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Chemistry

Preliminary Course
Stage 6

NEW SOUTH WALES DEPARTMENT OF EDUCATION AND TRAINING

TAFE NSW

OPEN TRAINING AND EDUCATION NETWORK - DISTANCE EDUCATION

Water

Incorporating October 2002
AMENDMENTS

Acknowledgments

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- Board of Studies NSW. *Chemistry Stage 6 Syllabus*. **Amended October 2002**
  The most up-to-date version is to be found at
  http://www.boardofstudies.nsw.edu.au/syllabus_hsc/index.html
- Messel, H (chair)). (1964.) *Science for high school.* The Foundation for Nuclear Energy. University of Sydney.

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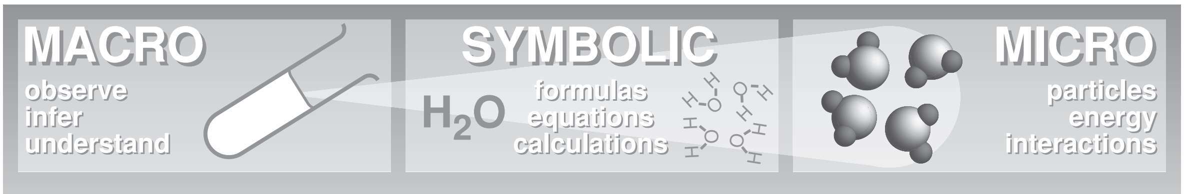
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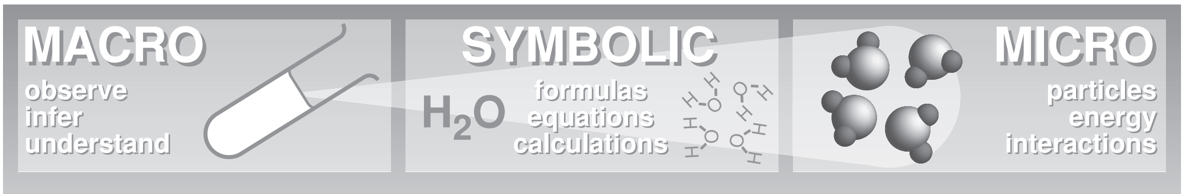
H2O
formulas
equations
calculations
MICRO
particles
energy
interactions

Contents



Module overview

The extracts below are from the Contextual Outline in the *Chemistry Stage 6 Syllabus*, Board of Studies NSW, November 2002.

> ‘The first astronauts who viewed the Earth from space commented on the beauty of our water-rich blue planet. Earth’s position in the solar system enables its retention of water in solid, liquid and gaseous forms on and around its surface. The particular properties of the water molecule assisted the evolution of life and continue to support life processes by maintaining a narrow temperature range on the Earth’s surface.’

The Earth is the only water rich planet in the solar system. Its appearance from space is dominated by the blue colour of thick layers of water, the white reflection of sunlight from ice crystals and water droplets in clouds, the brown landscape shaped by water and the green of water rich plants. Water is the only inorganic compound that is found as a liquid, the most abundant liquid on Earth and the only chemical substance found in solid, liquid and gaseous states where life is found. Nature’s fluid – water – moves between the spheres spreading heat energy absorbed from the sun and keeping the average temperature of the Earth’s surface close to 15°C. No known living thing can exist without water. Not all living things require oxygen (and in fact some are poisoned by it) but all living things need water. Water transports many chemicals between the living and non-living worlds.

> ‘The concepts of bonding and intermolecular forces are used to increase understanding of the special nature of the water molecule. The chemistry of solutions is examined in greater detail.

The unusual properties possessed by such a small molecule as water can be explained by the intramolecular bonding within a water molecule and the intermolecular forces between water molecules. Nature’s fluid dissolves more substances than any other solvent and transports them to, through and out of the bodies of living things. Water dissolves most substances without reacting with them and the fluid nature of the water ensures good contact between dissolved reactants.

The emphasis in this module is on Prescribed Focus Areas 2 and 4.

Indicative time

This module is designed to take a minimum of thirty hours. There are a number of practical activities and opportunities to complete an open–ended investigation integrating the skills and knowledge and understanding outcomes. Organising materials and equipment for carrying out all these activities could take additional time but in doing so you will better understand the type of work chemists do.

Resources

Materials and equipment you need to carry out activities are listed below. Access to a computer and the internet are sometimes important for the study of modern chemistry. An important skill to develop in chemistry is planning ahead and thinking things through before carrying out the action. Make sure the resources you need are available when you start an activity.

For Part 1 you will require:

- three small containers of uniform cross-section. The plastic containers that photographic film is supplied in are suitable. Ask at a photographic film processing shop for these containers as they are often discarded when a film is processed. The black plastic containers are best.
- a freezer or freezing compartment of a refrigerator
- half a cup of ice
- equipment to boil water
- a rule marked in millimetres (mm)
- an egg (raw and in its shell)
- a small saucepan of water
- up to four heaped tablespoons of table salt
- ice cubes
- two glasses of the same size and type

- a means of boiling mixtures of salt and water
- a thermometer reading to at lest 110°C
- ways of measuring mass or volume of table salt and water so that you can prepare aqueous solutions of different concentrations.
- three straight drinking straws
- a fine marker pen
- three small, wide mouthed heat resistant containers
- three pieces of clay/Blu-tack®
- some food colouring solution

For Part 2 you will require:

- modelling balloons
- an object that can be electrically charged by rubbing against another object, for example:
  - a plastic rule/pencil/biro and a woollen jumper
  - a blown up balloon and a woollen carpet

  Make sure the object and material used are dry.
- a tap where you can get a continuous thin stream of water
- clean dry paintbrush
- paper clip/dull razor blade/needle
- water insoluble powder eg. baby powder, talcum powder or cornflour
- one drop of oil
- liquid detergent
- glass of water
- glycerol
- two identical test tubes and a dense bead to fit into the test tubes OR

  funnel, plastic or rubber tube, clamp, eyedropper, liquid volume measurer.

For Part 3 you will require:

- table salt
- table sugar
- kerosene/mineral turps/candle wax
- graphite/sand

- cottonwool
- polyethylene plastic eg plastic for food bags, milk container
- six small test tubes
- anhydrous copper(II) sulfate ($CuSO_4$).

For Part 4 you will require:

- Epsom salts (magnesium sulfate crystals) from a supermarket, health food shop or chemist shop
- two containers of 200 to 1000 mL volume
- about 20 cm of drawstring cord (eg from an old swimming costume or pyjama pants) or rope or thick string
- a place where you can leave your equipment undisturbed for a week eg a low shelf away from pets or top of a flat toilet cistern.
- two ice blocks and a rubber band
- nine dropper solution bottles containing soluble salt solutions (5% salt by mass) separately containing the
  - cations Ba^{2+}, Ca^{2+}, Fe^{2+}, Cu^{2+}, Pb^{2+}
  - anions Cl^-, SO_4^{2-}, CO_3^{2-}, OH^-
- a flat plastic sheet such as an overhead projector sheet

  or 20 small test tubes

  or a plastic well plate onto or into which a drop of cation solution and a drop of anion solution can be placed
- a dark surface on which the sheet/test tubes/well plate can be placed for viewing
- cordial or fruit juice for dilution
- sugar cubes in a packet with information on the weight of each cube or the total weight and the number of sugar cubes
- containers and volume measuring equipment in the kitchen
- iceblocks, preferably cube shaped.

For Part 5 you will require:

- an electric kettle
- a watch timing to seconds
- a way of weighing water or measuring the volume of water
- two balloons
- matches
- small compact objects made of metal

- a means of weighing the small compact metal objects or a knowledge of their weights (metal masses with their masses marked are useful)
- a saucepan in which to heat water and metal objects
- a simple calorimeter with thermometer readable to the nearest degree Celsius. Examples are shown on page 8 of Part 5.
- a means of measuring the volume of water used in the calorimeter
- fork or tongs or tweezers or attached string to transfer metal from hot water to cold water
- a means of weighing solute to be dissolved
- a spoon for handling solute and stirring solution
- at least two suitable solutes such as:
  - sodium chloride $NaCl$ (table salt)
  - ammonium nitrate NH_4NO_3 (fertiliser)
  - sodium hydrogen carbonate $NaHCO_3$ (baking soda or bicarb soda).

Icons

The following icons are used within this module. The meaning of each icon is written beside it.

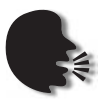
The hand icon means there is an activity for you to do. It may be an experiment or you may make something.

You need to use a computer for this activity.

There is a safety issue that you need to consider.

There are suggested answers for the following questions at the end of the part.

There is an exercise at the end of the part for you to complete.



The talk icon guides you to discuss a topic with others.

Additional resources

McIntyre A.K. (1977) *Water: Planets, Plants and People* Australian Academy of Science ISBN 0858470462

Smith, D. I. (1998) *Water in Australia* Oxford University Press ISBN 0195537041

Bucat, R.B. (1984) Elements of Chemistry, Vol. 2. Australian Academy of Science ISBN 0858471140

VisChem Videos from Video Education Australasia, VEA Multimedia

Video 1 The molecular world of water

Video 2 The molecular world of reactions in water – Part 1 dissolving, precipitation and complexation

Video 3 The molecular world of reactions in water – Part 2 ionic equilibrium, acid/base & oxidation/reduction chemistry.

CD ROM animation from the videos

Glossary

The following words, listed here with their meanings, are found in the learning material in this module. They appear bolded the first time they occur in the learning material.

alpha helices (α helices)	coiled part of protein molecule
amylose	soluble starch molecule
analogy	similarity
antibiotic	chemical that inhibits the growth of or destroys certain microbes
anti-freeze	chemical used to lower the freezing point of water
aqueous	containing or based on water
beta strand (β strand)	flat part of protein molecule
biogeochemical cycle	revolving series of changes involving movement of atoms through living and non-living things
bonding electron pair	pair of valence electrons involved in a covalent bond
brittleness	tendency of a solid structure to fracture
calorimeter	equipment used to measure heat change
calorimetry	measurement of heat change
capillarity/capillary action	ability of a liquid to rise up a narrow tube (capillary) against the force of gravity
closed system	a system that does not allow particles to enter or leave
convection currents	movement of heated parts of a liquid or gas
correlation	connection between values of two different quantities
coulomb (C)	unit of electric charge
coulomb metre (Cm)	10^{-30} Cm is used to measure dipole moments
delta H (ΔH)	change in heat energy
delta T (ΔT)	change in temperature
density	$\frac{\text{mass}}{\text{volume}}$

dichotomous key	table or diagram dividing a group into two smaller groups with different characteristics
dielectric constant	measure of the polarity of a medium; 80 for water
diffusion	movement of particles from a region of high pressure or concentration
dipole	polar molecule
dipole-dipole force	force of attraction between dipoles
dipole moment	measure of the strength of a dipole; measured in coulomb metres (Cm)
dispersion force	force of attraction occurring between all molecules
dissolution	dissolving
dynamic equilibrium	all chemical equilibria are dynamic because, although the composition of the reaction mixture is not changing, change is occurring at the micro level
electron cloud	region in space around a nucleus where electrons are likely to be found
electrostatic force	force between charged particles or objects; like charges repel, unlike charges attract
entropy	capacity to undergo spontaneous change/energy distribution/disorder/inability to do work
equilibrium	chemical equilibrium occurs when the rate of forward reaction equals the rate of the reverse reaction
fluid	material that flows – gas or liquid
glacier	river of ice
hazard	cause of a risk; a potential source of harm
heat of fusion	heat required to melt solid to liquid
heat of solution	heat change when a solute dissolves in a solvent
heat of vaporisation	heat required to vaporise liquid to gas
heat insulator	material with low heat conductivity
helices; helix	spirals; spiral
herbicide	chemical to kill unwanted plants

hGH	human growth hormone – a small protein (polypeptide) in human blood
hydration	reaction with water/attachment of polar water molecules to polar molecules or ion
hydrogen bonding	bond between hydrogen attached to N/O/F and N/O/F attached to a hydrogen in a nearby molecule
hydrophilic	water loving/water attracting group or molecule
hydrophobic	water fearing group or molecule
iceberg	detached part of a glacier floating on water
immiscible	not mixing together
interface	surface which separates two phases
intermolecular forces	forces of attraction between molecules – dispersion forces between all molecules, dipole-dipole forces between polar molecules, hydrogen bonding between molecules with a hydrogen bonded to N, O or F
intramolecular bonding	strong forces of attraction between atoms in molecules; covalent bonding
isoelectronic	same number of electrons
key	table or diagram used to classify a group into smaller groups
molar solution	solution measured in moles per litre; abbreviation M
molarity	moles of solute per litre of solution
net ionic equation	an equation showing reacting ions but not spectator ions
nonbonding electron pair	pair of valence electrons not involved in a covalent bond
nutrient	chemical that sustains life by promoting growth, replacing loss or providing energy
orientation	arrangement in space
pathology	study of the nature of disease
pesticide	chemical to kill a pest
pollution	any harmful effect on the natural environment caused by the release of an unwanted substance

physical hardness	resistance of a solid surface to scratching
precipitation	falling out of a solid from solution when solutions are mixed
radioactive	containing nuclei which can emit particles or high frequency electromagnetic radiation
relative viscosity	viscosity of a liquid compared with another liquid, usually water
reparting style	way of presenting information
residence time	time something is resident in a place
reversible reaction	reaction that can go in reverse as well as forward
risk	exposure to the chance of injury or loss
saline solution	salt solution in a medical context; solution of salt (s) in which cells can be placed without osmosis occurring
saline water	salty water, especially fresh water contaminated with salt
saturated solution	solution which will not dissolve any more solute
sedative	chemical reducing irritability and excitement
simulate	to have the appearance of
solute	substance that dissolves
solution	mixture of solute and solvent
solvent	liquid able to dissolve another substance
specific heat (capacity) C	joules of heat energy required to raise the temperature of 1 g of a substance by 1 K equal to kilojoules of heat energy required to raise the temperature of 1 kg of a substance by 1K
stalactite	$CaCO_3$ growth from the ceiling of a cave
stalagmite	$CaCO_3$ growth from the floor of a cave
surface tension	energy required to increase the surface area of a liquid; measured in J/m^2
surfactant	surface-active agent – chemical that concentrates at an interface
terminology	system of terms belonging to a science
terrestrial habitat	land place where life lives or grows

tetrahedral shape	shape when central atom has four electron clouds around itself
tetrahedron	four sided triangular pyramid
venn diagram	diagram using circles or ellipses to shown connections between sets (groups)
viscosity	resistance of a liquid to flow
volumetric analysis	quantitative analysis using volumes of known concentration solution
VSEPR	Valence shell electron pair repulsion model

