Surfing

# VCE CHEMISTRY 2

## Unit 2 What Makes Water Such a Unique Chemical?

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# Contents

# Introduction

This book covers the Chemistry content specified in the Victorian Certificate of Education Chemistry Study Design. Sample data has been included for suggested experiments to give you practice to reinforce practical work in class.

Each book in the *Surfing* series contains a summary, with occasional more detailed sections, of all the mandatory parts of the syllabus, along with questions and answers.

All types of questions – multiple choice, short response, structured response and free response – are provided. Questions are written in exam style so that you will become familiar with the concepts of the topic and answering questions in the required way.

Answers to all questions are included.

A topic test at the end of the book contains an extensive set of summary questions. These cover every aspect of the topic, and are useful for revision and exam practice.

# Words To Watch

**account, account for** State reasons for, report on, give an account of, narrate a series of events or transactions.

**analyse** Interpret data to reach conclusions.

**annotate** Add brief notes to a diagram or graph.

**apply** Put to use in a particular situation.

**assess** Make a judgement about the value of something.

**calculate** Find a numerical answer.

**clarify** Make clear or plain.

**classify** Arrange into classes, groups or categories.

**comment** Give a judgement based on a given statement or result of a calculation.

**compare** Estimate, measure or note how things are similar or different.

**construct** Represent or develop in graphical form.

**contrast** Show how things are different or opposite.

**create** Originate or bring into existence.

**deduce** Reach a conclusion from given information.

**define** Give the precise meaning of a word, phrase or physical quantity.

**demonstrate** Show by example.

**derive** Manipulate a mathematical relationship(s) to give a new equation or relationship.

**describe** Give a detailed account.

**design** Produce a plan, simulation or model.

**determine** Find the only possible answer.

**discuss** Talk or write about a topic, taking into account different issues or ideas.

**distinguish** Give differences between two or more different items.

**draw** Represent by means of pencil lines.

**estimate** Find an approximate value for an unknown quantity.

**evaluate** Assess the implications and limitations.

**examine** Inquire into.

**explain** Make something clear or easy to understand.

**extract** Choose relevant and/or appropriate details.

**extrapolate** Infer from what is known.

**hypothesise** Suggest an explanation for a group of facts or phenomena.

**identify** Recognise and name.

**interpret** Draw meaning from.

**investigate** Plan, inquire into and draw conclusions about.

**justify** Support an argument or conclusion.

**label** Add labels to a diagram.

**list** Give a sequence of names or other brief answers.

**measure** Find a value for a quantity.

**outline** Give a brief account or summary.

**plan** Use strategies to develop a series of steps or processes.

**predict** Give an expected result.

**propose** Put forward a plan or suggestion for consideration or action.

**recall** Present remembered ideas, facts or experiences.

**relate** Tell or report about happenings, events or circumstances.

**represent** Use words, images or symbols to convey meaning.

**select** Choose in preference to another or others.

**sequence** Arrange in order.

**show** Give the steps in a calculation or derivation.

**sketch** Make a quick, rough drawing of something.

**solve** Work out the answer to a problem.

**state** Give a specific name, value or other brief answer.

**suggest** Put forward an idea for consideration.

**summarise** Give a brief statement of the main points.

**synthesise** Combine various elements to make a whole.

SURFING

# VCE CHEMISTRY 2

## Area of Study 1

# How Do Substances Interact with Water?

# 1 Hydrides and Their Boiling Points

The element **hydrogen** has the smallest atoms of all elements. Hydrogen can form compounds with all other elements except group 8/18 of the periodic table. (Note that some periodic tables refer to the noble or inert gases as group 18 whereas others refer to it as group 8, and sometimes as group 0.)

A **hydride** is a compound formed from hydrogen and another element. Hydrides can be ionic or covalent.

**Ionic hydrides** consist of hydrogen combined with a metal. Examples include sodium hydride (NaH) and magnesium hydride ($MgH_2$). Here hydrogen is the more electronegative element in each compound, so it is considered as **negatively charged** ($H^-$).

**Covalent hydrides** consist of hydrogen bonded to one or more non-metals by sharing electrons, for example, hydrogen chloride (HCl), ammonia ($NH_3$), water ($H_2O$), silicon tetrahydride ($SiH_4$), phosphorus trihydride ($PH_3$), dihydrogen sulfide ($H_2S$). When these compounds dissolve in water they dissociate and a **positively charged hydrogen ion forms** ($H^+$). This is really just a positively charged proton; it cannot exist alone, so it bonds to a water molecule, forming an hydronium ion ($H_3O^+$).

## Hydrides of group 14/4

If we look at the boiling points of the hydrides of group 14/4 of the periodic table, we can see a trend.

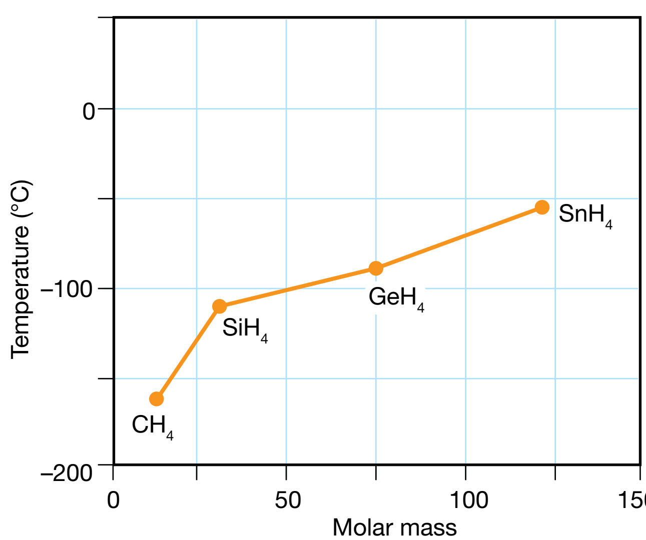



**Figure 1.1** Boiling points of the hydrides of group 14/4.

The **boiling points of the group 14/4 hydrides increase** as you go **down the group** – from methane, to silicon hydride, germanium hydride and tin hydride.

Remember, melting and boiling points of covalent substances are determined by the strength of forces or bonds between their particles – intermolecular forces. Melting and boiling points of ionic compounds are determined by the strength of their ionic bonds.

Across the graph in Figure 1.1, the elements bonded with hydrogen become **more metallic** so their bonds become more ionic. Carbon is a non-metal. It combines with hydrogen to form methane ($CH_4$), a small, non-polar, covalent molecule. Tin is a metal, it combines with hydrogen to form an ionic compound, $SnH_4$. As you already know, ionic compounds tend to have higher boiling and melting points than covalent compounds because ionic bonds are stronger than the intermolecular forces between covalent molecules.

## Comparing hydrides of groups 14/4 to 17/7

Now, let's compare the **boiling points of the hydride**s of groups 14/4 with those of groups15/5, 16/6, and 17/7 of the periodic table.

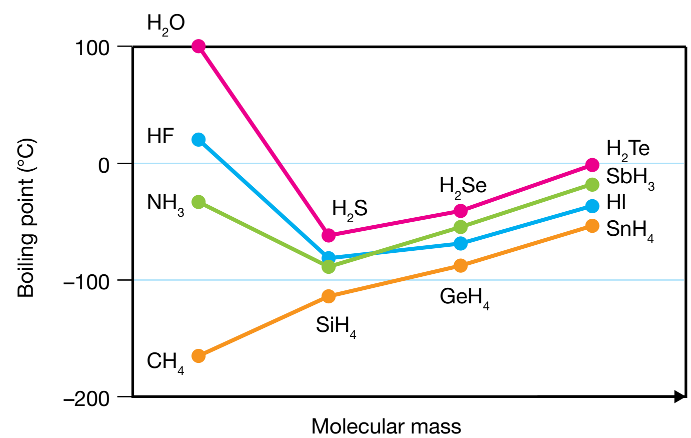



**Figure 1.2** Boiling points of hydrides of groups 14 to 17 (groups 4 to 7).

In Figure 1.2, we can see that the overall trend is still there. The bottom line, shows the graph we have already looked at – the boiling points of hydrides of group 14/4 elements – with its gradual increase in boiling point down the group from carbon hydride to tin hydride.

The other lines on the graph show the same overall trend of increasing boiling point down the group – *except* for the first member of each group (**ammonia**, **hydrogen fluoride** and **water**). These three compounds have **much higher boiling points than would be expected** when we look at the trends in boiling points of all the other hydrides.

**Water**, **hydrogen fluoride** and **ammonia** are all small, covalent molecules. The forces that determine their melting and boiling points are their **intermolecular forces** – the forces between their molecules – and these are dispersion forces, dipole-dipole forces and hydrogen bonds.

