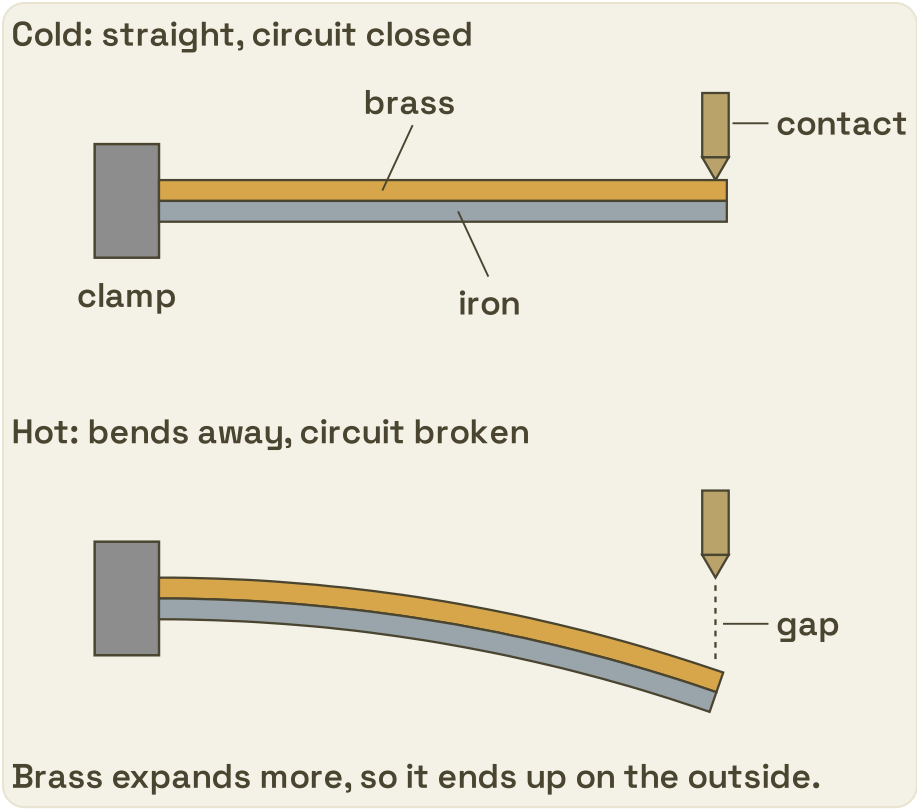


Figure 12.8 *A bimetallic strip used as a switch. Cold, it is straight and touches the contact, so current flows. Hot, the brass grows more than the iron, the strip bends away from the contact, and the circuit is broken. (The bend is drawn much bigger than it really is.)*

- **Circuit breakers** (MCBs): too large a current heats a bimetallic strip, which bends and trips the switch.
- **Flashing lights**: in older indicator lamps and Christmas-light flashers, a strip heated by the current bends, breaks the circuit, cools, and closes it again, over and over.
- **Dial thermometers**: a long bimetallic strip wound into a coil turns a pointer as the temperature changes (oven and meat thermometers).

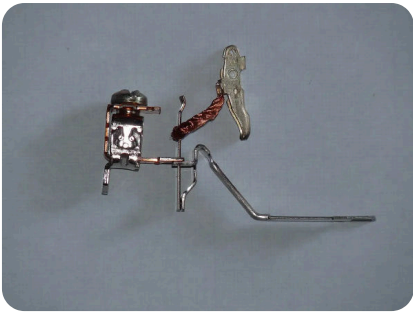


Figure 12.9 *The bimetallic strip from inside a circuit breaker. If too much current flows for too long, it heats up, bends, and trips the breaker, cutting off the current.*
Photo: Dmitry G, CC BY-SA 3.0

Check 6

A bimetallic strip is made of brass and iron, and brass expands more than iron. When the strip is heated, it:

a stays straight, getting longer

b bends, with the brass on the outside of the curve

c bends, with the iron on the outside of the curve

d gets shorter

12.3 Latent heat

Put a pan of crushed ice on a steady flame with a thermometer in it. The temperature rises to 0 °C and then **stops**, while the ice melts, even though the flame keeps giving heat. Only when the last of the ice has melted does the temperature rise again. At 100 °C it stops a second time, while the water boils away. The heat going in during those flat stretches doesn't show on the thermometer. The Scottish chemist Joseph Black, who studied it around 1761, called it **latent** (hidden) heat.