Chemistry
Preliminary Course
Stage 6

NEW SOUTH WALES DEPARTMENT OF EDUCATION AND TRAINING
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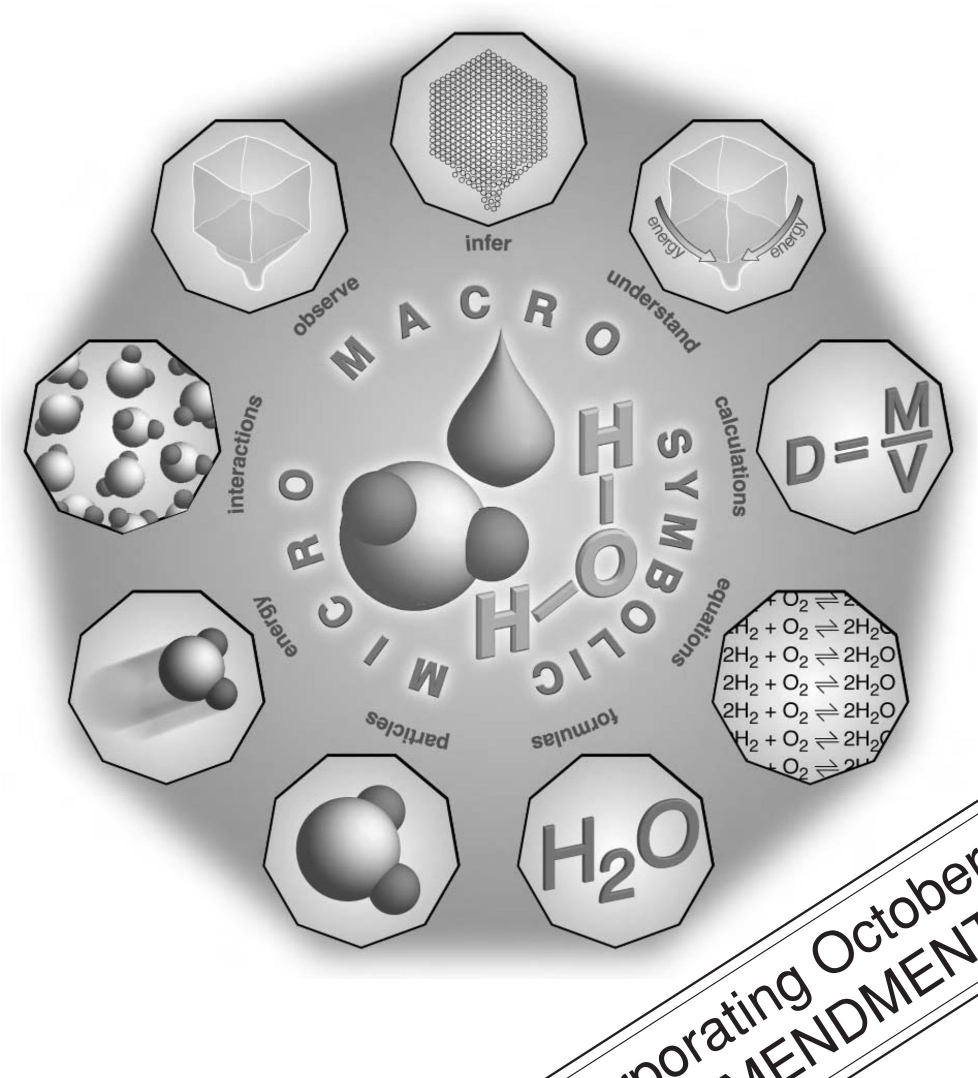
Water

Part 1: Water on Earth



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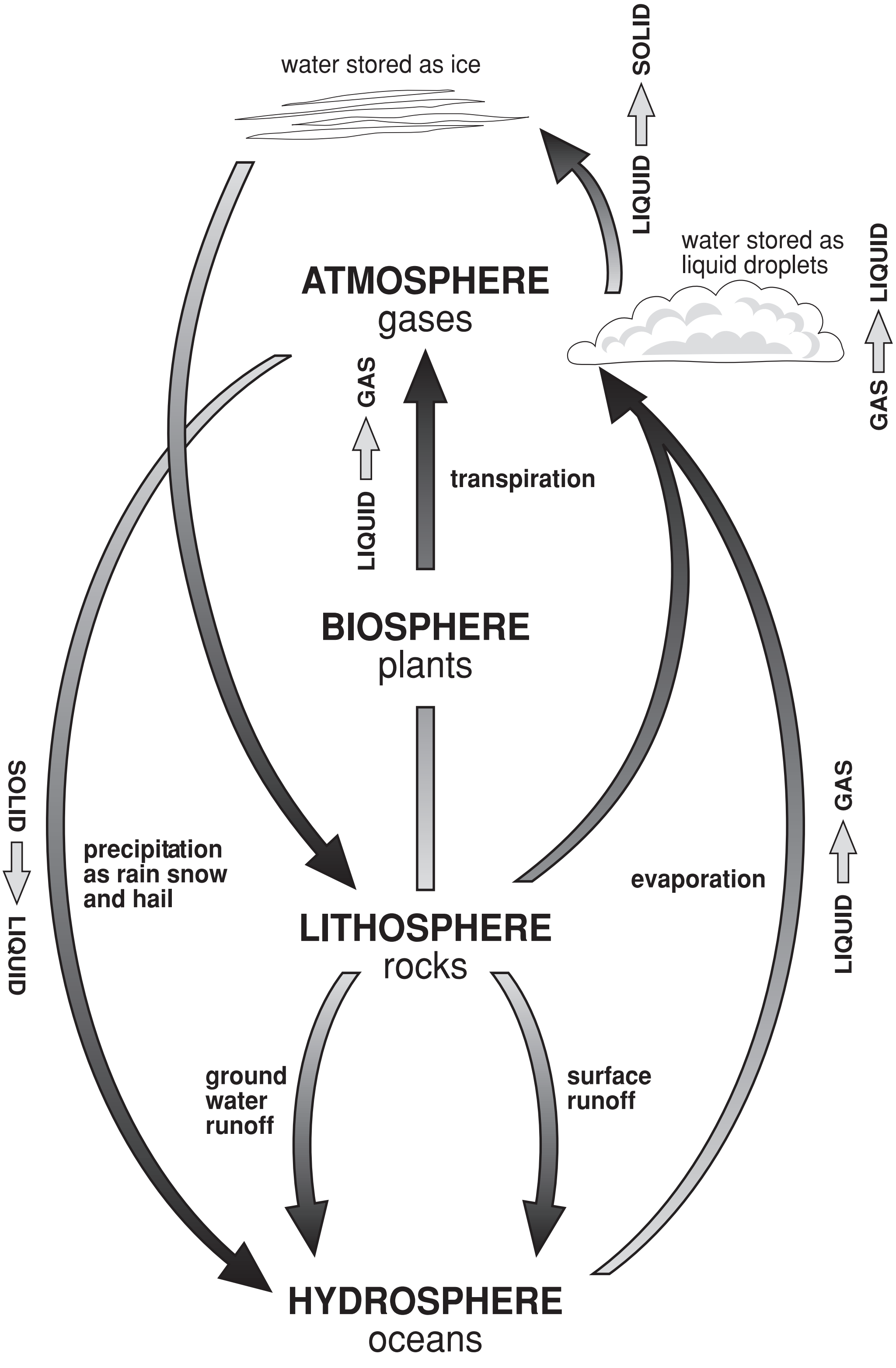

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BIOSPHERE
plants
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rocks
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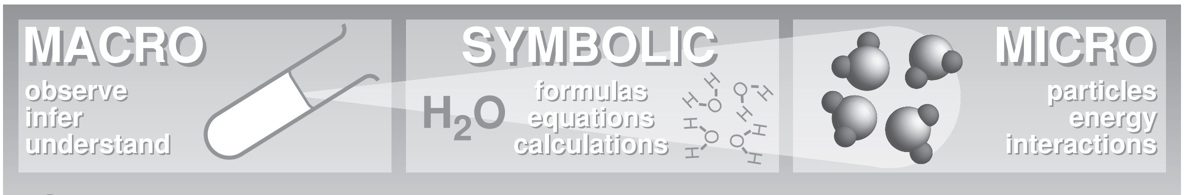
surface runoff
HYDROSPHERE
oceans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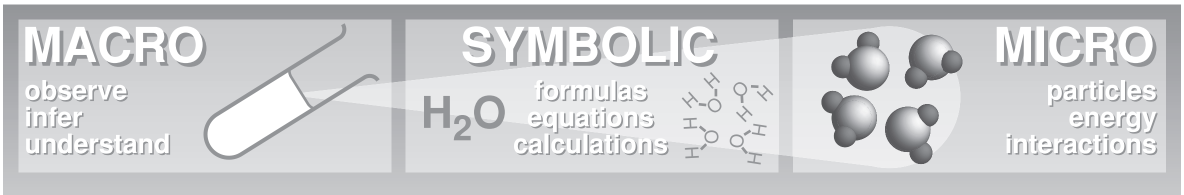
Contents



Introduction

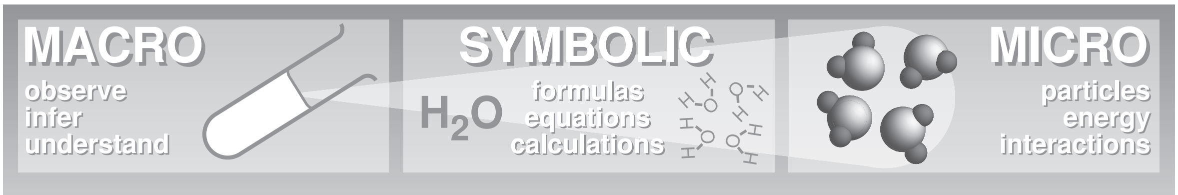
Water is the only chemical substance which can be found on Earth in all three states – solid, liquid or gas – where life exists. Water is a transparent, tasteless, odourless compound of hydrogen and oxygen that has unusual physical properties. Water has carved and moulded the surface of the Earth and its movement in the hydrosphere and atmosphere determines climate.

In Part 1 you will be given opportunities to learn to:

- define the terms, solute, solvent and solution
- identify the importance of water as a solvent
- compare the state, percentage and distribution of water in the biosphere, lithosphere, hydrosphere and atmosphere
- outline the significance of the different states of water on Earth in terms of water as:
  - a constituent of cells and its role as both a solvent and a raw material in metabolism
  - a habitat in which temperature extremes are less than nearby terrestrial habitats
  - an agent of weathering of rocks both as liquid and solid
  - a natural resource for humans and other organisms.

In Part 1 you will be given opportunities to:

- perform an investigation involving calculations of the density of water as a liquid and a solid using: $density = \dfrac{mass}{volume}$
- analyse information by using models to account for the differing densities of ice and liquid water
- plan and perform an investigation to identify and describe the effect of anti-freeze or salt on the boiling point of water.



Water distribution in the hydrosphere

The figures in the table below are estimates of the percentage (%) of water in different parts of the Earth. The mean **residence time** in the last column is the average time a water molecule would spend in that location eg. three days on average in living things.

Water location	State	Percentage of total water (%)	Percentage of fresh water (%)	Mean residence time
oceans		96.5	–	2600 years
ice caps and glaciers		1.8	69.6	1100 years
lithosphere (groundwater)		1.7	30.1	700 years
lakes		0.0013	0.2	13 years
soil		0.0012	0.05	155 days
atmosphere	s, l, g	0.0010	0.04	8 days
biosphere (living matter)	l, g	0.0001	–	3 days

1 Using *s* for solid, *l* for liquid and *g* for gas complete the state column. You only need to put one state, the main state, for each location.

2 Estimate the percentage (%) of total water on Earth in the:

a) solid state ______________________________

b) liquid state ______________________________

c) gaseous state. ______________________________

3 A human body of about 60 kg eats 1 kg of water, drinks 1 kg of water and makes 0.25 kg of water each day. The body breathes out 0.5 kg of water, sweats 0.5 kg of water and urinates about 1.25 kg of water each day.

Using the above information for a 60 kg human calculate:

a) the total water input per day ____________________________

b) the total water output per day. __________________________

c) the mean residence time for water in a human body
(assume the body is 60% water by mass)
= mass of water in body/mass of input per day.

The mean residence time for water in a human body is greater than the figure for all living matter. Most living things live in water and more water passes through their bodies each day than through a human body.

____

____

4 About 10% of the surface freshwater runoff is used by humans.
Most of this is consumed and not returned to runoff. Some is used for hydroelectricity generation and cooling of industrial plants and returned to runoff. In Australia total consumed water is about 2200 L per person per day. 1560 L of this is used for irrigation of plant crops, 320 L for urban domestic use and the remainder for industry and commerce.

a) Knowing that the density of water is 1 g/cm^3 = 1 kg/L calculate how many tonne (1000 kg) of water are consumed per person per day in Australia.