Hydrogen bonds are the strongest intermolecular forces. It is the presence or absence of strong hydrogen bonds that makes the difference. You will recall that a hydrogen bond forms between a **hydrogen** atom in a molecule and a very electronegative atom (nitrogen, fluorine or oxygen) in a nearby molecule of that same substance. So hydrogen bonds only form between molecules which contain a **hydrogen** atom and also an atom of one of the very electronegative atoms, **nitrogen**, **fluorine or oxygen**.

The only compounds in those four graphs that could form **intermolecular hydrogen bonding** are ammonia, hydrogen fluoride and water. So the reason these three compounds have higher melting and boiling points than expected is because their molecules are attracted together not just by weak dispersion and dipole-dipole forces, but also by strong hydrogen bonds.

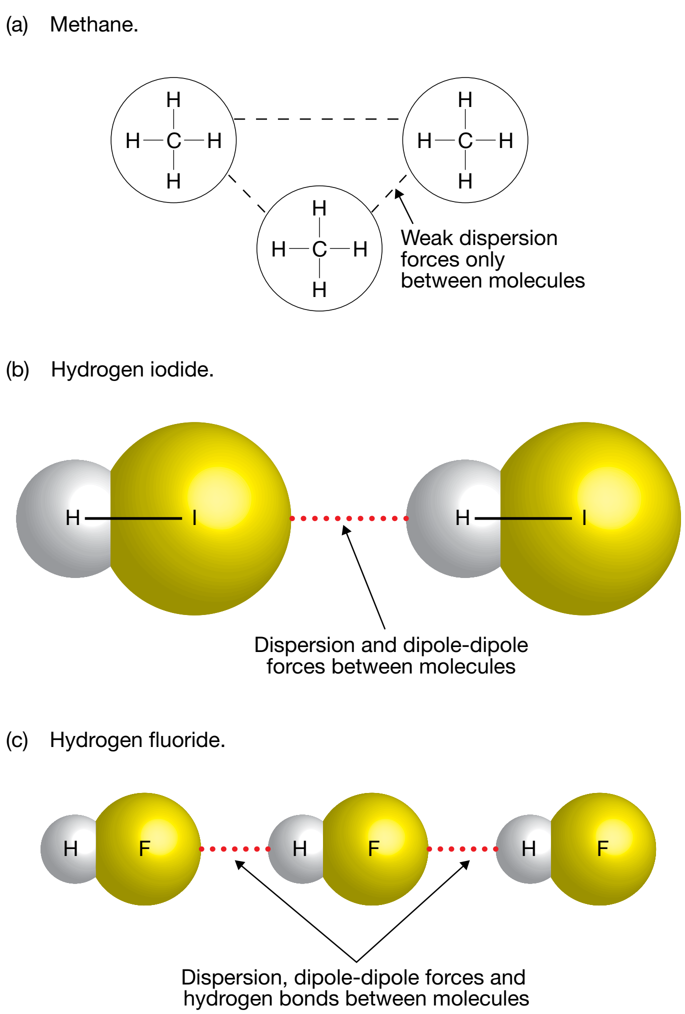



**Figure 1.3** Models of intermolecular bonds between some hydride molecules.

Stronger bonds cause stronger attractive forces between molecules, which means more energy and thus a higher temperature is needed to break these intermolecular attractions to cause a change in state.

## QUESTIONS

1. Hydrogen is the element with the smallest atoms.
   (a) What is a hydride?
   (b) Name and state the formulas for two covalent hydrides.
   (c) Name and state the formulas for two ionic hydrides.
2. Consider the graph shown in Figure 1.1.
   (a) Outline any trend in boiling point of hydrides down group 14/4 of the periodic table.
   (b) Suggest two causes for this trend in boiling point.
3. Consider Figure 1.2 showing trends in boiling points.
   (a) Name three hydrides that do not follow the trend in boiling point of most hydrides.
   (b) Are their boiling points higher or lower than expected, based on the boiling points of other hydrides of similar size?
   (c) How does the intermolecular bonding of these three compounds differ from the bonding between molecules of other hydrides.
   (d) Identify the points, on the blue and green graphs, which are not labelled.
4. Check your knowledge with this quick quiz.
   (a) Melting and boiling points of covalent compounds are determined by the strength of (ionic bonds/intermolecular forces/covalent bonds).
   (b) Identify three types of intermolecular forces and list them in order of increasing strength.
   (c) Which type of intermolecular force is the strongest?
   (d) Which is stronger, an ionic bond or a hydrogen bond?
   (e) Name three elements with strongly electronegative atoms.
   (f) Which of the following hydrides would have intermolecular hydrogen bonds?
   $H_2S$, $H_2O$, HF, $CH_4$, $NH_3$, HI, $H_2Te$, $SnH_4$

# 2 Melting and Boiling Points of Group 16 Hydrides

**Group 16 elements** (also called group 6) become increasingly metallic down the group, from oxygen and sulfur which are non-metals to polonium which is a metal.

All of the elements of group 16/6 react with hydrogen to form hydrides as shown in Table 2.1.

**Table 2.1** Group 16/6 elements and their hydrides.

| Group 16/6 elements | Hydrides of group 16/6 |
|---|---|
| Oxygen O | Water $H_2O$<br>Hydrogen peroxide $H_2O_2$ |
| Sulfur S | Hydrogen sulfide $H_2S$<br>Hydrogen disulfide $H_2S_2$ |
| Selenium Se | Hydrogen selenide $H_2S$ |
| Tellurium Te | Hydrogen telluride $H_2Te$ |
| Polonium Po | Hydrogen polende $H_2Po$ |

Except for water and hydrogen peroxide, these hydrogen compounds are all volatile, toxic gases with very unpleasant odours. They all form weak acids in water, the strength of the acid increasing down the group.

Melting and boiling points of group 16/6 hydrides show a trend down the group as you can see in Table 2.2. When working out the trend, don't be fooled by the negative numbers, for example, −83°C is a lower temperature than −66°C.

**Table 2.2** The melting and boiling points of group 16/6 hydrides.

| Hydride | Molar mass | Melting point (°C) | Boiling point (°C) |
|---|---|---|---|
| $H_2O$ | 18 | 0 | 100 |
| $H_2S$ | 34 | −83 | −62 |
| $H_2Se$ | 81 | −66 | −40 |
| $H_2Te$ | 130 | −49 | −2 |

You can see from these figures that both melting and boiling points increase down group 16/6, from $H_2S$ to $H_2Te$. However, water does not fit this trend. Instead it melts and boils at surprisingly high temperatures.

This occurs because of the **intermolecular forces between water molecules**. Water molecules are attracted together by dispersion forces, dipole-dipole forces and also by strong hydrogen bonds. For ice to melt or liquid water to boil, all these intermolecular forces and bonds need to be broken. Water can form up to four hydrogen bonds per molecule as it has two hydrogen atoms and two lone pairs on the oxygen atom. This accounts for its relatively high melting and boiling points.

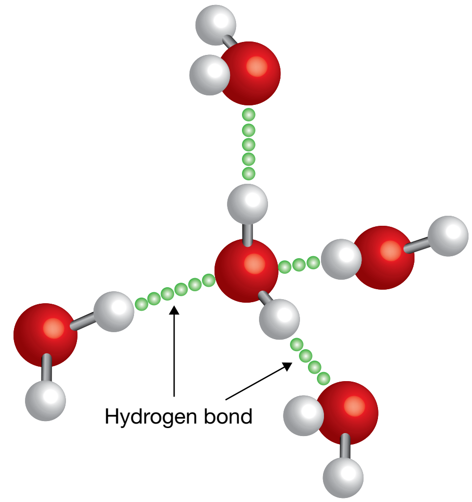



**Figure 2.1** A water molecule forming hydrogen bonds with four other water molecules.

## QUESTIONS

1. Research information about group 16/6 elements to outline:
   (a) Their discovery.
   (b) Their main uses.
2. (a) Graph the information in Table 2.2 to compare melting and boiling points of group 16/6 hydrides.
   (b) Based on trends in your graph, predict what would be the melting and boiling points of water if it followed these trends. Also comment on the effect this could have on the state of water found on Earth.
3. The boiling points of four covalent compounds with similar sized molecules are shown in the table.

| Feature | $CH_4$ | $NH_3$ | $H_2O$ | HF |
|---|---|---|---|---|
| Molar mass (g) | 16 | 17 | 18 | 20 |
| Boiling point (°C) | −160 | −33 | 100 | 19 |

   Account for the differences in boiling point of these substances.
4. Use formulas connected by dotted lines to illustrate hydrogen bonding between molecules of:
   (a) Hydrogen fluoride.
   (b) Ammonia.
   (c) Water.
5. Check your knowledge with this quick quiz.
   (a) Compared to compounds with molecules of similar size, the melting and boiling points of water are (higher/lower) than expected.
   (b) Which is the lower temperature −50 or −75?
   (c) Name three types of intermolecular forces between water molecules.