Latent heat is the heat absorbed or given out by a substance while it changes its state (solid to liquid, liquid to gas, or back), at a constant temperature.

The heat needed to melt a solid at its melting point is its **latent heat of fusion**; the heat needed to turn a liquid into vapor at its boiling point is its **latent heat of vaporization**. The same heat comes out again when the change goes the other way: water gives out its latent heat of fusion as it freezes, and steam gives out its latent heat of vaporization as it condenses.

Specific latent heat

The more ice there is, the more heat it takes to melt it: twice the mass, twice the heat. So the latent heat Q is proportional to the mass m that changes state:

$$Q \propto m \quad \text{so} \quad Q = mL \quad \text{or} \quad L = \frac{Q}{m}$$

The constant L is the **specific latent heat** of the substance. Put $m = 1$ kg and $L = Q$:

The specific latent heat of fusion of a substance is the heat needed to change 1 kg of it from solid to liquid at its melting point, without any change in temperature. For ice, it is **336,000 J/kg** (3.36×10^5 J/kg).

The specific latent heat of vaporization of a substance is the heat needed to change 1 kg of it from liquid to vapor at its boiling point, without any change in temperature. For water, it is **2,260,000 J/kg** (2.26×10^6 J/kg).

The SI unit of specific latent heat is the **joule per kilogram (J/kg)**. There is no °C in it, because the temperature doesn't change while the state does.

Substance	Melting point (°C)	Latent heat of fusion (J/kg)	Boiling point (°C)	Latent heat of vaporization (J/kg)
water	0	336,000	100	2,260,000
ethyl alcohol	−114	104,000	78	854,000
lead	327	23,000	1,749	866,000
aluminium	660	397,000	2,519	10,500,000
iron	1,538	247,000	2,861	6,090,000

So melting 1 kg of ice at 0 °C takes 336,000 J, as much as warming that kilogram of water from 0 °C to 80 °C. And boiling away 1 kg of water at 100 °C takes 2,260,000 J, more than five times what it took to warm it from 0 °C all the way to 100 °C.

CAREFUL

336,000 or 334,000?

The most careful modern measurement of the latent heat of fusion of ice is 333,550 J/kg, usually rounded to **334,000 J/kg**, and some books (including this site's Going Further books) use that. Nepal's textbooks and exams use **336,000 J/kg**, which comes from the older value of 80 calories per gram ($80 \times 4.2 = 336$). The difference is under 1%. In your exam, use 336,000 J/kg, or whatever value the question gives.

Latent heat explained by the kinetic molecular theory

The molecules of a substance have two kinds of energy. Because they move, they have **kinetic energy**; the temperature measures its average. Because they attract each other, they also have **potential energy**, which depends on how far apart they are, just as a stone lifted off the ground has potential energy because the Earth pulls it.

Below the melting point, heat makes the molecules of the solid vibrate faster: their kinetic energy rises, so the temperature rises.

At the melting point, the molecules are vibrating so hard that a little more energy lets them break out of their fixed places. Now all the heat goes into **pulling the molecules free of the forces holding them in place**: it raises their potential energy, not their kinetic energy. Their average speed doesn't change, so the temperature doesn't change, until every molecule is free and the solid has all melted. The molecules of water at 0 °C have the same kinetic energy as those of ice at 0 °C, but **more potential energy**.

At the boiling point, the same thing happens again, more strongly. Each molecule must be pulled **completely** away from its neighbors, to fly off on its own in the gas, and the vapor must also push back the air above the liquid to make room for itself. All the heat goes into that, and the temperature stays the same until the liquid has all boiled away.

Why the latent heat of vaporization is so much larger. When ice melts, each molecule only loosens its hold: in the water it still touches its neighbors and is still pulled by them. When water boils, each molecule must break away from all its neighbors entirely, and the vapor, about 1,600 times the volume of the water, must push the atmosphere back. So for water the latent heat of vaporization is about 6.7 times the latent heat of fusion. (Of the 2,260,000 J for each kilogram, about 170,000 J goes into pushing back the atmosphere: some books call it the **external** latent heat, and the rest, for pulling the molecules apart, the **internal** latent heat.)