____

____

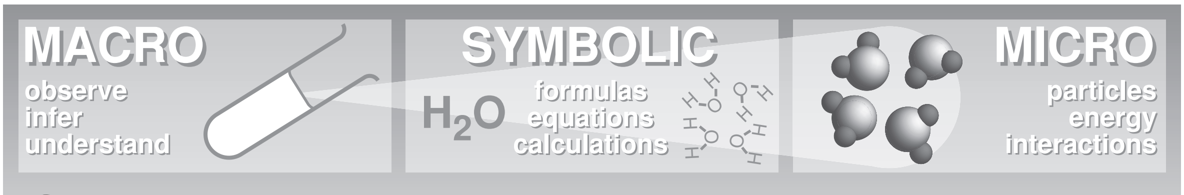
b) What percentage of Australia's consumed water is used for irrigation?

____

c) What percentage of the consumed water would be specially treated with filters and chemicals before human use?

____

Check your answers.



Solvents

solute + solvent → solution.

The solute is a substance that dissolves; the solvent is a liquid able to dissolve the solute; the solution is the mixture of solute and solvent.

Practically every solution you have studied in chemistry so far has been an **aqueous** solution – a solution where the solvent is water.

solute + water → aqueous solution

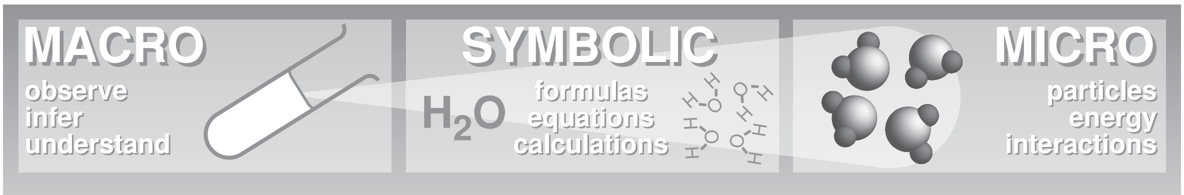
The fluids that flow through the bodies of plants and animals are aqueous solutions. The sea is an aqueous solution of many salts. Many of the mineral deposits that humans mine today were formed from aqueous solutions millions of years ago. No other solvent can dissolve as many different chemical substances as water.

As water cycles through the atmosphere, lithosphere, hydrosphere and biosphere it often acts as a solvent. Acid gases are dissolved out of the atmosphere to form acid rain. Water flowing through the lithosphere dissolves salts which are transported to the seas. The chemical reactions that provide energy to keep living things alive occur in water solutions.

The water that flows through the spheres of the earth consists of moving particles and so is fluid – either liquid or gas.

Solid ice particles are in fixed positions but the movement of solid ice can have a major impact on human life. Think of snow avalanches, the dangers to shipping of **icebergs** and the way **glaciers** moving over rocks scrape the surface to powder which can form soil.

For web sites that provide information about the importance of water monitoring and research in Australia, NSW and Sydney refer to the chemistry web pages at http://www.lmpc.edu.au/science



Density of water

Investigating the density of water

In this activity you will compare the density of cold water from melted ice, tap water and recently boiled water.

You may find it useful to review the section on calculating **density** in Part 1 of the first module *The chemical earth.*

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

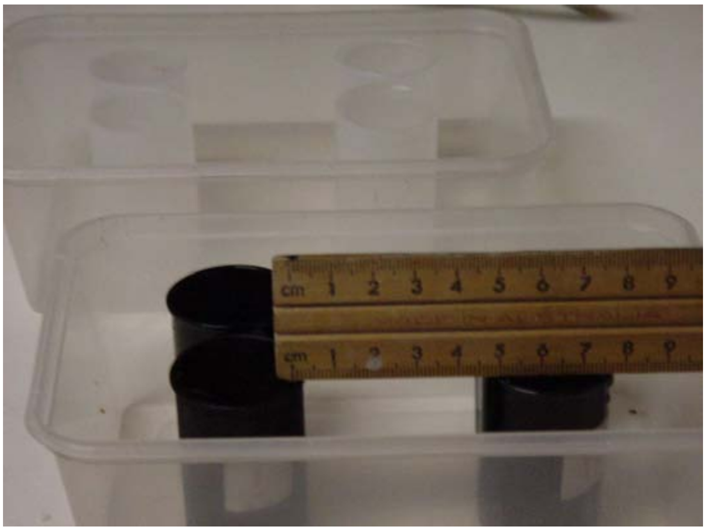
You will need:

- three small containers of uniform cross-section. The plastic containers that photographic film is supplied in are suitable. Ask at a photographic film processing shop for these containers as they are often discarded when a film is processed. The black plastic containers are best.
- a freezer or freezing compartment of a refrigerator
- half a cup of ice
- equipment to boil water
- a rule marked in millimetres (mm).

Method:

1. Add about a quarter of a cup of tap water to the half cup of ice.
2. Boil at least half a cup of water.
3. Mark the three containers – *C* for cold, *T* for tap and *H* for hot water.
4. Separately fill the containers to overflow with:
   – cold water from the cup of ice
   – tap water from a tap
   – hot water from the boiled water.

5 Hold the rule horizontally and move it over the top of each container to remove the curved upper section of water. This ensures that each container is filled to a level top and contains the same volume of water.



6 Place the three containers into the freezer and leave them for at least five hours for the water to freeze.

7 When ice has formed in each container compare the amounts of ice.

Results:

1 When the contents of all three containers have frozen to ice what do you notice about the volume of ice compared with the volume of water you started with?

2 Measure and record the height of the ice above the top of each container:

cold water ________________________________

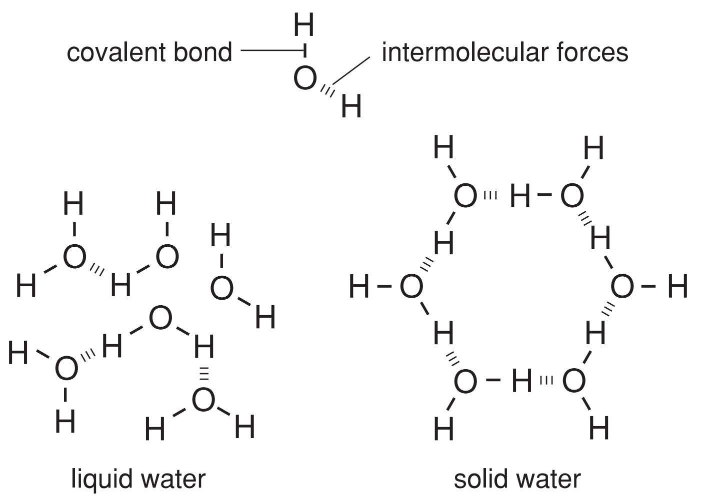
tap water ________________________________

hot water ________________________________

3 Consider the following statements.

- Hardly any of the water you started with in each container would have evaporated in the cold of the freezer.
- The equal volumes of liquid water in the containers at the start of the activity were at different temperatures.
- Each of the containers contain ice at the same temperature.

Use the diagram below to explain why the ice has a large volume than the liquid water.



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The expansion of water when it freezes is important in the breakdown of rock to smaller pieces. Liquid water can flow into and collect in rock cracks.

As the temperature drops, the liquid freezes producing tremendous forces which cause the crack to widen. Over time the rock breaks apart. This process and the grinding of ice over rock in glaciers help form soils.

4 Which of the three ice containers holds the:

- greatest volume of solid water ________________
- smallest volume of solid water? ________________

5 Which of the three ice containers holds the:

- greatest mass of water ______________________________
- smallest mass of water? ______________________________

6 Which container of liquid water that you started with had water at the greatest density? ______________________________

7 Which container of liquid water that you started with had water at the smallest density? Explain how you came to this conclusion:

____

____

8 How does the density of liquid water vary with temperature?

____

____

9 How does the density of solid water compare that of liquid water?

____

____

10 Explain why ice always floats on liquid water.

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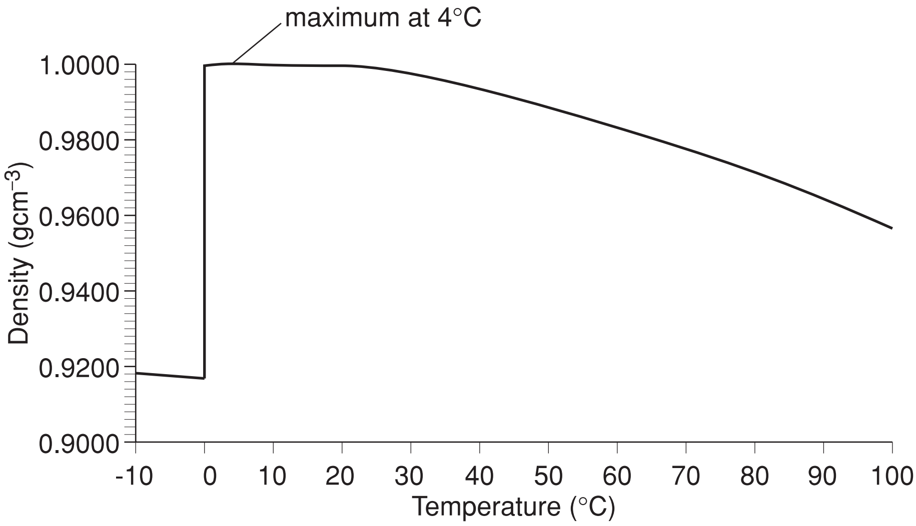
Density and temperature



Change in density of pure water with temperature.

Use the information in the graph to answer the following questions.

1 The average temperature on the surface of the earth is about 15°C. As the temperature of liquid water drops from 15°C to 0°C describe how its density changes.

____

____

____

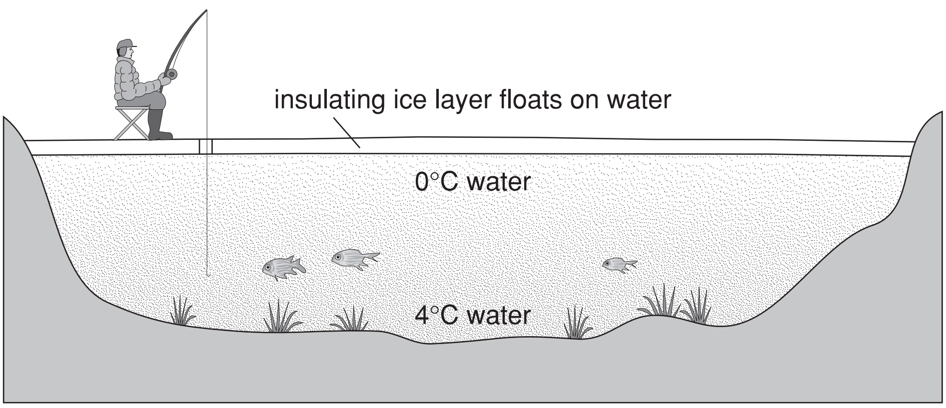
2 At what temperature does water reach its maximum density?

____

Check your answers.

Density in fresh water and salt water

When water at the surface of a freshwater lake is cooled it becomes denser and sinks. It is replaced by less dense water with a higher temperature. **Convection currents** continue until eventually all the water in the lake will be at the maximum density temperature. The surface layer of water then freezes to solid water. This surface layer of ice floats on the liquid water and acts as a **heat insulator**. It takes very cold weather for the liquid water below ice to freeze and thus most lakes do not completely freeze.
This provides a sanctuary for living things.



The formation of an ice layer on a water body.

In this activity you will make predictions, observe and explain observations. You will compare the density of salt water to fresh water.

You will need:

- an egg (raw and in its shell)
- a small saucepan of water
- up to four heaped tablespoons of table salt
- ice cubes
- two glasses of the same size and type.

1 Predict what will happen if you place a complete raw egg into a small saucepan of fresh water. ____________________________________

Did the egg float or sink? ____________________________________

If the egg floated it could be bad. It should be discarded and replaced with an egg that sinks. Decaying eggs release gases including hydrogen sulfide H_2S (rotten egg gas). This decreases the overall density of the bad egg so that it floats in fresh water.

2 Predict what will happen to the egg if you add salt to change the fresh water in the saucepan to salt water.

____

Add two heaped tablespoons of salt to the water and stir until the salt dissolves. If no change occurs try dissolving a second lot of salt. Any floating is not due to gas production inside the egg.

Explain why an egg that sinks in fresh water floats in salt water.

____

3 Using simple labelled diagrams predict how much of an iceberg would be above the water level in salt water compared with fresh water.