# 3 Specific Heat Capacity of Water

## Temperature versus heat

Temperature and heat are not the same thing. If two different sized copper cubes are placed together into hot water, they will both reach the same temperature, but they will not hold the same amount of heat. The larger one will hold more heat energy.

We can see this if we then place each cube into a separate beaker of cold water, both at the same temperature and measure any rise in temperature of the water. The larger cube will cause a larger rise in temperature of the cold water as it contains more heat energy.

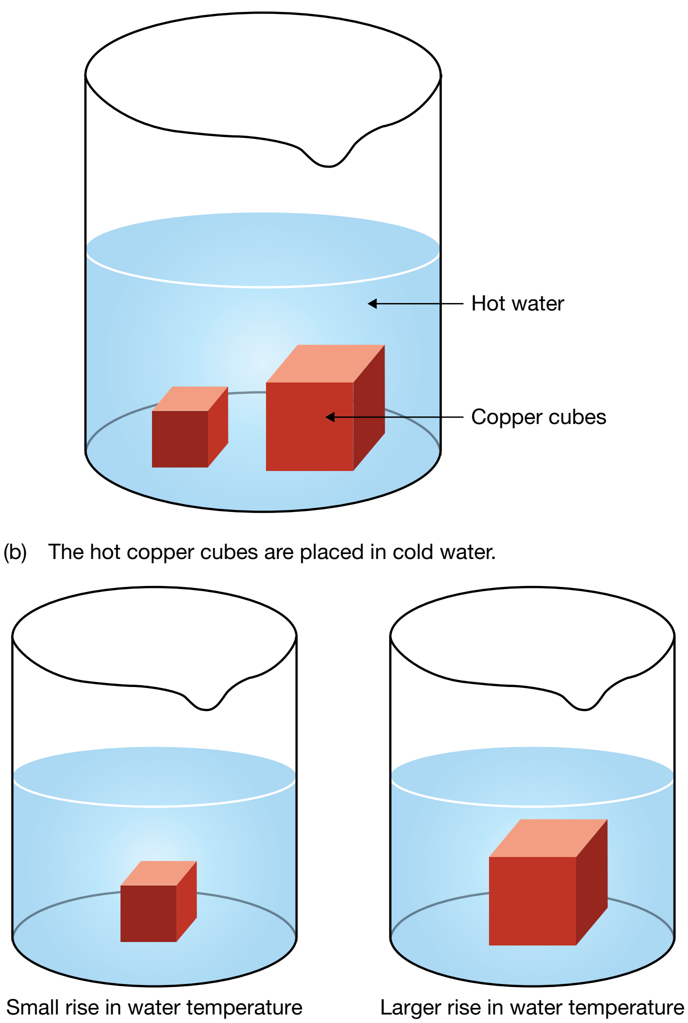



**Figure 3.1** Transferring heat.

**Temperature** is a measure of how hot an object is – the kinetic energy of its particles – and it is measured in degrees Celsius (°C) or kelvins (K). **Heat** is a measure of the thermal energy contained in an object and is measured in joules.

When two objects at different temperatures are in contact, heat flows from the hotter object to the cooler object as the heat energy causes the particles of the hotter object to move faster. Temperature is a measure of the kinetic energy of its particles, so making its particles move faster means that its temperature increases. The amount that temperature increases depends on:

- The amount of heat energy transferred.
- The mass of the object.
- The substance the object is made of.

Some materials absorb more heat more readily than others. If we place two identical masses of copper and aluminium into the same beaker of hot water, the copper will heat up faster than the aluminium. The aluminium will need to absorb more heat energy than the copper in order to reach the same temperature as the copper.

## Specific heat

Water is unusual in that it heats up even more slowly than metals such as copper and aluminium. In fact water can absorb a lot of heat energy without showing much of a change in temperature at all. We say that **water has a high specific heat capacity**. This property of water makes water very useful for storing heat energy and for transporting it – for example in central heating pipes, and via the blood in our bodies.

The specific heat of a substance can be **measured** by placing a measured quantity in an insulated container, applying a measured amount of heat and recording the temperature rise. An insulated container used in this way is called a calorimeter.

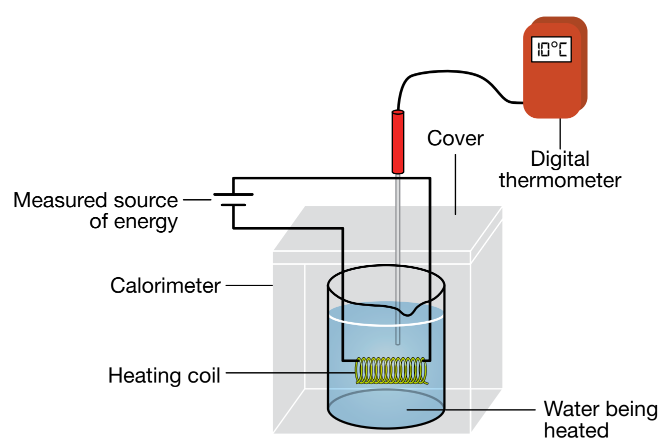



**Figure 3.2** Measuring specific heat.

**Specific heat capacity** is defined as the amount of heat energy needed to raise the temperature of 1 gram of a substance by one degree Celsius or one kelvin. Table 3.1 compares the specific heat capacity of some substances. You can see that the value for water is relatively high.

**Table 3.1** The specific heat capacity of some substances.

| Substance | Specific heat capacity (J $g^{-1}$ $K^{-1}$) |
|---|---|
| Water | 4.18 |
| Ethanol | 2.4 |
| Dry soil | 2.2 |
| Air | 1.0 |
| Aluminium | 0.9 |
| Glass | 0.67 |
| Copper | 0.4 |
| Mercury | 0.1 |

This means that when the same amount of heat is absorbed by water and by the same amount of most other substances, such as air, soil, or metals, there will be a smaller temperature increase in the water. Or, in other words, water needs more heat than other substances to raise its temperature by the same amount. We say that water has a high specific heat capacity, and this is related to its hydrogen bonding.

On a hot day, the sand at the beach or a lake will get very hot, while the temperature of the lake or ocean hardly rises at all. This happens despite the fact that the land and the water have each received the same heat energy per square metre from the Sun.

**Figure 3.3** Land gets hotter than water.

## Importance of water's high specific heat

The **high heat capacity of water** (specific heat) **is important for life on Earth**. A water environment stays at a **more constant temperature** than a land environment where there is more fluctuation of conditions. This means that the survival of organisms is easier in water than on land. The regulation of body temperature is easier for aquatic organisms, because in water they are not subjected to temperature extremes and drying winds. This may be why early organisms on Earth were able to evolve in the oceans.

The high heat capacity of water also plays an important role in the **regulation of the climate** on Earth. Coastal regions typically have cooler climates in summer than inland areas because of the high heat absorption of the ocean. Water is also slow to release its absorbed heat, so the oceans help to stabilise temperatures on Earth.

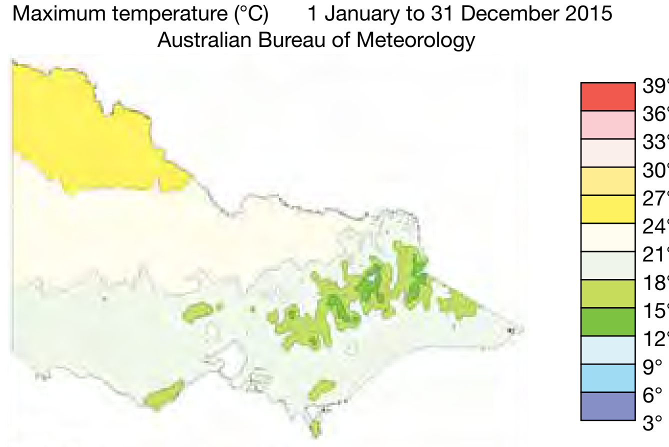



**Figure 3.4** Mean daily maximum temperatures (°C) in Victoria in 2015.

You can see from Figure 3.4 that the coolest areas in Victoria are along the coast and the warmest areas are inland. Coastal regions have a smaller range of temperature than the interior of continents because of the moderating effect of the ocean. The further an area is from the ocean, the more extreme its temperatures are likely to be.

Cool onshore breezes occur along the coast because the land heats up more than the water during the day, causing convection currents which bring cooling breezes.