When the change goes backward, the molecules come together again, their potential energy falls, and the same energy comes out as heat. That is why water freezing at 0 °C keeps giving out heat without its temperature falling, and why steam condensing on your skin gives out so much heat.

The ice is in a cylinder under a free piston, so it stays at the air pressure at sea level (water boils at 100 °C) and the steam can be heated further, and no heat is lost. The heater gives 1,000 J every second, so the length of a flat part, in seconds, measures the latent heat. Melting took 168 s: $Q = Pt = 1{,}000 \times 168 = 168{,}000$ J for 0.5 kg, so $L = Q \div m = 336{,}000$ J/kg. The sloping parts have different slopes too: the ice and the steam warm about twice as fast as the water, because their specific heat capacities (2,100 and about 2,040 J/(kg °C)) are about half of water's 4,200.

Check 7

While ice melts at 0 °C, the heat it takes in:

a raises the average kinetic energy of its molecules

b raises the potential energy of its molecules, freeing them from their fixed places

c is destroyed

d makes the molecules bigger

Check 8

Why is the latent heat of vaporization of water much larger than its latent heat of fusion?

a water boils at a higher temperature than ice melts

b steam is heavier than water

c in boiling, each molecule must break completely free of its neighbors and push back the air

d boiling happens more slowly than melting

Latent heat around us

Steam burns are worse than boiling-water burns. A gram of steam at 100 °C that condenses on your skin gives out 2,260 J as it turns to water, before the water even starts to cool. (Example 3 below works out how much worse.)

Ice cools better than ice-cold water. Every kilogram of ice at 0 °C takes in 336,000 J just to melt, before it starts to warm up. Water at 0 °C has no such store. That is why ice keeps fish fresh in a market and a cold drink cold.

Snow and glaciers melt slowly. On the first warm day of spring, the snow on the hills doesn't vanish in an afternoon, because every kilogram needs 336,000 J to melt.

Rivers fed by Himalayan snow rise steadily through spring and summer, not in a single flood.

Lakes freeze slowly. As the water of a lake like Gosaikunda freezes in winter, every kilogram gives out 336,000 J, which slows the cooling of the water below.

Protecting crops from frost. On frosty nights, fruit growers in some countries spray their trees with water. As the water freezes on the buds it gives out its latent heat, which holds the buds at about 0 °C instead of letting them fall to a few degrees below, the temperature that would kill them.

Refrigerators work by evaporation: a liquid evaporates in the pipes inside the freezer, taking its latent heat from the food; a pump then squeezes the vapor back into a liquid outside, in the coils at the back, where it gives that heat out to the room. That is why the back of a refrigerator is warm.

Sweat and the *matka* cool by the latent heat of vaporization, as you saw in 12.1.

WORKED EXAMPLE

Example 1: Melting ice

How much heat is needed to melt 2 kg of ice at 0 °C into water at 0 °C? (Latent heat of fusion of ice = 336,000 J/kg)

Answer

$m = 2$ kg, $L = 336{,}000$ J/kg

$$Q = mL = 2 \times 336{,}000 = 672{,}000 \text{ J}$$

672,000 J (6.72×10^5 J). The water is still at 0 °C: all the heat went into melting.

WORKED EXAMPLE

Example 2: From cold ice to warm water

How much heat is needed to change 200 g of ice at −10 °C into water at 40 °C? (Specific heat capacity of ice = 2,100 J/(kg °C), of water = 4,200 J/(kg °C); latent heat of fusion of ice = 336,000 J/kg)

Answer

Do it in three steps, one for each stretch of the heating curve. $m = 200\text{ g} = 0.2\text{ kg}$.

a. Warm the ice from −10 °C to 0 °C: $Q_1 = ms_{\text{ice}}(T_2 - T_1) = 0.2 \times 2{,}100 \times 10 = 4{,}200\text{ J}$
b. Melt the ice at 0 °C: $Q_2 = mL = 0.2 \times 336{,}000 = 67{,}200\text{ J}$
c. Warm the water from 0 °C to 40 °C: $Q_3 = ms_{\text{water}}(T_2 - T_1) = 0.2 \times 4{,}200 \times 40 = 33{,}600\text{ J}$

Total: $Q = Q_1 + Q_2 + Q_3 = 4{,}200 + 67{,}200 + 33{,}600 = 105{,}000\text{ J}$
105,000 J. Melting took nearly two-thirds of it.