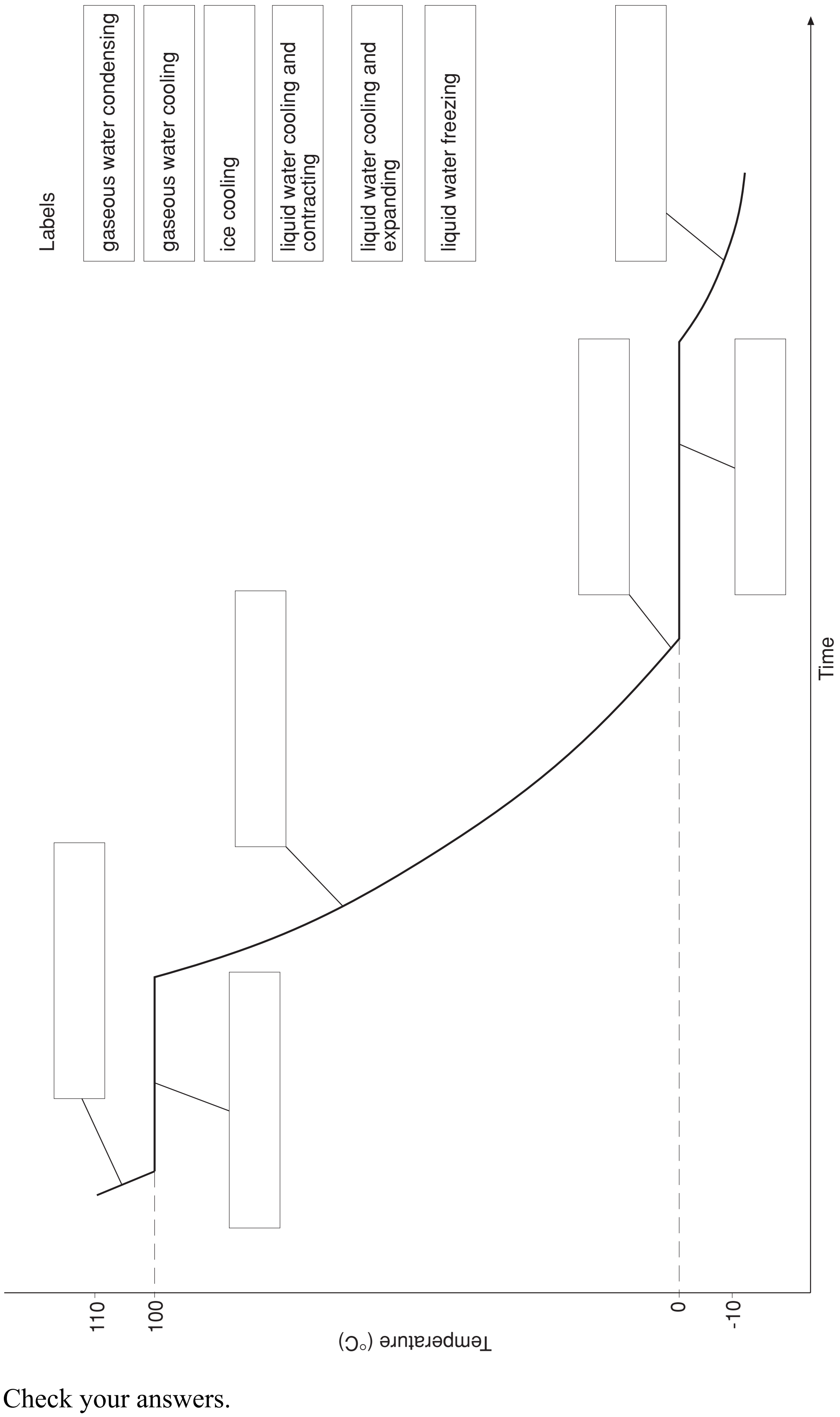
Check your prediction by observing the floating of an ice cube in a glass of fresh water compared with salty water. Explain your observations.

____

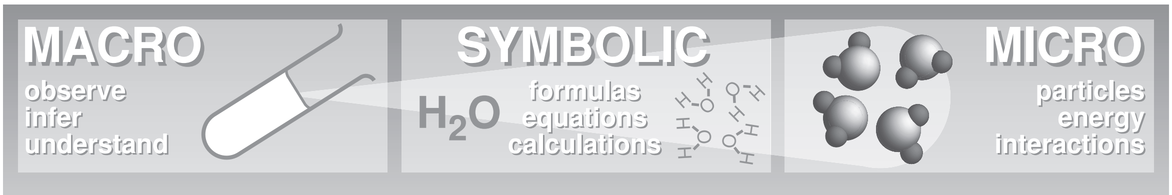
____

Check your answers.

Use the six labels to label six parts of this curve for the cooling of a water sample.



Check your answers.



An open ended investigation

Planning an open ended investigation

If you have a *Resource book* read the section on Experiment design skills. This will enable you to design an experiment which is:

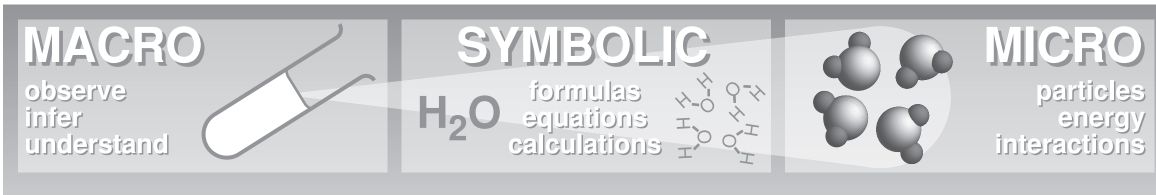
- fair
- valid (leading to effective results and worthwhile conclusions) and
- reliable (trustworthy).

The investigation you need to plan is to identify and describe the effect of table salt (sodium chloride, NaCl) on the boiling point of water.

You will need:

- a means of boiling mixtures of salt and water
- a thermometer reading to at least 110°C
- ways of measuring mass or volume of table salt and water so that you can prepare aqueous solutions of different concentrations.
  You may find density values useful for your calculations. Take the density of table salt as 2.0 g cm^{-3}.

Once you have worked out what materials and equipment you have available turn to Exercise 1.2. Draft your plan. Do NOT perform the investigation until your plan has been returned to you with comments from your teacher.



Biogeochemical cycles

Take the word **biogeochemical.** Can you divide it into three parts?

a) meaning life or living things ________________________________

b) meaning the Earth ____

c) meaning substance used or produced by a chemical process.

____

Check your answer.

The water cycle

The water cycle you can see on the inside cover of this part shows water changing from state to state as it moves through the environment. These state changes are physical changes because no new substance is produced. Physical changes are characteristic of the water cycle.

While cycling through living things and the surface of the Earth, water molecules can also undergo chemical changes. The water become part of the chemicals in the biosphere and lithosphere eg HNO_3 produced by lightning contains H and O from water molecules. When this happens the water is part of biogeochemical cycles which provide the nutrients for all living things. Water is essential as a reactant and a solvent in the cycling of C, O, N, P and S in nature.

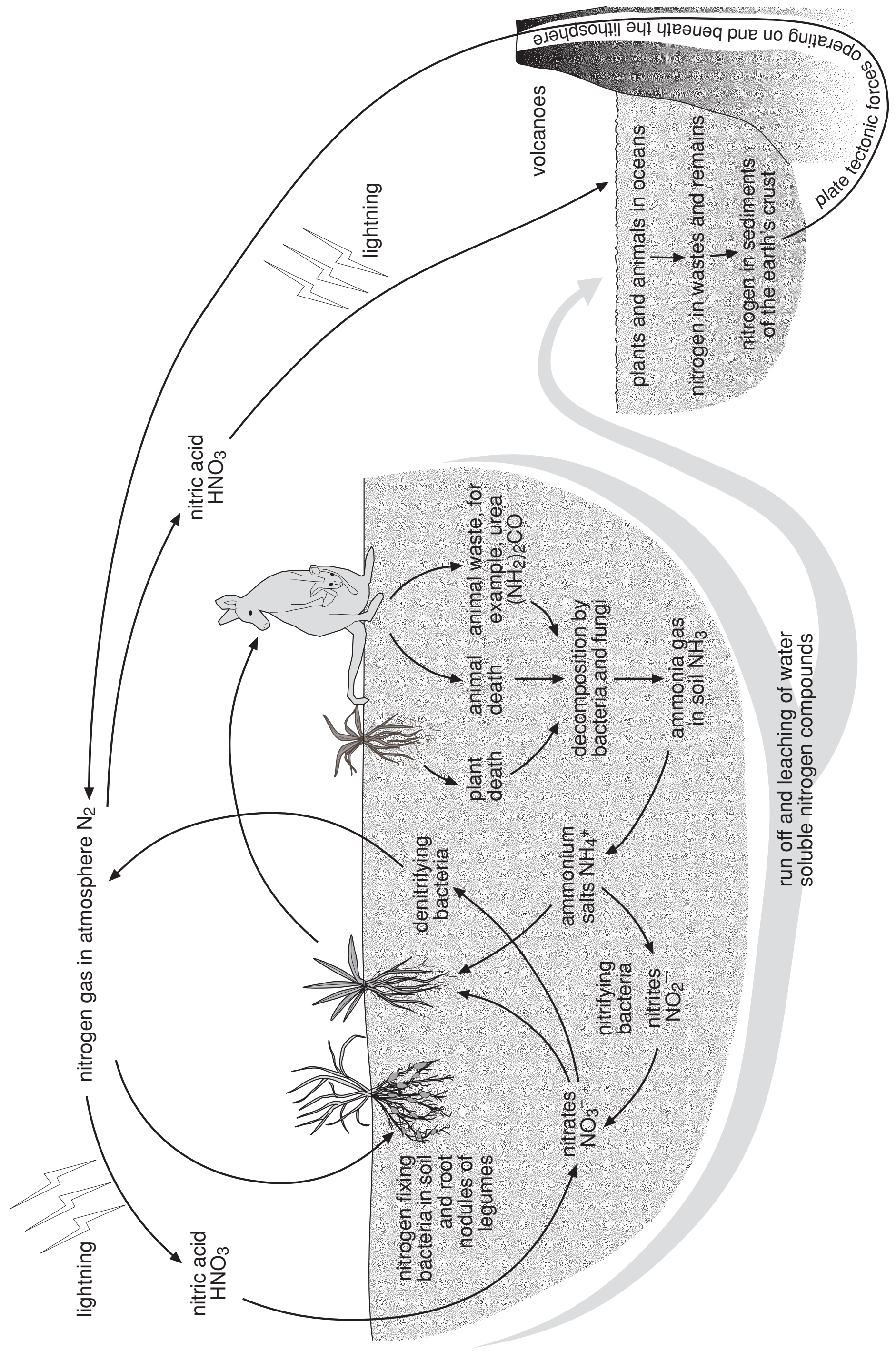
The main function of water in living things is as a solvent. Most chemicals in living things such as gases, sugars, acids, salts and so on are water soluble. Sometimes the water undergoes a chemical change forming new substances such as in photosynthesis. Water is part of many chemical changes in the cycling of C, H, O, N, P and S between living things, the atmosphere, the lithosphere and the hydrosphere.

The nitrogen cycle

Remembering that water is usually at least 50% of every living thing:

a) tick each part of the nitrogen cycle (shown following) where water is needed by a living thing

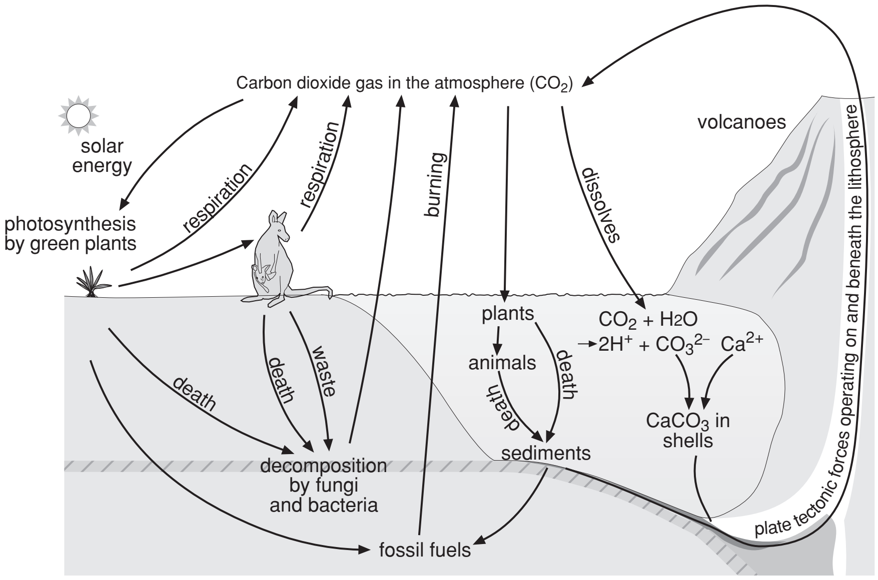
b) underline the name of any chemical that is transported by water.



The nitrogen cycle in nature.

The carbon cycle

Remember that water is usually at least 50% of every living thing.
Tick each step in this carbon cycle where water is needed by a living thing.

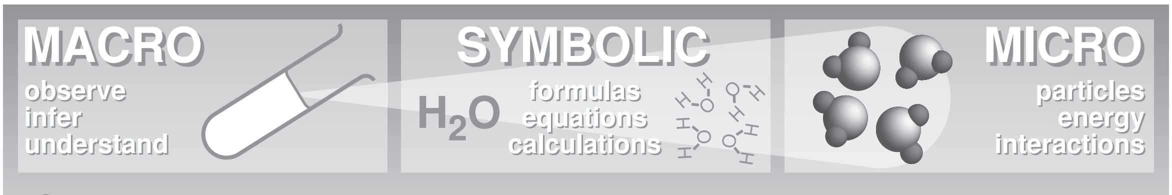


The carbon cycle in nature.

Water has a huge capacity for absorbing heat energy without much change in temperature. The water covering 70% of the Earth's surface moderates temperature changes and prevents temperature extremes. Just as Australians living on the coast experience less extreme temperatures compared with inland so living things in aquatic habitats experience less extremes of temperature than living things in **terrestrial habitats**.

Water is an agent of weathering of rocks both as liquid and solid. Liquid water causes chemical changes such as its reaction with the common mineral feldspar to form clay, salts and silica. The freezing of liquid water to solid water is a physical change producing tremendous pressures in cracks which can split rocks. Moving waters such as in glaciers and rivers carve out valleys. Rocks rolled over by rivers and waves and rocks carried along by ice at the bottom of glaciers wear down other rocks.

Water – nature's fluid, the most versatile solvent, the main transporter of chemicals between the non-living and living worlds – is the basis of life as we know it. To understand the fundamentals of biochemistry, the reactions of ions in lithosphere and hydrosphere, and how water transfers energy around the atmosphere you need to relate water's properties to its molecular structure – this is what you will study in Part 2.