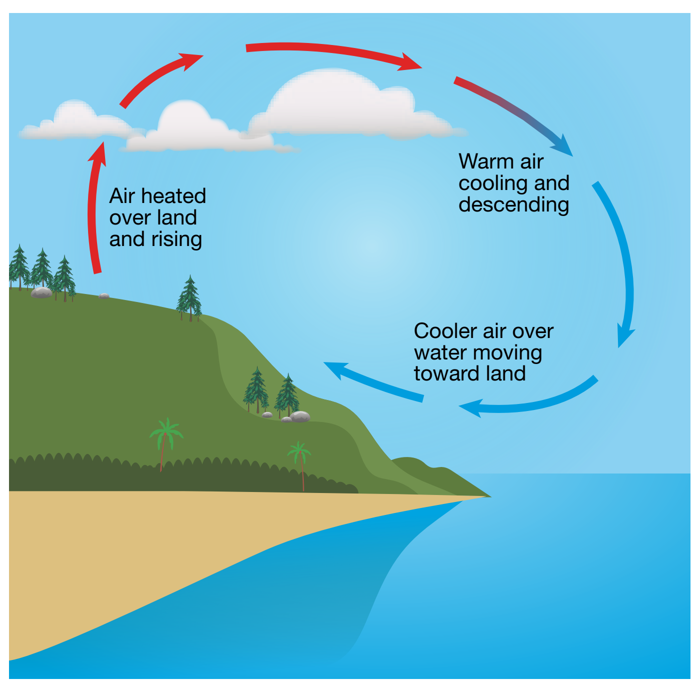



**Figure 3.5** Cool sea breezes on summer afternoons.

## Temperature scales

There are three scales used for measuring temperature. The Celsius scale is the one used in Australia today to measure temperature. In some parts of the world, such as the United States, the Fahrenheit scale is still used. The Kelvin scale was designed so that zero K is absolute zero, which is approximately −273°C. This is the coldest temperature thought to be possible, the temperature at which particles would theoretically stop moving.

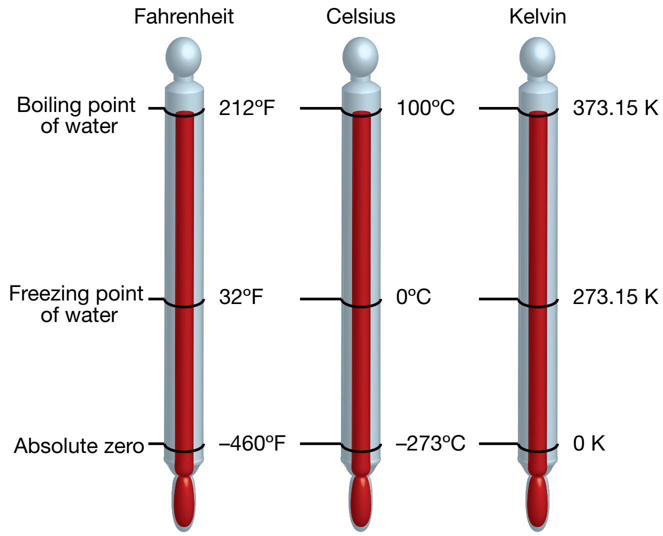



**Figure 3.6** Three temperature scales.

Notice that, on the scale above, 1 Celsius degree is equal to 1 kelvin.

To convert between degrees Celsius and kelvins, use the following relationships:

K = °C + 273 or °C = K − 273

## Calculations

1. Specific heat is usually expressed as joules per gram per kelvin. However, it can also be expressed in terms of **moles**.

   To raise the temperature of 1 gram of water through 1 kelvin, 4.18 joules is needed.

   1 mole of water has a mass of 18.0 grams.

   4.18 × 18 joules = 75.24 joules.

   So 75.24 joules of energy is needed to raise the mass of 1 mole of water by 1 kelvin.

2. The heat needed to raise the temperature of a body can be calculated as follows:

$$\text{Specific heat capacity} = -\frac{\text{change in heat energy}}{\text{mass} \times \text{change in temperature}}$$

   We can rearrange this formula as follows.

$$\begin{array}{ccccccc} \text{Change in energy} & = & -\text{mass} & \times & \text{specific heat} & \times & \text{temperature change} \\ \Delta H & = & -m & \times & C & \times & \Delta T \end{array}$$

   Where: $\Delta H$ is the change in heat energy in joules
   $m$ is the mass in grams
   $C$ is the specific heat capacity in J $g^{-1}$ $K^{-1}$
   $\Delta T$ is change in temperature.
   Note that $\Delta T$ can be in either °C or K.

The use of the negative sign in this equation is a convention that lets us determine whether energy is being absorbed or released in a reaction.

You will see that when the temperature increases, the reaction releases heat and the value of $\Delta H$ is negative. When the temperature drops, heat is absorbed and the value of $\Delta H$ is positive.

Later on you will see that $\Delta H$ is negative for exothermic reactions and positive for endothermic reactions.

## Sample questions

1. Calculate the heat needed to raise the temperature of 50 mL of water from 25°C to 40°C.

   *Solution:*

   50 mL of water has a mass of 50 grams.
   Specific heat ($C$) of water = 4.18 J $g^{-1}$ $K^{-1}$
   Heat = − mass × specific heat × temperature rise
   $\Delta H = -m \times C \times \Delta T$
   $\Delta H = -50 \times 4.18 \times (40 - 25)$
   $= -3135$ joules or 3.135 kJ

   This tells us that 3.1 kJ of energy has been absorbed by the water.

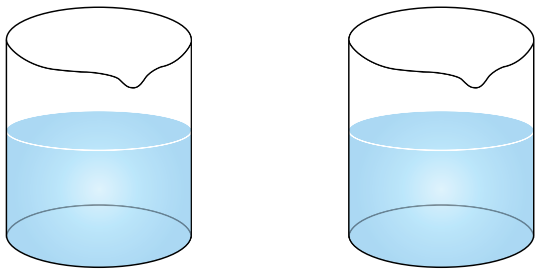
2. Calculate the heat needed to raise the temperature of a 50 g block of aluminium from 25°C to 40°C.

   *Solution:*

   Specific heat ($C$) of aluminium = 0.9 J $g^{-1}$ $K^{-1}$
   Heat = − mass × specific heat × temperature rise
   $\Delta H = -m \times C \times \Delta T$
   $\Delta H = -50 \times 0.9 \times (40 - 25)$
   $= -675$ joules or 0.675 kJ

   0.68 kJ of energy has been absorbed.

3. The temperature of a 40 gram piece of glass (specific heat = 0.67 J $g^{-1}$ $K^{-1}$) is changed by transferring it between two containers of water maintained at different temperatures.




The piece of glass is heated to a temperature of 25°C by allowing it to sit in the water of container A for 5 minutes.

It is then transferred from container A to B and allowed to remain in the water of container B until its temperature rises to 45°C.

(a) How much heat must be added to this piece of glass for its temperature to rise from 25°C to 45°C?

(b) The hot glass is then placed into a new container with 100 mL of water at a temperature of 20°C. The water in this container is warmed by the hot piece of glass. Calculate the new temperature of the water.

*Solution:*

(a) Heat = mass × specific heat × temperature rise
$\Delta H = -m \times C \times \Delta T$
$\Delta H = -40 \times 0.67 \times (45 - 25)$
$= -5360$ joules or 5.36 kJ

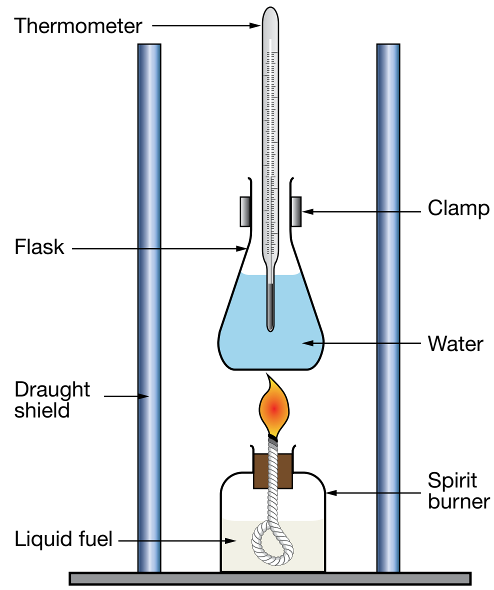
5.4 kJ of energy has been added to the glass.

(b) Specific heat ($C$) of water = 4.18 J $g^{-1}$ $K^{-1}$
Heat = −mass × specific heat × temperature rise
$\Delta H = -m \times C \times \Delta T$
$-5360 = -100 \times 4.18 \times \Delta T$
$\Delta T = 12.823$°C

The new temperature of the water is 33°C.