WORKED EXAMPLE

Example 3: Steam against boiling water

Compare the heat given to the skin by 10 g of steam at 100 °C and by 10 g of water at 100 °C, each ending as water at the skin's temperature, 37 °C. (Latent heat of vaporization of water = 2,260,000 J/kg)

Answer

$m = 10\text{ g} = 0.01\text{ kg}$. The fall in temperature from 100 °C to 37 °C is 63 °C.
Boiling water only cools: $Q = ms(T_1 - T_2) = 0.01 \times 4{,}200 \times 63 = 2{,}646\text{ J}$
Steam first condenses at 100 °C, then the water cools:
$Q = mL + ms(T_1 - T_2) = 0.01 \times 2{,}260{,}000 + 2{,}646 = 22{,}600 + 2{,}646 = 25{,}246\text{ J}$
The steam gives 25,246 J, about 9.5 times as much as the boiling water's 2,646 J. That's why a burn from the steam of a pressure cooker's whistle is so much worse than a splash of boiling water.

WORKED EXAMPLE

Example 4: Finding a latent heat

A 2,000 W kettle, kept boiling with its lid open for 3 minutes, boils away 0.16 kg of water. Find the latent heat of vaporization of water from this experiment.

Answer

Heat given, $Q = Pt = 2{,}000 \times (3 \times 60) = 360{,}000$ J

$$L = \frac{Q}{m} = \frac{360{,}000}{0.16} = 2{,}250{,}000 \text{ J/kg}$$

2.25×10^6 J/kg, within half a percent of 2,260,000 J/kg. (A real kettle loses some heat to the air, so a real experiment usually gives a little **more** than the true value, as in numerical 5j; these readings are tidied for the example.)

YOUR TURN

Freezing the ice tray

An ice tray holds 300 g of water at 20 °C. How much heat must the freezer take out of it to turn it into ice at 0 °C?

Answer

Cool the water from 20 °C to 0 °C: $0.3 \times 4{,}200 \times 20 = 25{,}200$ J
Freeze it at 0 °C: $0.3 \times 336{,}000 = 100{,}800$ J
Total = **126,000 J**. Freezing takes four times as much as cooling the water all the way to 0 °C.

YOUR TURN

Sweat on a hot day

A porter carrying a load up a hill on a hot day evaporates 0.5 kg of sweat in an hour. Taking the latent heat of vaporization of water at skin temperature as 2,400,000 J/kg, how much heat does that take from the porter's body? How many joules is that each second?

Answer

$Q = mL = 0.5 \times 2{,}400{,}000 = 1{,}200{,}000$ J in an hour, which is $1{,}200{,}000 \div 3{,}600 \approx 333$ J **every second** (333 W), more than three times what a resting body gives out.

12.4 Heat equation

Mix hot water with cold, drop ice into juice, or blow steam into water, and heat flows from the hotter to the colder until both reach the same temperature. If no heat escapes to the surroundings, no energy is lost or made on the way, so

$$\text{heat lost by the hot substance} = \text{heat gained by the cold substance}$$

This is the **heat equation**, also called the **principle of calorimetry** or the principle of mixtures:

Heat equation: when a hot body and a cold body are in contact in a closed, insulated system, the heat lost by the hot body is equal to the heat gained by the cold body, until both reach the same temperature.

It is the law of conservation of energy, applied to heat. To use it, we write each side using the two formulas we have:

- for a change of temperature, $Q = ms(T_2 - T_1)$ (the heat for a body of mass m and specific heat capacity s);
- for a change of state, $Q = mL$.

Some books give the name "heat equation" to $Q = ms(T_2 - T_1)$ itself, and Grade 10 Science calls "heat lost = heat gained" the principle of calorimetry. In an exam, state the principle in words, write $Q = ms(T_2 - T_1)$ and $Q = mL$, and you have covered both.