Suggested answers

Distribution of water in the hydrosphere

1

Water location	State	Percentage of total water (%)	Percentage of freshwater (%)	Mean residence time
oceans	l	96.5	–	2600 years
ice caps and glaciers	s	1.8	69.6	1100 years
lithosphere (groundwater)	l	1.7	30.1	700 years
lakes	l	0.0013	0.2	3 years
soil	l	0.0012	0.05	155 days
atmosphere	s, l, g	0.0010	0.04	8 days
biosphere (living matter)	l, g	0.0001	–	3 days

2 a) solid state 1.8%

b) liquid state 98.2%

c) gaseous state $<$ 0.001%

3 a) 2.25 kg

b) 2.25 kg

c) (60/100 x 60) kg/2.25 kg per day = 16 days

4 a) 2200 L = 2200 kg = 2.200 t per person per day

b) 1560/2200 x 100/1 = 71%

c) 640/2200 x 100/1= 29%

Investigating the density of water

1. Volume of ice is greater in all three containers. Volume increase greatest for cold water, least for hot water.
2. Sample measurements:

   cold water 4 mm

   tap water 3 mm

   hot water 1.5 mm
3. Solid water (ice) forms a structure with large hexagonal spaces. Thus liquid water has a smaller volume than solid water.
4.
   - greatest volume of solid water is in the cold water container
   - smallest volume of solid water is in the hot water container.
5.
   - greatest mass of water must be in the container with the greatest volume of solid water, that is , the cold water container
   - smallest mass of water must be in the container with the smallest volume of solid water, that is, the hot water container.
6. The container of cold water had water at the greatest density (as it produced the greatest mass of ice and all containers had the same volume).
7. The container of hot water had water at the smallest density. The hot water container produced the smallest mass of ice. Thus this container had the smallest mass in the same volume as the other containers. The hot water had the smallest density but its density was still larger than the density of ice because when it froze the volume of ice was greater than the volume of the container.
8. Liquid water has highest density at low temperatures.
   Density decreases as the temperature increases.
9. Solid water takes up a bigger volume than liquid water. Thus the density of solid water is less than the density of liquid water.
10. Ice floats because solid water is less dense than liquid water.

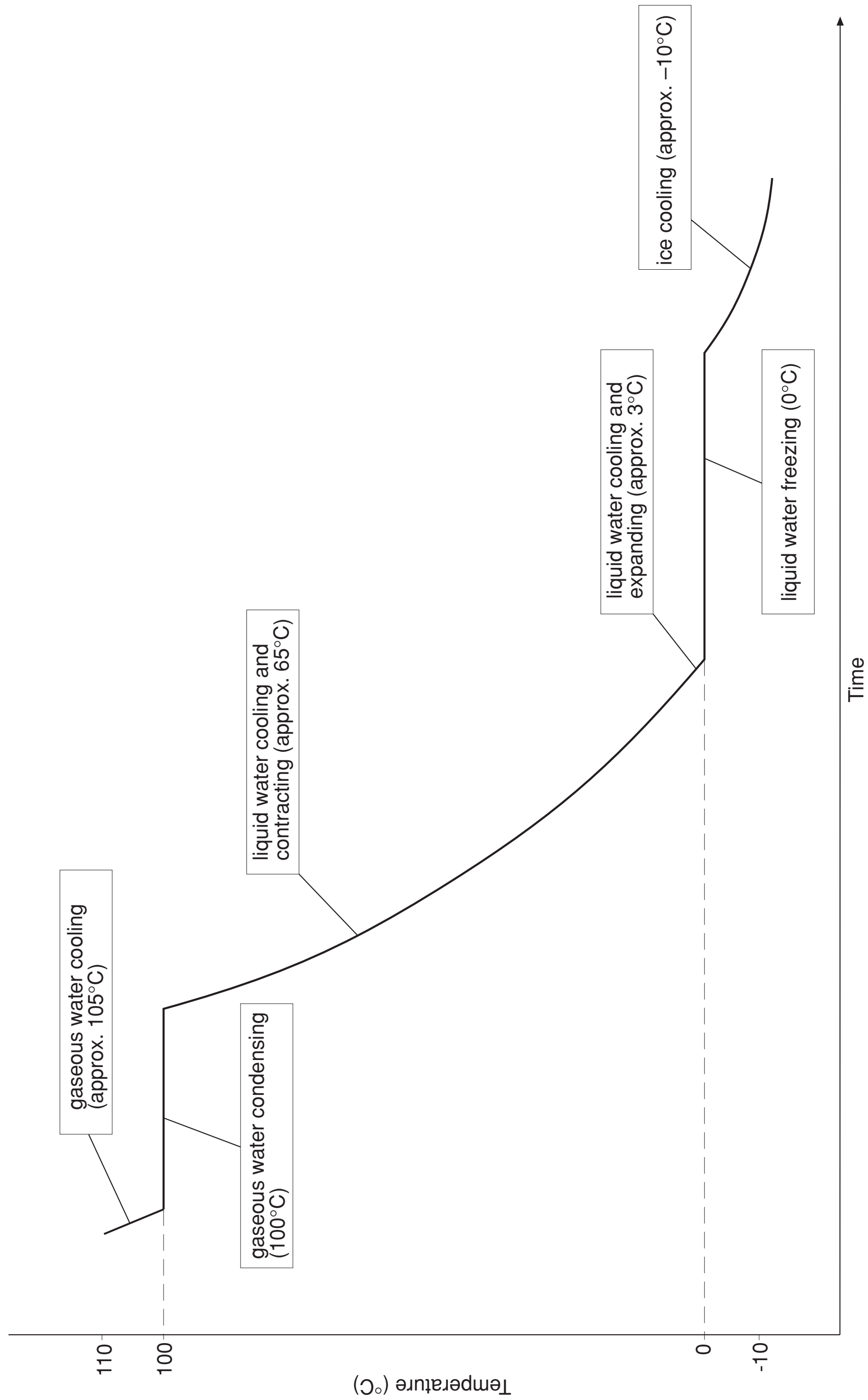
Density of water

1. The density of water increases from 15°C until 4°C where it is at maximum density. The density of liquid water decreases slightly from 4°C to 0°C. When liquid water freezes at 0°C to solid the density decreases significantly.
2. 4°C

- egg should sink in fresh water
- egg should float in salt water because egg density is between that of fresh water and salt water

- diagrams should show more of an iceberg above salt water than fresh water. Density of salt water > density of fresh water therefore an iceberg floats higher in salt water.

Water cooling curve



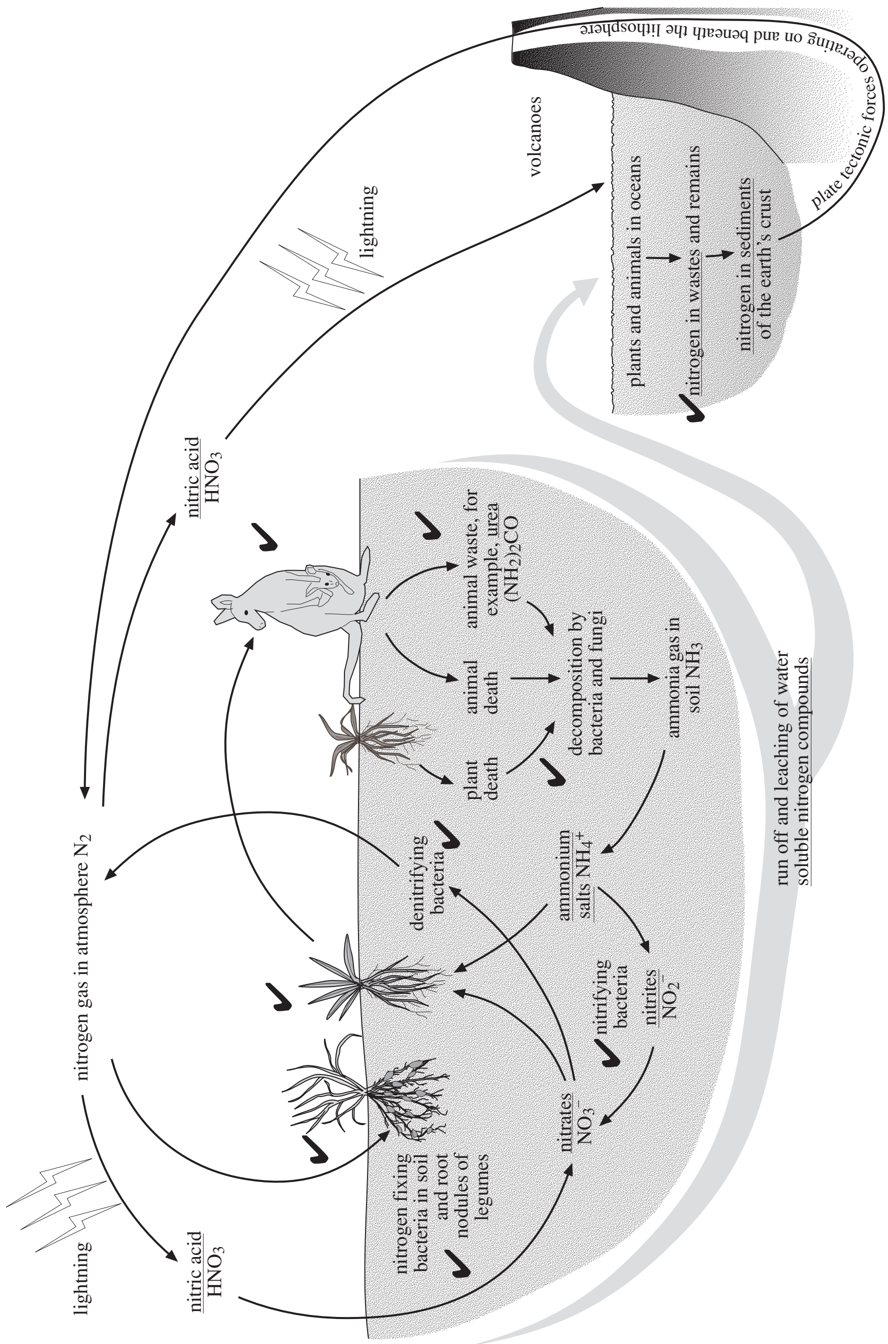
Biogeochemical cycles

a) meaning life or living things – *bio*

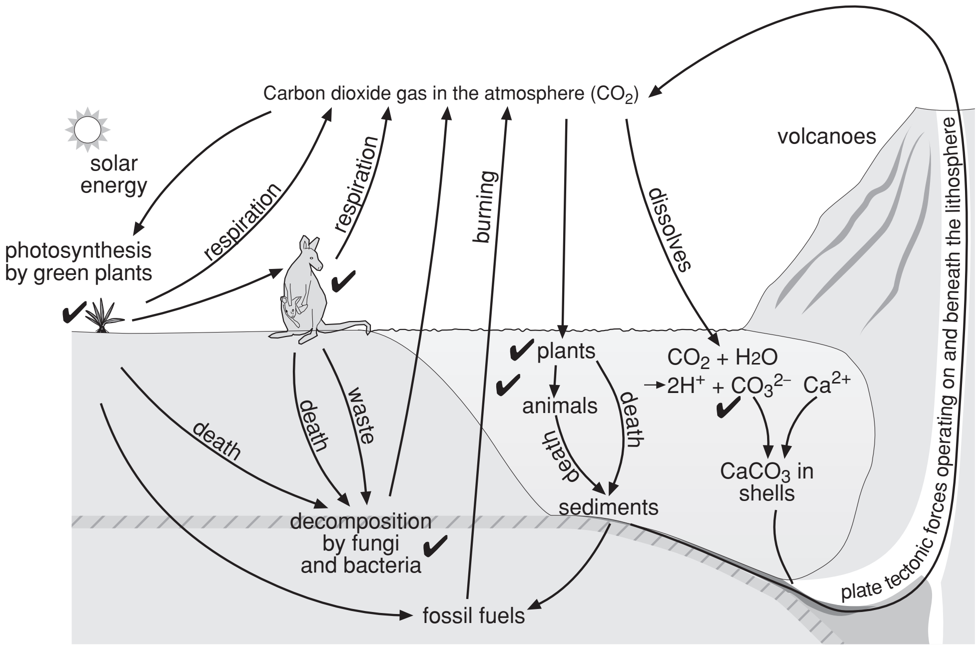
b) meaning the earth – *geo*

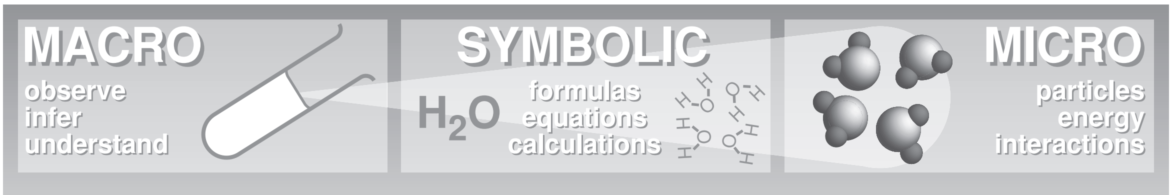
c) meaning substance used or produced by a chemical process – *chemical*

Nitrogen cycle



Carbon cycle





Exercises – Part 1

Exercises 1.1 to 1.2 Name: ________________________________

Exercise 1.1: A more quantitative approach

You will need:

- a thermometer measuring from –10°C to 110°C
- three straight drinking straws
- a fine marker pen
- three small, wide mouthed heat resistant containers
- three pieces of clay/Blu-tack®
- some food colouring solution
- a freezer or freezing compartment of a refrigerator
- half a cup of ice
- a way of boiling water
- a rule marked in millimetres (mm).