## QUESTIONS

1. (a) Define specific heat capacity.
   (b) Place the following substances in order of increasing heat capacity – ethanol, copper, water, mercury.
   (c) Given the same exposure to the Sun's heat, which would show a greater temperature rise after one hour, an area of water or the same sized area of land?
   (d) Justify your answer to part (c).
   (e) Explain how this is an advantage for organisms living in aquatic environments.
2. Describe an experiment you carried out to study specific heat.
3. Research applications of specific heat.
4. Calculate the heat needed to raise the temperature of:
   (a) 5 litres of water from 25°C to 35°C.
   (b) 400 mL of water from 20°C to 40°C
5. Use the specific heats in Table 3.1 to calculate the following.
   (a) Calculate the heat needed to raise the temperature of 5 kg of copper from 25°C to 40°C.
   (b) Calculate the heat used if 5 grams of aluminium is to undergo a temperature rise of 8 K.
6. A conical flask containing 100 mL (100 g) of water is suspended over a spirit burner containing ethanol as shown in the diagram.




   (a) Calculate the amount of heat absorbed by the water in the flask if its temperature increases by 16°C.
   (b) The amount of ethanol burned to bring about this temperature rise was 0.49 g. Assuming all of the heat produced enters the water, how much heat energy would be released by one mole of ethanol?
7. (a) Calculate to compare the rise in temperature when 100 g of water and 100 g of dry soil each absorb 500 joules of heat energy. See Table 3.1 for specific heats.
   (b) Calculate to compare the heat released when 100 g of water and 100 g of soil each show a drop in temperature of 3.0 K.
8. Check your knowledge with this quick quiz.
   (a) The specific heat of water is (higher/lower) than most other substances.
   (b) This means that water is (slower/faster) to heat and then, once it is hot, it is (slower/faster) to cool down.
   (c) It takes (more/less) heat to raise the temperature of water the same amount as other liquids.

# 4 Latent Heat of Water

**Latent heat** refers to the energy absorbed or released by a substance when it changes state.

Latent heat is expressed as the amount of heat in joules per mole or joules per gram of the substance undergoing change of state.

You will recall that when a substance:

- Melts or evaporates it **absorbs heat energy** as **bonds are being broken**.
- Freezes or condenses it **releases heat energy** as **new bonds are being formed**.

The **latent heat of fusion** is the heat absorbed when a substance melts or the heat released when a substance freezes (solidifies).

Solid + heat of fusion $\rightleftharpoons$ liquid

The **latent heat of vaporisation** is the heat absorbed when a substance evaporates or the heat released when a substance condenses.

Liquid + heat of vaporisation $\rightleftharpoons$ gas

Note that during a change of state, the temperature does not change (see Figure 4.1), so latent heat is sometimes called **hidden heat** – it goes into bond changes (which we cannot see) and does not cause a change in temperature (which we could see). The heat is being absorbed to break bonds (melting and evaporating) or being released when new bonds form (condensing and freezing).

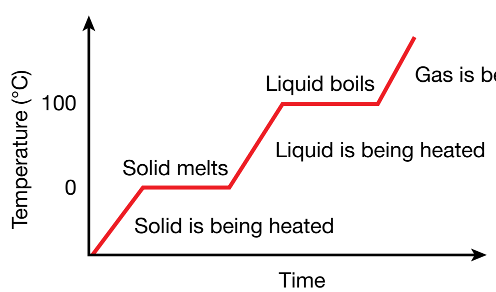



**Figure 4.1** Boiling water.

If you very gently heat an ice cube at 0°C, it will change state to liquid water (it melts). This happens without any change in its temperature and it will form liquid water at 0°C. The heat has been used to break bonds (mainly hydrogen bonds) between water molecules, which were held together in the ice, so they can then move more freely.

If you then further heat the water, you see its temperature rise – until it boils. But then if you keep on heating the boiling water there is no further temperature rise. The heat energy is breaking more bonds between water molecules so they can leave the surface of the water. The water is changing from a liquid to a gas.

**Table 4.1** Latent heat of some substances.

| Substance | Latent heat of fusion ($kJ\ mol^{-1}$) | Latent heat of vaporisation ($kJ\ mol^{-1}$) |
|---|---|---|
| Water | 6.01 | 40.66 |
| Ammonia | 5.65 | 23.35 |
| Methane | 0.94 | 8.18 |
| Methanol | 3.16 | 35.27 |

Notice that the latent heat of vaporisation is higher than the latent heat of fusion and also that water has a relatively high latent heat – once again due to its hydrogen bonding between molecules.

## Significance of water properties

The high specific heat and the high latent heat of vaporisation of water are important to the roles played by water on Earth.

- They allow **plants and animals to be cooled efficiently** by the evaporation of water. A large amount of heat is needed and this comes from the plant or animal, leaving their tissues cooler.
- Organisms are composed largely of water. The high specific heat and high latent heat of water help to **maintain organisms at a relatively constant temperature**, preventing sudden changes in their temperature.
- The high specific heat and latent heat of water in the oceans plays an important role in moderating Earth's **climate**. Heat, absorbed during evaporation, is transported in water vapour from the oceans to high altitudes and latitudes where it is released as the vapour condenses.

## QUESTIONS

1. (a) Define latent heat.
   (b) Identify the unit for latent heat.
2. Distinguish between latent heat of fusion and vaporisation.
3. Describe any experiments you performed on latent heat.
4. Explain each of the following.
   (a) A burn caused by 1 gram of steam will do more damage than one caused by 1 gram of boiling water, when both are at a temperature of 100°C.
   (b) One way in which the heat of vaporisation can influence weather systems.
   (c) The high heat of vaporisation means that water evaporating from our skin is effective at cooling us.
   (d) Picnic coolers (eskys) make use of the latent heat of fusion.
5. Research one industrial application of heat of vaporisation.

# 5 Density of Water

You have been looking at some unique properties of water that are related to its bonding. Other unique properties of water which are related to its bonding and which affect our lives include its high surface tension, its ability to dissolve a vast number of substances and the way its hydrogen bonding causes changes in its density as it freezes.

Density is one of the fundamental physical properties of a substance. It is calculated by dividing the mass of a substance by its volume.

$$\text{Density} = \frac{\text{mass}}{\text{volume}} \text{ or } d = m/V$$

In chemical data books, the density of water is given as 1.0 g $mL^{-1}$. In other words, 1 gram of water has a volume of 1 millilitre. However, the density of water changes with temperature, as shown in the graph.

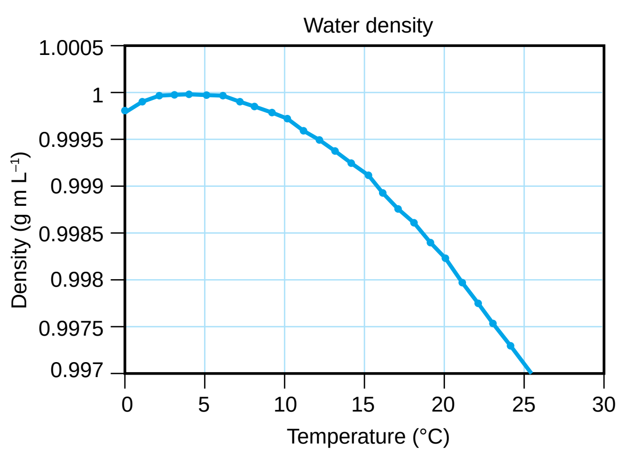



**Figure 5.1** Change in density of water with change in temperature.

This graph shows that temperature changes affect the density of water. For most substances, as you cool them, they shrink, occupying a smaller volume, so becoming more dense. Most reach maximum density at their freezing points. This increase in density also happens to water – but only until the temperature drops to 4°C. At this temperature, water reaches its maximum density. As the temperature drops lower, from 4°C towards its freezing point at 0°C, the volume increases so it becomes less dense again.

Ice floats on the surface of liquid water because ice is less dense than liquid water. This is an important property of water because it allows aquatic organisms to survive in very cold climates, living in the warmer water under the ice.

This low density of ice between 4°C and 0°C is due to changes in the hydrogen bonds between the water molecules. As water turns to ice, its molecules move less and the bonds become more stable. Ice consists of a stable three-dimensional structure of hexagonal rings of water molecules. Because this arrangement of the molecules takes up more volume, ice is less dense than liquid water.

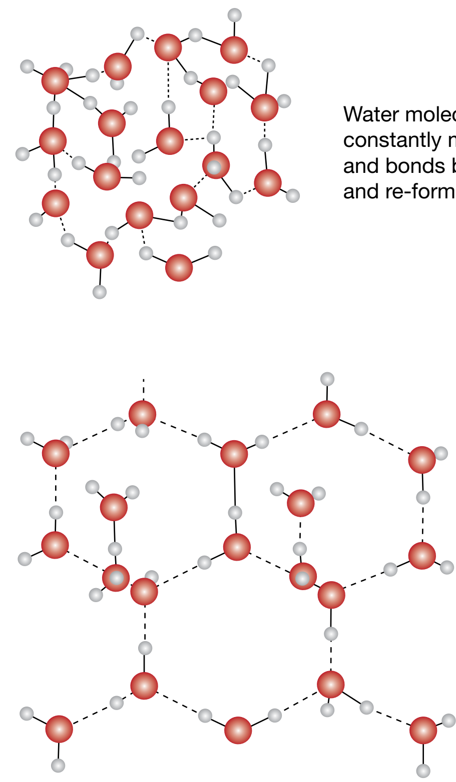



**Figure 5.2** Hydrogen bonding between molecules in liquid water and ice.

## QUESTIONS

1 (a) State the relationship between density, mass and volume of a substance.
(b) The density of water at 8°C is 0.9999 g $mL^{-1}$. Calculate the volume in millilitres of 100.0000 grams of water at this temperature.