The calorimeter

Heat measurements are made in a **calorimeter**: a small vessel of copper or aluminium (metals, so they quickly reach the same temperature as what is inside), with a lid, a stirrer and a thermometer, standing on a cork or felt pad inside a larger insulated container. The pad, the air gap and the lid stop heat leaking in or out, so "heat lost = heat gained" holds. The calorimeter itself is warmed or cooled with its contents, so in careful work it goes on the right side of the equation too, with its own mass and specific heat capacity (Example 8).

Solving problems with the heat equation

1. List each substance with its mass, specific heat capacity, starting temperature, and any latent heat.
2. Decide which substances lose heat (start hot) and which gain heat (start cold), and call the final temperature T.

3. Write the heat lost and the heat gained, using $ms\times$ (the fall or rise in temperature) and mL for any change of state. Each change of temperature is written as the larger temperature minus the smaller, so every term is positive.
4. Put heat lost = heat gained and solve for the unknown.
5. Check: the final temperature must lie between the starting temperatures.

WORKED EXAMPLE

Example 5: Ice in a glass of juice

50 g of ice at 0 °C is put into 250 g of juice at 30 °C. Find the final temperature, once all the ice has melted. (Treat the juice as water: $s = 4{,}200$ J/(kg °C); latent heat of fusion of ice = 336,000 J/kg; no heat is lost.)

Answer

Let the final temperature be T.

Heat lost by the juice, cooling from 30 °C to T: $0.25 \times 4{,}200 \times (30 - T) = 1{,}050(30 - T)$

Heat gained by the ice: it melts at 0 °C, $0.05 \times 336{,}000 = 16{,}800$ J, then the melted water warms from 0 °C to T: $0.05 \times 4{,}200 \times T = 210T$

Heat lost = heat gained:

$$1{,}050(30 - T) = 16{,}800 + 210T$$

$$31{,}500 - 1{,}050T = 16{,}800 + 210T$$

$$14{,}700 = 1{,}260T \quad \Rightarrow \quad T = 11.7\ ^\circ\text{C}$$

The juice ends at about 11.7 °C. Of the 31,500 J the juice could give by cooling to 0 °C, melting the ice took 16,800 J.

WORKED EXAMPLE

Example 6: Not all the ice melts

Dawa drops 100 g of ice at 0 °C into 100 g of water at 20 °C in a flask. What is the final temperature, and how much ice is left? (Latent heat of fusion of ice = 336,000 J/kg)

Answer

First check whether the water has enough heat to melt all the ice.
The most heat the water can give is by cooling all the way to 0 °C: $0.1 \times 4{,}200 \times 20 = 8{,}400$ J
To melt all the ice needs: $0.1 \times 336{,}000 = 33{,}600$ J
8,400 J is far less than 33,600 J, so **not all the ice melts**, and the mixture ends at **0 °C**, ice and water together.
Ice melted by 8,400 J: $m = \frac{Q}{L} = \frac{8{,}400}{336{,}000} = 0.025$ kg $= 25$ g
The final temperature is 0 °C, with 75 g of ice left (and 125 g of water).
If you had simply written heat lost = heat gained with a final temperature T, you would have got T below 0 °C, which is impossible for water that hasn't frozen: the check in step 5 catches it.

WORKED EXAMPLE

Example 7: Steam into water

10 g of steam at 100 °C is passed into 290 g of water at 20 °C, and all of it condenses. Find the final temperature. (Latent heat of vaporization of water = 2,260,000 J/kg; no heat lost.)

Answer

Heat lost by the steam: it condenses at 100 °C, $0.01 \times 2{,}260{,}000 = 22{,}600$ J, then the condensed water cools from 100 °C to T: $0.01 \times 4{,}200 \times (100 - T) = 42(100 - T)$
Heat gained by the water: $0.29 \times 4{,}200 \times (T - 20) = 1{,}218(T - 20)$

$$22{,}600 + 4{,}200 - 42T = 1{,}218T - 24{,}360$$

$$51{,}160 = 1{,}260T \quad \Rightarrow \quad T = 40.6\ ^\circ\text{C}$$

The water ends at about 40.6 °C. Just 10 g of steam warmed nearly 300 g of water by more than 20 °C, mostly with its latent heat.