Method:

1 Add about a half cup of tap water to the half cup of ice.

2 Boil at least half a cup of water.

3 Mark the three containers *C* for cold, *T* for tap and *H* for hot water.

4 Place a piece of clay/Blu-tack® in the centre of the inside bottom of each container. Push the pieces down so they are about 1 mm thick.

5 Separately fill each container to near the top with:
   - cold water from the cup of ice
   - tap water from a tap
   - hot water from the boiled water.

6 Add about five drops of food colouring to each container and stir.

7 Slowly lower a straw vertically into each container until the bottom end is stuck in the clay/Blu-tack®.

8 Measure the temperature of the liquid in the cold water container. Record this temperature on the side of the container. Slowly pour the coloured water out of the container keeping you finger over the top of the straw so no liquid escapes from the straw. Mark the height of the water in the straw. Place the container in the freezer.

9 Repeat step 8 for the tap water, then for the hot water.

10 Allow time for all three lots of water in the straws to freeze.

Results:

1 The straws are uniform bore (constant area) so lengths of straw are proportional to volume. Measure length in mm from the clay/Blu-tack® surface to the mark (volume of liquid) and to the top of the ice (volume of solid). The numbers you get represent volume units. Calculate vol. of liquid/vol. of solid to two significant figures.

2 Remembering that density is inversely proportional to volume (that is, density decreases as volume increases) use the results in the third last and second last rows to calculate the density of each liquid.

3 Complete the last three columns of the following table:

Type of water	*sample*	ice water	tap water	hot water
Temperature (°C)	*80*			
Volume of liquid (mm vol. units)	*46*			
Volume of solid (mm vol. units)	*49*			
$\frac{\text{vol. of solid}}{\text{vol. of liquid}}$	*49/46*			
Density of solid (g cm^{-3})	*0.92*	0.92	0.92	0.92
Density of liquid (g cm^{-3})	*0.98*			

Exercise 1.2 Draft plan for my open-ended investigation

Chemistry

Preliminary Course

Stage 6

NEW SOUTH WALES DEPARTMENT OF EDUCATION AND TRAINING

TAFE NSW

OPEN TRAINING AND EDUCATION NETWORK - DISTANCE EDUCATION

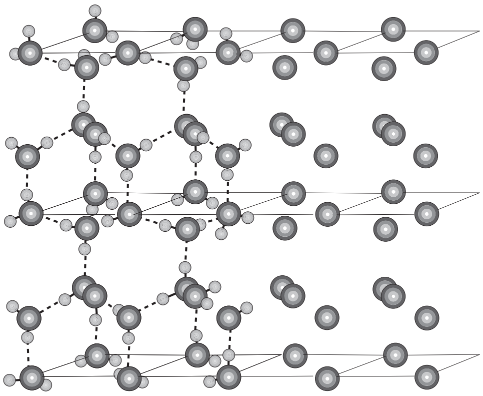
Water

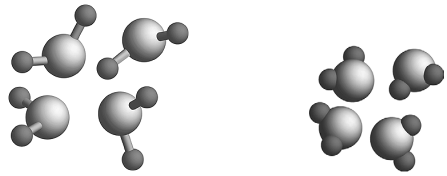
Part 2: Structure and properties

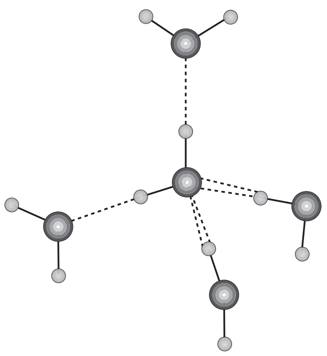
Incorporating October 2002

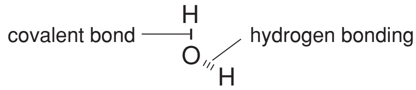
AMENDMENTS

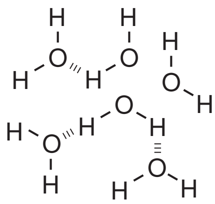
Structures representing water



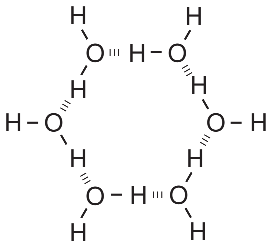




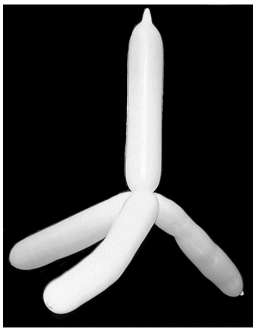


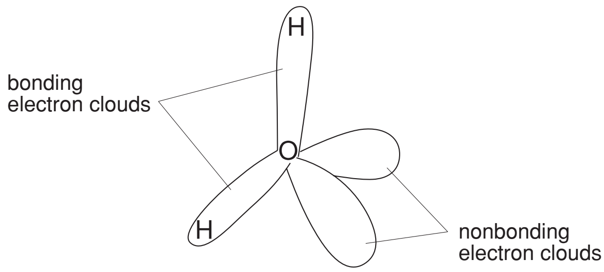


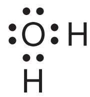
liquid water



solid water

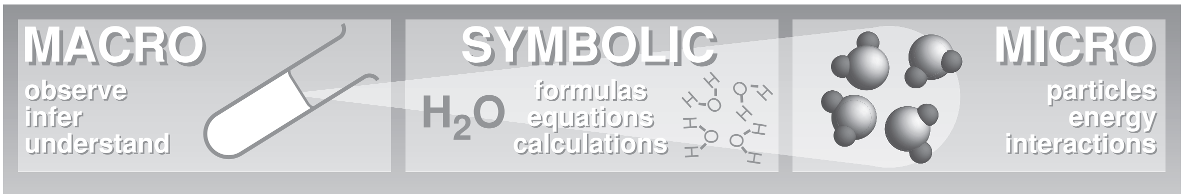




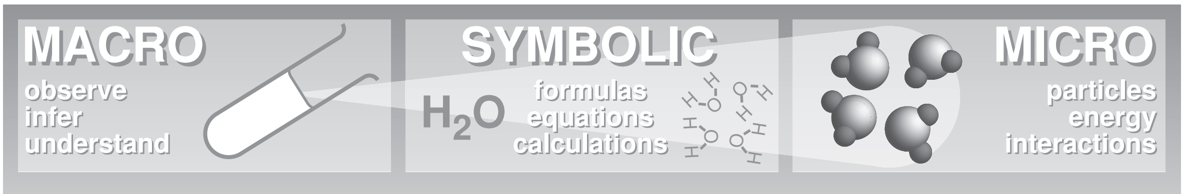


Properties of water

Molar mass 18.016 g/mol MP 0°C BP 100°C
Maximum density (at 3.98°C) 1.000 g cm^{-3}
Density at 25°C 0.997 g cm^{-3}
Density $H_2O(s)$ at 0°C 0.92 g cm^{-3} Density of $H_2O(l)$ at 0°C 1.00 g cm^{-3}
Viscosity at 25°C 0.9 x 10^{-3} N s m^{-2}
Surface tension at 25°C 0.072 J m^{-2}
Steam volume at 100°C = 1700 x liquid volume at 100°C


MACRO
observe
infer
understand
SYMBOLIC
H_2O
formulas
equations
calculations
MICRO
particles
energy
interactions

Contents



Introduction

The properties of water are so unusual for such a small molecule that it has been described as wacky, weird and even bizarre! Compared with other chemicals water is really quite extraordinary!!

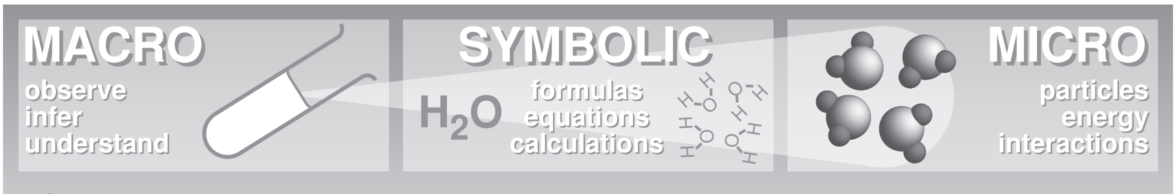
The uniqueness of this chemical – the chemical upon which all life as we know it is based – deserves careful study of its micro level structure. You will study water as the single molecule in the gas phase, groups of molecules (called clusters) in the liquid phase and molecules bonded together in the crystal lattice of the solid phase.

In Part 2 you will be given opportunities to learn to:

- construct Lewis electron dot structures of water, ammonia and hydrogen sulfide to identify the distribution of electrons
- compare the molecular structure of water, ammonia and hydrogen sulfide, the differences in their molecular shapes and in their melting and boiling points
- describe hydrogen bonding between molecules
- identify the water molecule as a polar molecule
- describe the attractive forces between polar molecules as dipole-dipole forces
- explain the following properties of water in terms of its intermolecular forces:
  - surface tension
  - viscosity
  - boiling and melting points

In Part 2 you will be given opportunities to:

- process information from secondary sources to graph and compare the boiling and melting points of water with other similar sized molecules
- identify data and process information from secondary sources to model the structure of the water molecule and effects of forces between water molecules
- choose equipment and perform first-hand investigations to demonstrate the following properties of water:
  - surface tension
  - viscosity



Comparing molecules

Comparing similar sized molecules

Most of the volume of a molecule is the space taken up by the electrons. The bonding between atoms within polyatomic molecules and the attractive forces between molecules involve electrons. To fairly compare water with other similar sized molecules, it is best to consider molecules with the same number of electrons (**isoelectronic** molecules).

Consider methane CH_4, ammonia NH_3, water H_2O, hydrogen fluoride HF and neon Ne. All of these are molecules (particles that can move independently of other particles) and all contain a total of ten electrons. They are isoelectronic.

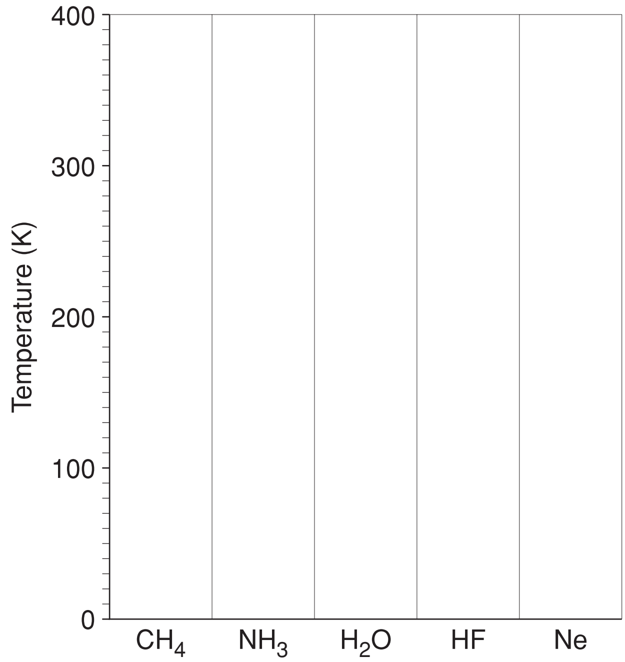
How do you know a molecule contains ten electrons?

Look at the atoms that make up the molecule eg. methane is made up of one C and four H. The number of electrons in an atom equals the atomic number of that atom eg. C has six electrons, each H has one electron. Total the number of electrons in the atoms making up the molecule
eg. CH_4 has 6 +(4 x1) = 10 electrons.

Use this type of mental calculation and a periodic table to check that NH_3, H_2O, HF and Ne all contain the same number of electrons. If this is the case you can make a fairer comparison of their properties. Whenever you wish to make comparisons, keep as many variables constant as possible.

The table below lists the MP and BP for each of these isoelectronic molecules in °C. Convert these readings to kelvin (K = °C + 273). Then plot the points on the following graph page. Draw straight lines between points so that you have a plot of the MPs and the BPs.

Molecule	CH_4	NH_3	H_2O	HF	Ne
MP (°C)	-182	-78	0	-83	-249
BP (°C)	-161	-33	100	20	-246
MP (K)					
BP (K)					



1 How does the MP and BP of water compare with similar sized molecules?

____

____

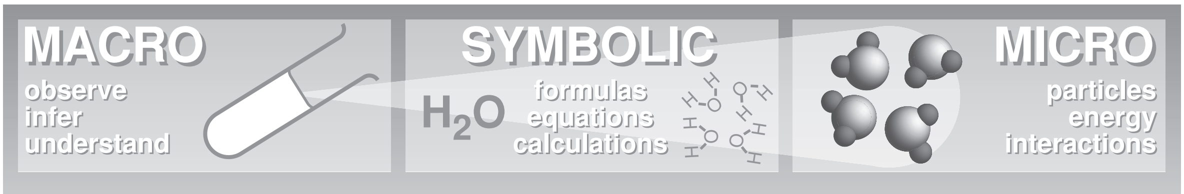
____

2 Do you think the difference in MP and BPs is due to the different number of atoms in the isoelectronic molecules?

____

____

Check your answers.



Lewis electron dot structures

Constructing Lewis electron dot structures

In Part 3 of *The chemical earth* module you learnt how to construct Lewis electron dot structures for H_2, Cl_2 and HCl. Here is how you can apply this method to constructing a Lewis diagram for H_2O.

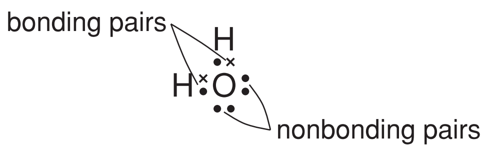
1 Choose a ‘central’ atom for the molecule – this is either the largest atom or a single atom of the molecule.