2. In cold climates, water pipes sometimes burst. Explain.

3. (a) At what temperature is the density of water at a maximum?
(b) Explain the implications of this for aquatic life in cold climates.

4. Check your knowledge with the following quick quiz.
(a) To calculate density, we divide the mass by its .......................... .
(b) At what temperature is the density of water at a maximum?
(c) As temperature cools below 4°C, the density of water (increases/decreases).

# 6 Surface Tension of Water

Surface tension refers to the forces acting on the surface of liquids and is a measure of the resistance of a liquid to any increase in its surface area.

Table 6.1 compares the surface tension of some liquids and you can see that the surface tension of water is relatively high. Only mercury is higher.

**Table 6.1** The surface tension of some liquids at 20°C.

| Liquid | Surface tension (mN/m) |
|---|---|
| Octane | 21.62 |
| Ethanol | 22.10 |
| Water | 72.8 |
| Mercury | 425.41 |

(a) Round water drops.

(b) A water strider walking on water.

(c) Capillary action – coloured liquid rising up thin capillary tubes against the force of gravity.

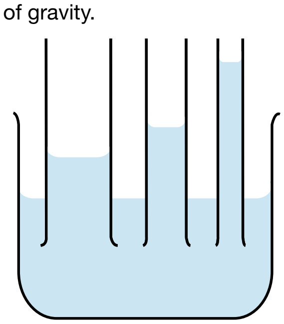


**Figure 6.1** Effects of the high surface tension of water.

The high surface tension of water is responsible for water droplets and bubbles being round. Surface tension makes the surface of water act as if it is a skin, allowing some insects to walk on the surface, and a carefully placed needle to float. Surface tension also causes the capillary action of water in fine (capillary) tubes. This capillary action helps plants to move water, with dissolved minerals, upwards from the soil to their leaves.

The surface tension of water is due to the water molecules being attracted to each other. Attraction of identical molecules to each other is called **cohesion**, and in water this is a strong force due largely to the presence of intermolecular attractive forces, especially **hydrogen bonds** which form between the water molecules.

Water is also attracted to other substances and this is called **adhesion**. The meniscus of water, turning up where it is in contact with glass is a result of adhesion forces between water and glass being stronger than cohesion forces between water molecules. The meniscus of mercury turns down because cohesion forces between mercury particles are stronger than adhesion forces between mercury and the container surface.

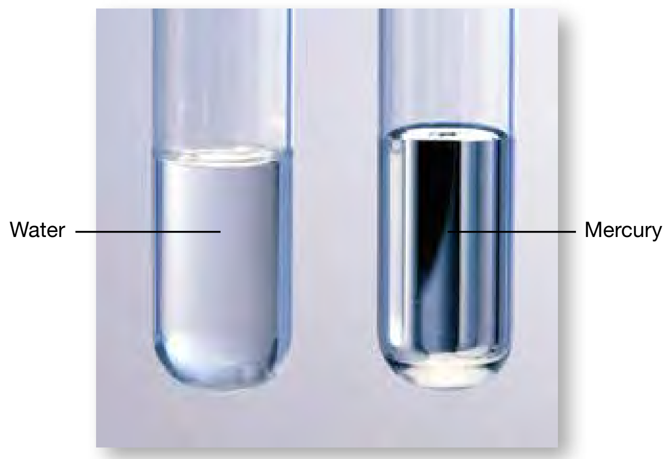



**Figure 6.2** Meniscus of water and mercury.

## QUESTIONS

1. (a) Define surface tension.
   (b) What causes surface tension in water?
2. Outline a first-hand investigation you carried out to investigate surface tension of water.
3. Identify reasons why the surface tension of water is important for life.
4. Check your knowledge with this quick quiz.
   (a) The surface tension of water is higher/lower than most other liquids.
   (b) Surface tension in water is mainly due to strong attractive forces called ………………………… .
   (c) The attraction of water to glass and mercury to glass is called (adhesion/cohesion.)

# 7 Revision of Properties of Water

Here are some multiple choice questions to help you revise the first section of this book. When you have finished, check your answers in the back and see your teacher if you are finding any difficulties.

1. In a molecule of water, the two hydrogen atoms are bonded to a single oxygen atom by:
   (A) Non-polar covalent bonds.
   (B) Polar covalent bonds.
   (C) Ionic bonds.
   (D) Hydrogen bonds.
2. A hydrogen bond can occur between which of the following?
   (A) Two hydrogen atoms in a water molecule.
   (B) A hydrogen atom and an oxygen atom in one water molecule.
   (C) A hydrogen atom and an oxygen atom in a nearby molecule.
   (D) Two oxygen atoms in nearby molecules.
3. Identify the maximum number of hydrogen bonds that can form between one water molecule and other nearby water molecules.
   (A) 1 (B) 2 (C) 3 (D) 4
4. Water has unique properties such as high melting and boiling points and a low density below 4°C because of the presence of:
   (A) Non-polar covalent bonds.
   (B) Adhesion and cohesion.
   (C) Strong dispersion forces.
   (D) Hydrogen bonds.
5. Adding an ice cube to a drink results in:
   (A) Heat from the drink changing the state of the ice.
   (B) Evaporation of the water.
   (C) Coolness of the ice transfering into the water by convection.
   (D) More collision between the water molecules in the drink.
6. For liquid water to change to a gas, which are the strongest bonds that must be broken?
   (A) Covalent bonds.
   (B) Ionic bonds.
   (C) Dispersion forces.
   (D) Hydrogen bonds.
7. When gaseous water condenses:
   (A) Heat is absorbed into the condensing water molecule.
   (B) Heat is released from the condensing water molecule.
   (C) The water molecules move around more freely.
   (D) The water molecules become solid.
8. At what temperature is water the most dense?
   (A) 0°C (B) 4°C (C) 40°C (D) 100°C
9. The property of water which allows it to maintain a relatively constant temperature is its high:
   (A) Boiling point.
   (B) Specific heat.
   (C) Surface tension.
   (D) Density.
10. A hydride is a compound formed between:
   (A) Hydrogen and another element.
   (B) Two hydrogen atoms.
   (C) Elements which have hydrogen bonding.
   (D) Water and an ion.
11. Which compound forms the strongest hydrogen bonding between its molecules?
   (A) Methane.
   (B) Ammonia.
   (C) Water.
   (D) Hydrogen fluoride.
12. As you go down group 16 of the periodic table, the elements become:
   (A) More metallic.
   (B) Less metallic.
   (C) More like water.
   (D) Less like water.
13. The amount of heat needed to raise the temperature of 1 gram of a substance by one degree Celsius (or one kelvin) is called:
   (A) Latent heat.
   (B) The freezing point.
   (C) Specific heat.
   (D) Density.
14. The heat absorbed or released when a substance changes state is called its:
   (A) Melting point.
   (B) Boiling point.
   (C) Specific heat.
   (D) Latent heat.
15. The latent heat of vaporisation of water is the:
   (A) Heat required to raise the temperature of a substance by one degree.
   (B) Heat required to break hydrogen bonds between liquid water molecules.
   (C) Energy required to break H–O bonds in water molecules.
   (D) Energy required to form hydrogen bonds between water molecules.
16. A student raised the temperature of 200 g water in an insulated container. The initial temperature was 21°C and the final temperature was 29°C. The energy change was:
   (A) 6.7 kJ (B) 1699 J
   (C) 33 J (D) 1.6 kJ

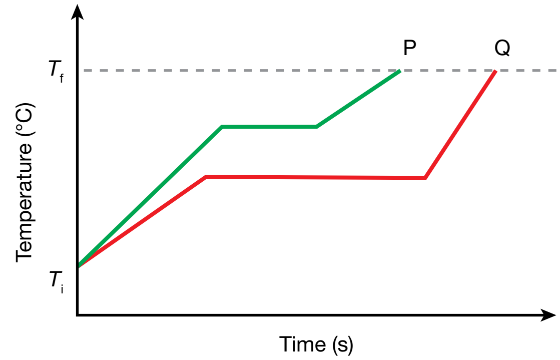
17. Which statement about surface tension is not correct?
(A) Surface tension is a downward force acting on molecules at the surface of a liquid.
(B) The stronger the intermolecular forces, the greater the surface tension.
(C) Water has a high surface tension due to strong hydrogen bonding.
(D) Water has a higher surface tension than mercury.