WORKED EXAMPLE

Example 8: Finding the latent heat of fusion by the method of mixtures

A copper calorimeter of mass 100 g holds 300 g of water at 40 °C. Small pieces of ice at 0 °C, dried on blotting paper, are added and stirred in until all have melted. The temperature falls to 10 °C, and the calorimeter's contents now weigh 400 g. Find the latent heat of fusion of ice. (s of copper = 385 J/(kg °C), of water = 4,200 J/(kg °C))

Answer

Mass of ice added $= 400 - 300 = 100$ g $= 0.1$ kg.

Heat lost by the water and the calorimeter, cooling from 40 °C to 10 °C:
$0.3 \times 4{,}200 \times 30 + 0.1 \times 385 \times 30 = 37{,}800 + 1{,}155 = 38{,}955$ J

Heat gained by the ice: melting, $0.1 \times L$, then warming the melted water from 0 °C to 10 °C, $0.1 \times 4{,}200 \times 10 = 4{,}200$ J

Heat lost = heat gained:

$$38{,}955 = 0.1L + 4{,}200 \quad \Rightarrow \quad L = \frac{34{,}755}{0.1} = 347{,}550 \text{ J/kg}$$

About 3.48×10^5 J/kg. It is a little more than 336,000 J/kg: a real experiment has small errors (heat exchanged with the room, a reading taken a moment too soon), which is why careful experimenters repeat it.

YOUR TURN

How much ice for the tea?

Samikshya's 200 g of tea at 80 °C is too hot. How much ice at 0 °C must she add to bring it to 50 °C? (Treat tea as water; no heat lost.)

Hint

The tea cools by 30 °C. The ice melts, then its water warms from 0 °C to 50 °C. Call the ice's mass m.

Answer

Heat lost by the tea: $0.2 \times 4{,}200 \times 30 = 25{,}200$ J

Heat gained by m kg of ice: $m \times 336{,}000 + m \times 4{,}200 \times 50 = 546{,}000\,m$

$546{,}000\,m = 25{,}200$, so $m = 0.046$ kg = **about 46 g of ice**. (Cold water at 0 °C would need 120 g to do the same.)

Check 9

In a mixing problem, your answer for the final temperature of hot water mixed with cold water comes out higher than the hot water's starting temperature. What does it mean?

a the mixture got hotter, which can happen

b there is a mistake in the working, such as a sign or a wrong subtraction

c the cold water had a higher specific heat capacity

d heat was lost to the surroundings

IN SHORT

- **Evaporation** is the change of a liquid into vapor from its surface, at any temperature. The fastest molecules escape, so the liquid left behind **cools**.
- The rate of evaporation increases with the **temperature** of the liquid, a larger **surface area**, **wind**, **dry** air, a more **volatile** liquid and lower air **pressure**.
- **Boiling** happens all through the liquid, with bubbles, at a fixed boiling point; evaporation only at the surface, at any temperature.
- **Thermal expansion**: heated molecules vibrate harder and, because the force between them is lopsided, settle a little farther apart. Gases expand most, then liquids, then solids.
- **Linear** expansion is an increase in length, **superficial** in area (a hole grows too), **cubical** in volume (the only kind for liquids and gases).
- Expansion is allowed for in rails, bridges, roads and wires, and used in **bimetallic strips** (thermostats, fire alarms, circuit breakers), thermometers and shrink fits.
- **Latent heat** is the heat taken in or given out during a change of state at constant temperature. It changes the molecules' **potential energy**, not their kinetic energy.
- $Q = mL$. For water, the latent heat of **fusion** is 336,000 J/kg and of **vaporization** 2,260,000 J/kg.
- **Heat equation:** heat lost by the hot body = heat gained by the cold body, with $Q = ms(T_2 - T_1)$ for changes of temperature and $Q = mL$ for changes of state.

Exercise

1. Choose the correct option

a. Evaporation takes place:

- **a** only at 100 °C
- **b** all through the liquid
- **c** only at the surface, at any temperature
- **d** only when the liquid is heated

b. Which of these does **not** increase the rate of evaporation of water?