Choose O

2 Around the symbol for the ‘central’ atom draw the electrons of the valence (outer) shell. Draw a single electron for each hydrogen atom attached to the ‘central’ atom and pair up the remaining electrons.

O has electron configuration 2.6 so there are six electrons in the valence shell. Two of these are drawn as single electrons for the two H attached to the ‘central’ O and the remaining four are paired up into two pairs.

3 Place the remaining atoms around the ‘central’ atom. Use their valence electrons to pair up with the single electrons around the central atom. Each pair of electrons between the central atom and a surrounding atom represents a bonding pair and is a single covalent bond. Hydrogen has a combining power of one so there will be one covalent bond. That is one bonding pair of electrons, between each H and the central atom.



Use this procedure to draw Lewis electron dot structures below for CH_4, NH_3 and HF.

1 Use your electron dot diagrams to complete this table:

Molecule	CH_4	NH_3	H_2O	HF
Central atom	C	N	O	F
Combining power of central atom	4	3	2	1
Number of valence electrons around the central atom by itself	4		6	
Number of valence electrons around the central atom in the molecule		8	8	
Number of valence electron pairs around the central atom in the molecule			4	4
Number of bonding electron pairs around central atom in the molecule		3	2	
Number of nonbonding electron pairs around central atom in the molecule	0		2	

2 What is the connection between combining power of an element and the number of electrons in the valence shell of its atoms?

Sulfur is in the same periodic table group as oxygen and forms the hydride H_2S. Draw the Lewis electron dot structure for hydrogen sulfide and then complete the table below.

Molecule	H_2O	H_2S
Central atom	O	
Combining power of central atom	2	
Number of valence electrons around the central atom **by itself**	6	
Number of valence electrons around the central atom **in the molecule**	8	
Number of valence electron pairs around the central atom **in the molecule**	4	
Number of **bonding** electron pairs around the central atom **in the molecule**	2	
Number of **nonbonding** electron pairs around the central atom **in the molecule**	2	

3 Comment on the number of electrons around the central atom in each of these molecules:

Check your answers.

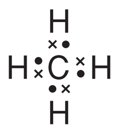
Electron clouds

You would have noticed in your electron dot diagrams of molecules that valence shell electrons are arranged in pairs. Each pair of electrons can be visualised as an **electron cloud**.

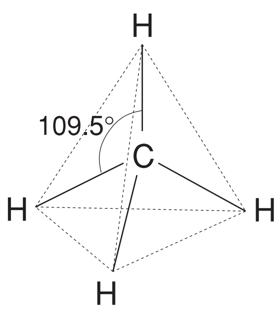
An electron cloud is a region in space where electron(s) are likely to be found. Electron clouds contain negative electrons and so electron clouds will repel one another because they have like charges. This model is called **VSEPR** for valence shell electron pair repulsion.

All of the polyatomic molecules you have just studied have four electron clouds (four pairs of electrons) around the central atom. These electron clouds repel one another and arrange themselves as far away as possible from one another in three dimensions.

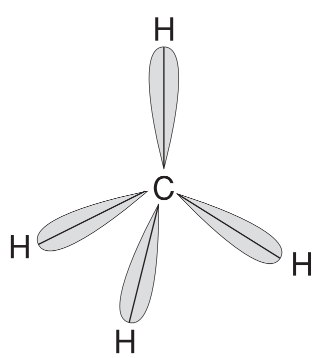
The most stable way for four equal electron clouds to arrange themselves is at 109.5° to one another. This is called a **tetrahedral shape** after the four sided triangular pyramid called the **tetrahedron**. If the central atom was at the centre of a tetrahedron each electron cloud would point towards a corner of the tetrahedron. This is the shape taken up by the four equal electron clouds in CH_4 shown below.



Lewis electron dot struture



tetrahedron



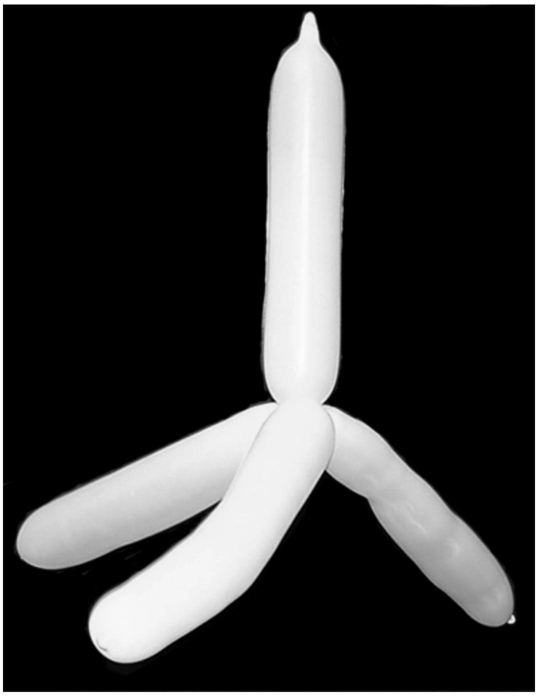
tetrahedral shape of four equal electron clouds

Modelling electron clouds

Modelling balloons can be used to represent electron clouds around central atoms in molecules. Modelling balloons are long thin balloons that can be twisted around one another to make up shapes. They can be purchased at many supermarkets in the party supplies section or in toy shops.

Select two thin modelling balloons. You will probably need to spend some time stretching them before they can be filled with air. Blow into the end while stretching the balloon to three or four times its unfilled length. Seal the end and squeeze the air into part of the balloon so that it swells up. Push the swollen part the full length of the balloon. Continue doing this until the unfilled balloon is about twice its original length and width.

Fill both balloons with air so they are about the same length. Hold a balloon in the centre with two hands and twist one half two or three turns so the balloon remains divided in the centre. Repeat this for the other balloon. Bring the centre parts of the two balloons together and twist the centres over one another. The four half balloons should arrange themselves tetrahedrally as far away from one another as possible.



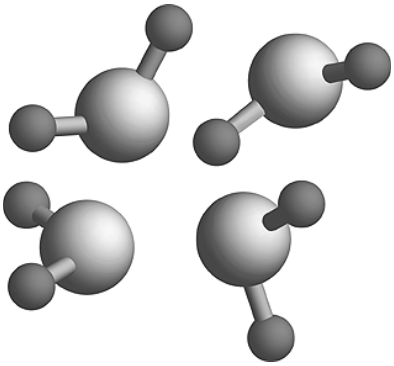
Each of the four electron pairs (two bonding and two nonbonding) around the central oxygen atom in water is represented by a half balloon.

You can distinguish between the bonding and nonbonding pairs in your model of water in two ways:

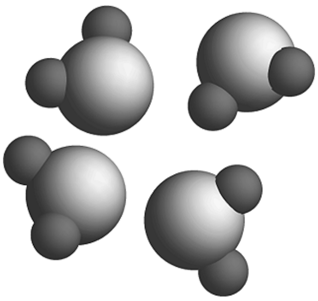
1 Mark the ends of two half balloons with two dots (representing the nonbonding electron pair). Mark the end of the other two half balloons with H to show the hydrogen nucleus at the end of a bonding pair. A marker pen or correction fluid could be used.

2 Use a thicker balloon or a different coloured balloon for the nonbonding pair electron cloud.

In Part 3 of the first module, *The chemical earth,* you saw a ball and stick model.

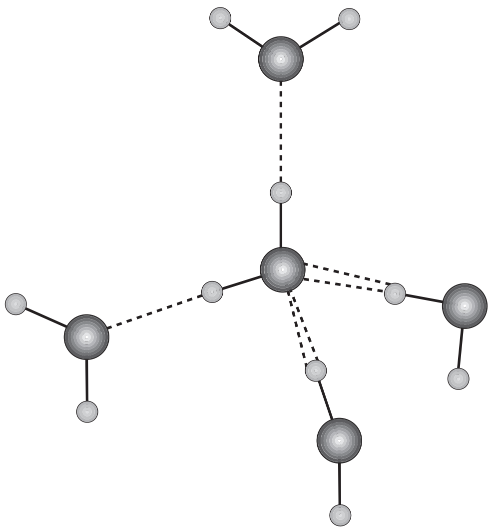


Then you saw a space-filled model for water.

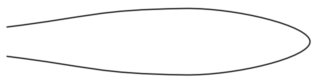


The tetrahedral model you have just made using modelling balloons is very useful for understanding how water molecules are arranged in the solid state.

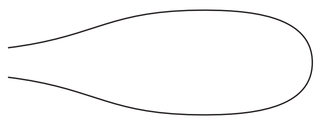
For most of the diagrams in this *Water* module the ball and stick model will be used. The sticks can show the tetrahedral direction of bonding between water molecules, especially in the solid state.



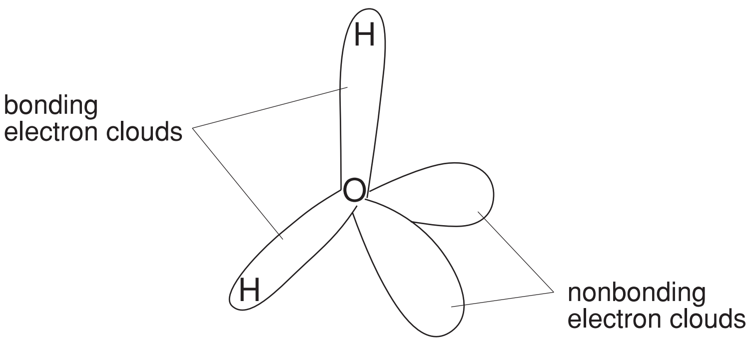
The electron cloud representing a bonding pair of electrons is narrow as it represents the covalent bond electron pair held between two nuclei.



The electron cloud representing a nonbonding pair of electrons is wider as it is not restrained between two nuclei.



The electron cloud model for water can be shown as:



If the nonbonding electron clouds are ignored it is easy to see why the water molecule is sometimes described as boomerang shaped!

The greater repulsion between the larger nonbonding electron clouds and the smaller repulsion between the smaller bonding electron clouds reduces the bond angle to 105° in water. The bond angle is the angle between the bonding electron clouds, that is between the two O–H bonds.

1 Complete this table to show the electron distribution in methane, ammonia and water.

Molecule	CH_4	NH_3	OH_2
Electron dot formula			
Number of electron pairs around central atom	4	4	4
Number of nonbonding electron pairs around central atom			
Number of bonding electron pairs around central atom			
Number of electron clouds around the central atom			
Bond angles expected around the central atom	109.5°	109.5°	109.5°
Actual bond angles	109.5°	107°	105°

2 Explain why the angle between the N–H bonds and between the O–H bonds are smaller than between the C–H bonds.

Check your answers.

1 Complete the table to show the electron distribution in water and hydrogen sulfide:

Molecule	H_2O	H_2S
Electron dot formula		
Number of electron pairs around central atom	4	4
Number of nonbonding electron pairs around central atom		
Number of bonding electron pairs around central atom		
Electron clouds around central atom		

2 Refer to the periodic table to explain the similarities between the structures of water and hydrogen sulfide:

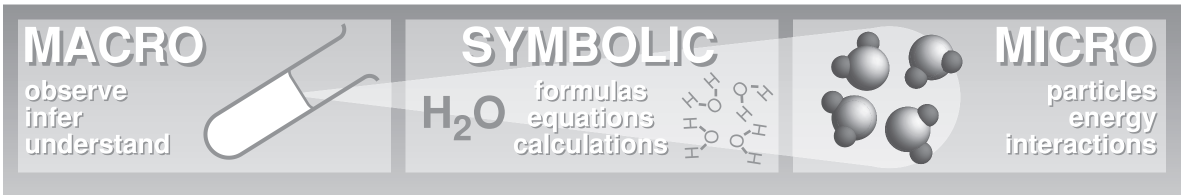
____

____

____

____

Check your answers.



Dipoles

Do water molecules have dipoles?

A dipole has two (di) poles or regions with opposite electric charges (+ and –). If a substance has molecules with dipoles (+ and – regions) the molecules will orientate to be attracted to charged objects.

What you will need:

- an object that can be electrically charged by rubbing against another object, for example:
  - a plastic rule/pencil/biro and a woollen jumper
  - a blown up balloon and a woollen carpet.

  Make sure the object and material used are dry.
- a tap where you can get a continuous thin stream of water.

What you will do:

1 Turn on the tap to get a continuous thin stream of water.

2 Charge the plastic rule/pencil/biro/blown up balloon by rubbing it against wool.

3 Bring the charged object from the side towards the thin stream of water; be careful that the charged object does not get wet.

4 Observe what happens to the stream of water.

5 Carry out steps 3 and 4 again but bring the charged object from the other side of the stream of water.

Results:

1 Describe what you observed.

____

____

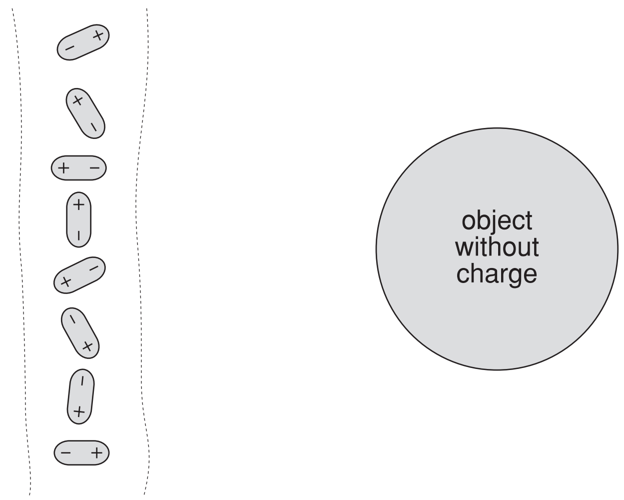
____

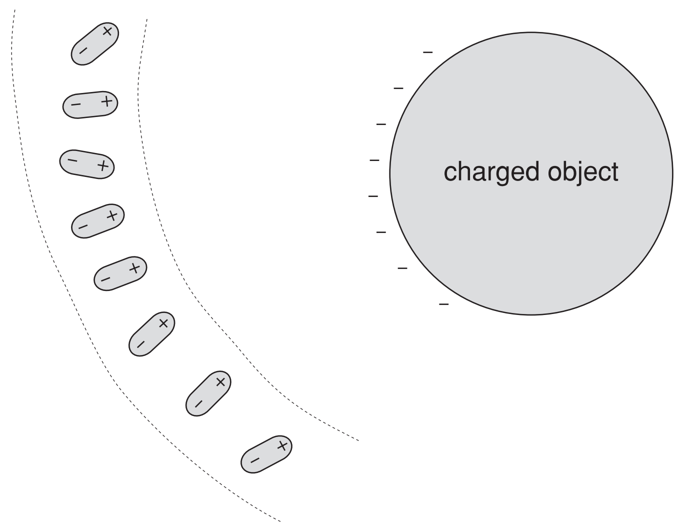
2 Do water molecules have dipoles?