18. Which of the following statements about the properties of water is correct?
(A) Ice is more dense than water and floats on its surface.
(B) The temperature of water is slow to rise and fall.
(C) Water molecules cling together because of adhesive forces.
(D) Water can wet other substances because of cohesion.

19. Four 1 gram cubes of different metals, were heated to 100°C and then each cube was placed in a separate identical beaker containing 100 mL water at the same temperature. The largest temperature increase would occur in the water containing the metal with the:
(A) Highest latent heat.
(B) Greatest density.
(C) Highest specific heat.
(D) Strongest hydrogen bonding.

20. As a solid changes to a liquid, heat is absorbed and the temperature:
(A) Increases.
(B) Decreases.
(C) Increases and then decreases.
(D) Stays the same.

21. Two solids P and Q of identical mass are heated, using identical heaters, from temperature $T_i$ (initial temperature) to $T_f$ (final temperature). Changes in the temperatures of the masses are shown in the graph.




Choose the most correct statement below.
(A) The latent heat of fusion of P is larger than that of Q.
(B) The specific heat capacity of liquid P is larger than that of Q.
(C) The latent heat of fusion of P equals the specific heat capacity of Q.
(D) This graph has nothing to do with specific heat.

# 8 Solubility of Molecular Substances

The nature and strength of forces between the particles of a substance can explain observable properties such as solubility. In the case of covalent molecules, the relevant forces are intermolecular forces.

When one substance dissolves in another:

- The **solute** is the substance that gets dissolved.
- The **solvent** does the dissolving.
- A **solution** is formed – this is a homogeneous mixture.

Some solvents are polar, e.g. water, and others are non-polar, e.g. kerosene and petrol. Water is such a good polar solvent that it is sometimes called the **universal solvent**.

A substance dissolves when attractive forces form between its particles and particles of the solvent. In general, ionic compounds and polar molecular compounds tend to be soluble in **polar solvents** such as water. Non-polar molecular substances tend to be soluble in **non-polar solvents**, e.g. petrol.

To help you remember this, you can think that 'like dissolves like'. However, this cannot be used as an explanation of why substances dissolve.

## Solubility of molecular substances in water

Most covalent molecules that are non-polar, both elements and compounds, are insoluble in water or only slightly soluble.

Some elements such as **oxygen gas** ($O_2$) are partly soluble in water; the water molecules becoming attached to oxygen molecules by weak dispersion forces.

Sucrose (table sugar, $C_{12}H_{22}O_{11}$) however is a polar covalent compound and it dissolves in water. The intermolecular forces in the sugar crystal break, as do the hydrogen bonds between water molecules. New hydrogen bonds then form – this time between water molecules and sugar molecules, so sugar is soluble.

Some very polar covalent compounds such as hydrogen chloride, hydrogen bromide and hydrogen sulfide react with water molecules to form ions. This is called **ionisation**. ($H_3O^+$ is called a hydronium ion.)

$HCl(g) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$ or

$HCl(g) \rightarrow H^+(aq) + Cl^-(aq)$

Notice that hydrogen chloride gas is covalent, but it forms ionic hydrochloric acid in solution.

Some molecular substances such as carbon dioxide and sulfur dioxide react with water.

$CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$

$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$

Covalent network substances such as silicon dioxide are insoluble in water. Strong covalent bonds exist throughout the silicon dioxide network and water molecules cannot break these strong covalent bonds, so covalent network substances such as silicon dioxide are insoluble (cannot dissolve) in water.

Macromolecules, e.g. in polymers such as polyethylene are also insoluble. The very large molecules of polyethylene have very strong dispersion forces between their molecules. Water molecules cannot separate these large molecules so polymers are insoluble in water.

## QUESTIONS

1. (a) Distinguish between the terms solute, solvent and solution.
   (b) Explain why water is called the 'universal solvent'.
2. (a) Explain why sodium chloride is soluble in water and silicon dioxide is not.
   (b) To part (a) one student answered, '… because like dissolves like.' Explain why this answer scored no marks.
3. Hydrogen chloride and sucrose are both molecular compounds and both dissolve in water. Explain why a solution of hydrogen chloride will conduct electricity and a solution of sugar will not.
4. Copy and complete the table.

| Type of chemical | Example | Solubility |
|---|---|---|
| Polar molecular compound | | |
| Molecular element | | |
| Highly polar molecular compound | | |
| Non-polar molecular compound | | |
| Covalent network structure | | |
| Macromolecules | | |

5. Research the solvent properties of ethanol ($C_2H_5OH$).
6. Check your knowledge with this quick quiz.
   (a) Is water a polar or non-polar solvent?
   (b) Polar solvents dissolve (non-polar/polar) solutes.
   (c) Polar molecules are (soluble/insoluble) in water.
   (d) Non-polar molecules are (soluble/insoluble) in water.

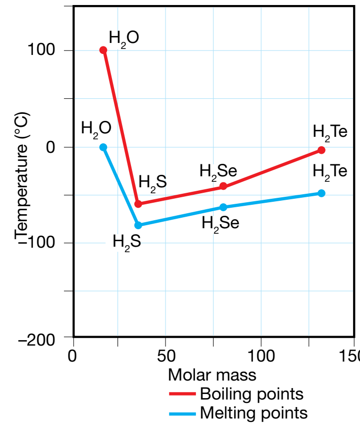
# Answers

## 1 Hydrides and their Boiling Points

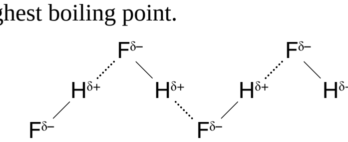
1. (a) A hydride is an anion of hydrogen, a negative hydrogen ion ($H^-$).
The term hydride is also used more generally to refer to compounds formed from hydrogen and another non-metal element.
(b) Various. Covalent hydrides include hydrogen chloride (HCl), ammonia ($NH_3$).
(c) Various. Ionic hydrides include sodium hydride (NaH) and magnesium hydride ($MgH_2$).
2. (a) Boiling point of group 14 hydrides increases down the group.
(b) Down the group, mass increases and the elements become more metallic and thus more ionic. Each of these changes can cause an increase in boiling point.
3. (a) Water, ammonia and hydrogen fluoride.
(b) These three compounds have much higher boiling points than expected based on the boiling points of other similar hydrides.
(c) These three compounds have hydrogen bonding. The other hydrides do not have hydrogen bonding. Having stronger intermolecular forces (due to the presence of hydrogen bonds) causes these compounds to have higher boiling points.
(d) Blue – HCl (hydrogen chloride) and HBr (hydrogen bromide).
Green – $PH_3$ (phosphorus trihydride) and $AsH_3$ (arsenic trihydride).
4. (a) Intermolecular forces.
(b) Dispersion forces, dipole-dipole forces, hydrogen bonds.
(c) Hydrogen bonds.
(d) Ionic bond.
(e) Nitrogen, fluorine, oxygen.
(f) $H_2O$, HF, $NH_3$.

## 2 Melting and Boiling Points of Group 16 Hydrides

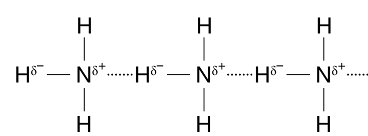
1. (a) *Oxygen* – Discovered in 1771 when Swedish pharmacist Carl Scheele heated oxides (such as manganese oxide and mercury oxide) and also silver carbonate and produced a colourless, odourless gas that supported combustion better than air. Joseph Priestley also discovered oxygen about the same time, and his paper was published first so he is credited with the discovery.
*Tellurium* – Discovered accidentally by Austrian chemist Franz Joseph Muller von Reichenstein in 1782. He found it difficult to analyse and it was not until 1798 that another Austrian chemist, Martin Klaproth, demonstrated that it was a new element and named it.
*Selenium* – Isolated in 1817 by the Swedish chemist Jöns Jakob Berzelius when investigating a foul smelling contaminant in the manufacture of sulfuric acid. (Tellurium and selenium occur as contaminants in sulfide ores.)
*Polonium* – First isolated by Marie Curie, and named after her native country, Poland. She was awarded a Nobel Prize in 1911 for the discovery of radium and polonium.
(b) Various.
*Oxygen* – Used in the manufacture of steel, in oxyacetylene welding and the cutting of metals, as an oxidant in rocket fuel, in the production of the white pigment titanium oxide.
*Tellurium* – Used to colour glass and ceramics, in blasting caps, and in thermoelectric devices. It is added to steel to increase its ability to be machined. Its addition to lead makes the lead stronger, harder and less susceptible to attack by acid.
*Selenium* – Used to decolourise glass, to make red glass and enamel, in solar cells and in photocopying.
*Polonium* – Emits alpha particles and is used as an antistatic device in brushes. It is also used as a source of heat for space equipment.
2. (a)