- **a** warming the water
- **b** a breeze over its surface
- **c** humid air above it
- **d** pouring it into a wide dish

c. The SI unit of specific latent heat is:

- **a** J/(kg °C)
- **b** J/kg
- **c** J °C
- **d** kg/J

d. While a pure solid is melting, its temperature:

- **a** rises steadily
- **b** falls
- **c** stays the same
- **d** rises and falls

e. When a bimetallic strip of brass and iron is heated, it bends because:

- **a** iron expands more than brass
- **b** brass expands more than iron
- **c** both metals contract
- **d** the rivets expand

f. The increase in the area of a metal sheet on heating is called:

- **a** linear expansion
- **b** superficial expansion
- **c** cubical expansion
- **d** anomalous expansion

g. How much heat is needed to boil away 0.5 kg of water already at 100 °C? (L = 2,260,000 J/kg)

- **a** 4,520,000 J
- **b** 1,130,000 J
- **c** 210,000 J
- **d** 226,000 J

h. The heat equation says that, in an insulated system:

- **a** heat gained is always more than heat lost
- **b** heat lost by the hot body equals heat gained by the cold body
- **c** temperature lost equals temperature gained
- **d** the hot body and the cold body change temperature by the same amount

2. Give reasons

a. A wet cloth on the forehead brings down a high fever.

b. Small gaps are left between the rails of a railway track.

c. Ice at 0 °C cools a drink better than the same mass of water at 0 °C.

d. Clothes dry faster on a windy day than on a still day.

e. A burn from steam at 100 °C is more severe than one from boiling water at 100 °C.

f. Electric wires are hung with some sag between the poles.

g. The temperature of water does not rise while it is boiling, though it is still being heated.

3. Differentiate between

i. Evaporation and boiling

ii. Linear, superficial and cubical expansion

iii. Latent heat of fusion and latent heat of vaporization

iv. Specific heat capacity and specific latent heat

4. Answer the following questions

a. What is evaporation? Explain, using the kinetic theory of matter, why evaporation causes cooling.

b. Describe the factors that affect the rate of evaporation, with an example of each.

c. Explain thermal expansion of a solid in terms of the motion of its molecules.

d. Write four impacts of thermal expansion in daily life, and how each is dealt with or used.

e. Define latent heat. Explain, using the kinetic molecular theory, why the temperature of melting ice stays at 0 °C while it is heated.

f. Why is the latent heat of vaporization of water greater than its latent heat of fusion?

g. State the heat equation. Write the formulas used with it for a change of temperature and a change of state.

h. Sketch the graph of temperature against time when ice at −20 °C is heated steadily until it becomes steam at 120 °C. Label each part and say what is happening in it.

i. How is a bimetallic strip used in the thermostat of an electric iron?

5. Numerical problems

a. How much heat is needed to melt 500 g of ice at 0 °C? ($L = 336{,}000$ J/kg)

b. How much heat is needed to change 2 kg of water at 100 °C into steam at 100 °C? (L = 2,260,000 J/kg)

c. Find the heat needed to change 0.5 kg of ice at −5 °C into water at 0 °C. (s of ice = 2,100 J/(kg °C); L = 336,000 J/kg)

d. How much heat does it take to turn 100 g of water at 20 °C into steam at 100 °C? (s of water = 4,200 J/(kg °C); L = 2,260,000 J/kg)

e. Find the heat needed to change 1 kg of ice at −10 °C into steam at 100 °C. What share of it is used in boiling?

f. A 500 W heater is put in a large block of ice at 0 °C. How long does it take to melt 1 kg of the ice, if all its heat goes into the ice?

g. Rizwan adds 40 g of ice at 0 °C to 200 g of water at 25 °C. Find the final temperature, assuming all the ice melts and no heat is lost.

h. How much steam at 100 °C must be passed into 500 g of water at 20 °C to warm it to 80 °C?

i. A copper calorimeter of mass 200 g contains 300 g of water at 30 °C. 50 g of ice at 0 °C is added. Find the final temperature once all the ice has melted. (s of copper = 385 J/(kg °C))

j. In an experiment, a 1,500 W kettle kept boiling for 10 minutes boils away 0.39 kg of water. Find the latent heat of vaporization from these readings. Why is your answer larger than 2,260,000 J/kg?

k. **Think:** 50 g of ice at 0 °C cools 250 g of juice at 30 °C to about 11.7 °C (Example 5). Find the final temperature if 50 g of water at 0 °C is used instead, and explain the difference.

Found a mistake? Please tell us.