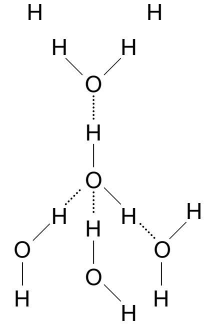

(b) If the melting and boiling points of water followed the trend, water would melt at approximately −120°C and boil at −100°C. There would be no liquid or solid water on Earth, it would all be in the gas state. Without liquid water, there could be no rivers, lakes, oceans, or living things as we know them. Hydrogen bonds keep water liquid over a wider range in temperature than is found for any other molecule of its size, and the evaporation of sweat, essential for the cooling of many mammals, has that cooling effect because of the large amount of heat needed to break the hydrogen bonds between water molecules so they can evaporate.
3. These four substances all consist of small, covalent molecules of similar size, so any differences in their boiling points must be due to differences in their intermolecular forces. They all have similar dispersion forces.
*Methane*, $CH_4$ only has dispersion forces as it is a non-polar molecule and there are no hydrogen bonds between its molecules. Hence it has by far the lowest boiling point.
*Ammonia, hydrogen fluoride and water* all have polar molecules so they have similar dipole-dipole forces between their molecules, although water molecules are more polar than ammonia.
*The main difference lies in their polar bonding*. They all have hydrogen bonds, but these are not the same between the three molecules.
*Hydrogen fluoride* has the strongest hydrogen bonds because fluorine is more electronegative than oxygen or nitrogen. However, *water* can form up to four hydrogen bonds per molecule, whereas ammonia has three hydrogens, but only one lone pair on the nitrogen atom, so it can only form one hydrogen bond per molecule. Hydrogen fluoride also only has one hydrogen atom although it has three lone pairs, so it is also limited to one hydrogen bond per molecule.
Hence the hydrogen bonding between water molecules is the strongest and water has the highest boiling point.
4. (a) Hydrogen bonding between HF molecules.



(b) Hydrogen bonding between $NH_3$ molecules.



(c) Hydrogen bonding between water molecules.



5. (a) Higher.
(b) −75.
(c) Dispersion forces, dipole-dipole forces, hydrogen bonds.

## 3 Specific Heat Capacity of Water

1. (a) Specific heat capacity is the amount of heat energy needed to raise the temperature of 1 gram of a substance by one kelvin.
   (b) Mercury, copper, ethanol, water.
   (c) Land.
   (d) Land has a lower specific heat capacity than water. It heats up more quickly.
   (e) Water maintains a more uniform temperature, so aquatic organisms do not need to cope with the large temperature changes in the environment to which terrestrial organisms are subjected.
2. Various. You may have compared the specific heat of identical pieces of different substances such as brass, aluminium and plastic. You could heat each cube in boiling water and then place it into another container of cool water and see how much the temperature of the water rises. The greater the increase in temperature of the water, the higher is the specific heat of the substance.
3. Various. The high specific heat capacity of water means that water is slow to heat up and then slow to cool down again. This means water can be used to store and transport heat, e.g. in central heating pipes and radiators. Nuclear reactors use water to absorb heat released during radioactive decay and thus cool the reactor, car radiators use water to cool the engine. Water filled radiators are used to heat homes and businesses in cold climates because water once heated retains the heat for a long time.
   Cooking utensils are made of metals with a relatively low specific heat so less heat is needed to heat them. The handles are made of substances with high specific heat so they do not become hot.
4. (a) 5 L of water has a mass of $5 \times 10^3$ g
   $\Delta H = -m \times C \times \Delta T$
   $= -5 \times 10^3 \times 4.18 \times 10$
   $= -209 \times 10^3$ J $= -200$ kJ (to one significant figure)
   200 kJ of heat is needed.
   (b) 400 mL of water = 400 g
   $\Delta H = m \times C \times \Delta T$
   $= 400 \times 4.18 \times 20$
   = 33 440 joules or 30 kJ (to 1 significant figure).
5. (a) $\Delta H = -m \times C \times \Delta T$
   $= -5 \times 10^3 \times 0.4 \times 15$
   $= -30$ J
   30 joules of energy is needed.
   (b) $\Delta H = -m \times C \times \Delta T$
   $-5 \times 0.9 \times 8 = -36$ J
   40 joules of energy is needed (to 1 significant figure).
6. (a) $\Delta H = -m \times C \times \Delta T = -100 \times 4.18 \times 16 = 6688$ J or 6.69 kJ
   (b) 1 mol $C_2H_5OH$ = 46 g
   0.49 g releases 6.688 kJ
   46 g releases $\frac{46}{0.49} \times 6.688 = 627.85$ kJ energy
7. (a) $\Delta H = -m \times C \times \Delta T$
   For water:
   $500 = -100 \times 4.18 \times \Delta T$
   $\Delta T = \frac{500}{418} = 1.196$ K
   For dry soil:
   $500 = -100 \times 2.2 \times \Delta T$
   $\Delta T = \frac{500}{220} = 2.273$ K
   The dry soil shows a larger rise in temperature than the water (using 1 significant figure – 1 K compared to 2 K) when the same amount of heat is added to the same mass of water and of soil.
   (b) For water:
   $\Delta H = -m \times C \times \Delta T$
   $= -100 \times 4.18 \times 3 = -1254$ J or 1.254 kJ
   For soil:
   $\Delta H = -m \times C \times \Delta T$
   $= -100 \times 2.2 \times 3 = -660$ J or 0.66 kJ
   The heat released from the dry soil is only half as much (0.66 kJ) as the heat released from the same quantity of water (1.3 kJ – both to 2 significant figures).
8. (a) Higher.
   (b) Slower, slower.
   (c) More.

## 4 Latent Heat of Water

1. (a) Latent heat is the energy absorbed or released by a substance when it changes state.
   (b) Joules per mole or joules per gram.
2. Latent heat of fusion is the heat change during melting or freezing (solidifying).
   Latent heat of vaporisation is the heat change during evaporation or condensation.
3. Various. For example, measure the temperature of crushed ice in a beaker as it is very gently heated and melts. Graph your results.
4. (a) Steam causes a burn because of its high temperature – the same as boiling water. But when steam touches the skin it also condenses to water and this process gives out more heat energy (heat of vaporisation) which further burns the skin.
   (b) Various, e.g. Much of the Sun's heat absorbed by tropical oceans is used to evaporate surface water. As moist tropical air moves upwards with air currents, water vapour condenses to form rain and this condensation releases heat. Also because the latent heat of vaporisation of water is high, a large amount of heat is transferred to the atmosphere by condensation. Water vapour in the atmosphere travels over land and to higher latitudes, distributing energy and moderating temperatures especially in coastal areas.
   (c) Water has a high heat of vaporisation which means a lot of heat is absorbed to evaporate a small volume of water. This heat is absorbed from our skin, thus having a cooling effect. The skin is cooled without losing too much water from the body.
   (d) Ice in the cooler or esky melts. The heat needed to do this, the latent heat of fusion, comes from the food and drink in the container, thus cooling the food or drink.
5. Various. The heat of vaporisation is used in steam heating plants and in industrial distillation apparatus. In electronics, cooling can occur by vaporisation in a closed system and the heat can be released elsewhere. Examples include heat pumps, pumped refrigerants and spray cooling.

## 5 Density of Water

1. (a) Density is proportional to the mass and inversely proportional to volume of an object and is measured in grams per millilitre (cubic centimetre). This is shown by the following equation.
   $\text{Density} = \frac{\text{mass}}{\text{volume}}$
   (b) $\text{Volume} = \frac{\text{mass}}{\text{density}} = \frac{100.0000}{0.9999} = 100.01$ mL
2. When the temperature drops below 4°C, water starts to expand as it drops to freezing point of 0°C. If there is any water in pipes this will expand as water molecules take up the hexagonal ring structure of ice. It can expand so much that it cracks the pipes.
3. (a) 4°C
   (b) As temperature drops, the volume of a mass of water gets smaller while the mass stays the same. Thus the density increases. But when the temperature of water reaches 4°C, its density is at a maximum – it cannot get any bigger. In fact, as the temperature drops below 4 degrees, the density starts to decrease as the water freezes. This means that ice is less dense than liquid water – so it floats on top of the water. This provides an insulating layer over the water, protecting aquatic life from the frigid conditions on land.
4. (a) Volume.
   (b) 4°C
   (c) Decreases.

## 6 Surface Tension of Water

1. (a) Surface tension refers to the way the surface of a liquid acts as if it is covered with a skin.
   (b) Surface tension is due to the cohesive forces between the molecules of water. These forces attract the molecules together so they seem to form a skin. The high surface tension of water is due mainly to the strong hydrogen bonds that form between water molecules.