____

____

Check your answers. The two diagrams below will help you understand what is happening at the micro level:

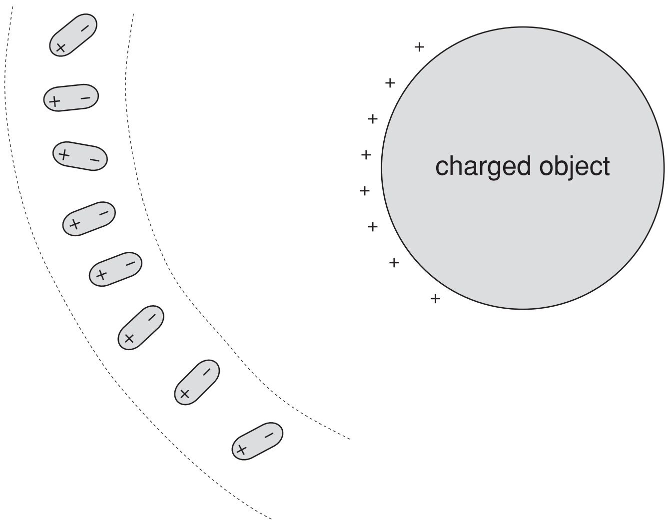




A molecule with a dipole is called a polar molecule because it will have + and – poles. The greater the dipole the more a thin stream of liquid polar molecules will be deflected by a charged object. A molecule without a dipole is called a nonpolar molecule. A thin stream of liquid nonpolar molecules will not be deflected by charged objects.

The previous diagrams show deflection of a stream of polar molecules by a negatively charged object. If a positively charged object had been used deflection would also have been towards the charged object.

Polar molecules are always attracted to a charged object no matter whether the charged object is positive or negative.



Unlike charges attract. Like charges repel.

Look at the previous two diagrams. Compare the distance between the attracted ends of the molecules and the charged object with the distance between the repulsed ends and the charged object. The smaller the distance between charged objects the greater the force.

In the diagrams when attraction forces are greater than repulsion forces the stream is attracted to the charged object.

Molecules

Polar
- dipole
- attracted to charged particles

Nonpolar
- no dipole
- not attracted to charged particles

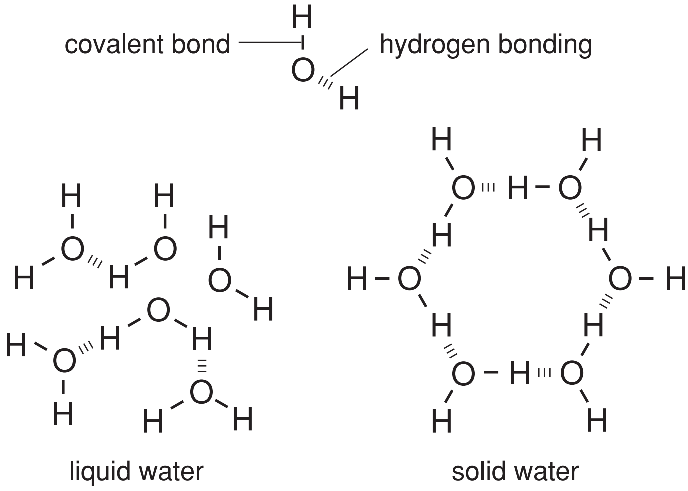
A polar molecule can be represented by

δ+ δ− OR + −

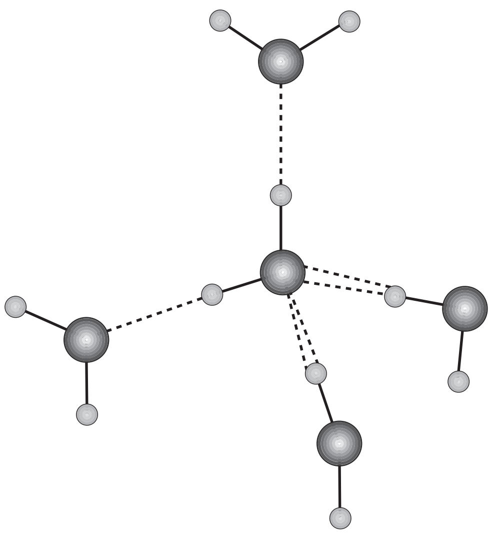
The δ+ and + represent a region of positive charge (less than the charge on a proton).

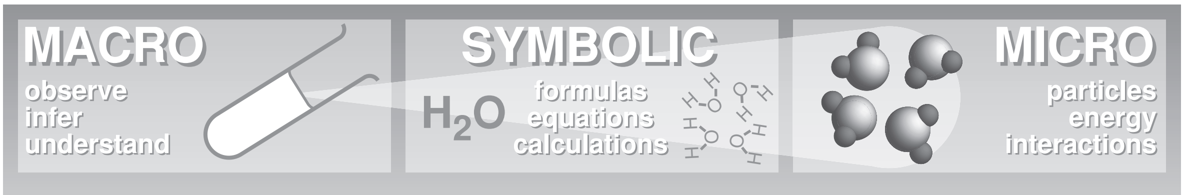
The δ– and – represent a region of negative charge (less than the charge on an electron).

The attraction of a δ+ region on one molecule for a δ– region on another molecule is called a dipole-dipole force. Dipole-dipole forces only occur between polar molecules. The dipole-dipole forces between the O of one water molecule and the H of another water molecule are so strong that they are given the special name **hydrogen bonding**.



Hydrogen bonding between water molecules is represented by - - - - in the diagram below. (The double dashed lines are to show the bottom centre water molecule is above the plane of the paper and the right hand water molecule is below the plane of the paper.)





Polar bonds

When two identical atoms are joined as in H_2, Cl_2, O_2, and N_2 the bonds between the identical atoms are nonpolar.

 Cl – Cl

 O = O

 N ≡ N

The two identical atoms have the same electronegativity (ability to attract electrons) so the electron pair(s)/electron cloud between the identical atoms are shared equally. The bond between the identical atoms is nonpolar.

The electronegativity values in the periodic table below enable you to predict how polar the bond is between two different atoms.

H 2.20																	
Li 0.98	Be 1.57						H 2.20	Symbol Electronegativity (EN)				B 2.04	C 2.55	N 3.04	O 3.44	F 3.98	
Na 0.93	Mg 1.31											Al 1.61	Si 1.90	P 2.19	S 2.58	Cl 3.16	
K 0.82	Ca 1.00	Sc 1.36	Ti 1.54	V 1.63	Cr 1.66	Mn 1.55	Fe 1.83	Co 1.88	Ni 1.91	Cu 1.90	Zn 1.65	Ga 1.81	Ge 2.01	As 2.18	Se 2.55	Br 2.96	
Rb 0.82	Sr 0.95	Y 1.22	Zr 1.33	Nb 1.6	Mo 2.16		Ru 2.2	Rh 2.28	Pd 2.20	Ag 1.93	Cd 1.69	In 1.78	Sn 1.96	Sb 2.05	Te 2.1	I 2.66	
Cs 0.79	Ba 0.89	LANTHANIDES	Hf 1.3	Ta 1.5	W 2.36		Os 2.2	Ir 2.20	Pt 2.28	Au 2.54	Hg 2.00	Tl 2.04	Pb 2.33	Bi 2.02	Po 2.0	At 2.2	
Fr 0.7	Ra 0.9	ACTINIDES															

La 1.10	Ce 1.12	Pr 1.13	Nd 1.14	Pm 1.2	Sm 1.17	Eu 1.2	Gd 1.20	Tb 1.2	Dy 1.22	Ho 1.23	Er 1.24	Tm 1.25	Yb 1.1	Lu 1.27
Ac 1.1	Th 1.3	Pa 1.5	U 1.38	Np 1.36	Pu 1.28	Am 1.3	Cm 1.3	Bk 1.3	Cf 1.3	Es 1.3	Fm 1.3	Md 1.3	No 1.3	

If the difference in electronegativities is greater than 1.7 the bond is usually regarded as ionic.

- The O–H difference of 1.24 indicates a very polar bond.
- The N–H difference of 0.84 indicates a polar bond.
- The C–H difference of 0.35 indicates very slight (negligible) polarity.

A polar double bond such as in C=O can be visualised as a single electron cloud which is more towards the electronegative element.

This produces a positive end δ+ and a negative end δ– C=O

δ+ δ–

As a result the CO molecule has a slight dipole. The dipole can also be represented as an arrow showing the direction of electron drift: C=O

$\rightarrow$

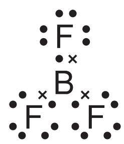
In the carbon dioxide molecule O=C=O there are two electron clouds around the central C atom. These regions of negative electron cloud arrange themselves as far away as possible from one another and so are pointing in opposite directions. The two equal polar bonds act in opposite directions. O=C=O

$\leftarrow \ \rightarrow$

Just as two equal and opposite forces cancel out so the two equal and opposite electron drifts cancel out so that CO_2 has no net dipole. CO_2 contains polar bonds but is a non polar molecule.

Boron trifluoride BF_3 has six electrons (three bonding pairs) around its central atom.

The three electron clouds around the boron atom arranged themselves as far away from one another as possible. They will lie in a plane at angles of 120° to one another. The three B–F bonds are very polar. Just as three equal forces at 120° to one another in a plane will cancel out, the three very polar B–F bonds cancel out. Boron trifluoride, BF_3 , is a nonpolar molecule.



Modelling the electron clouds

Take three equal sized modelling balloons. Fill two of them to the same size and then twist their ends over one another.

1 Did the two balloons arrange themselves the same way as two electron clouds around a central atom? Describe the arrangement you made for CO_2.

____

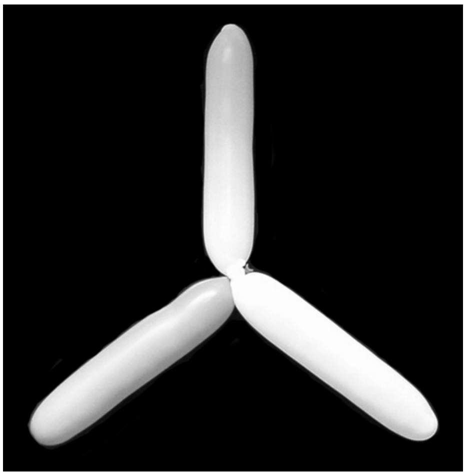
____

Fill the third modelling balloon to the same size as the other two balloons. Twist the end of the third balloon around the centre of the two balloons.

2 Did three balloons arrange themselves the same way as three electron clouds would around the central atom? Describe the arrangement you made.

____

____



3 Does this represent the electron clouds in CO_2 or BF_3? _________

4 Draw the electron dot formula for carbon tetrachloride (tetrachloromethane) CCl_4.

The four C–Cl bonds are polar and tetrahedrally arranged at 109° to one another. Just as four equal forces arranged tetrahedrally cancel out so these four polar C–Cl bonds cancel out. Carbon tetrachloride is a nonpolar molecule.

Check your answers.

Polar bonds in ammonia

1 Draw the electron dot formula of ammonia NH_3.

2 How many electron clouds are around the central N atom?

3 How many of these electron clouds are involved in bonding to H atoms? ____________________

4 Use the periodic table of electronegativities to explain why the N–H bond is polar.____________________

____

____

5 Explain why the three polar N–H bonds at about 107° to one another do not cancel. ____________________

____

____

____

6 Is the ammonia molecule polar or nonpolar? Explain your answer.

____

____

____

____

Check your answers.

Polar bonds in water

1 Draw the electron dot formula of water H_2O.

2 How many electron clouds are around the central O atom?

3 How many of these electron clouds are involved in bonding to H atoms?

4 Use the periodic table of electronegativities to explain why the O–H bond is polar. ____

____

____

5 Explain why the two polar O–H bonds at about 105° to one another do not cancel. ____

____

____

____

6 Is the water molecule polar or nonpolar? Explain your answer.

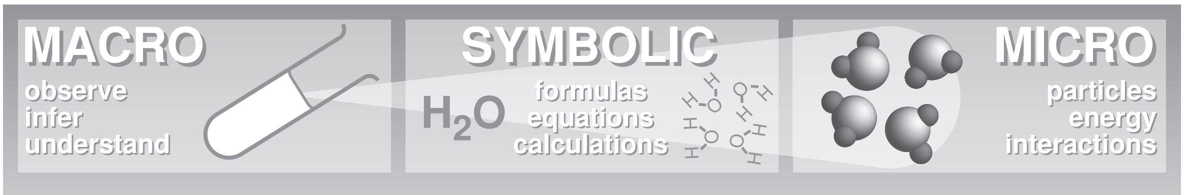
____

____

____

____

Check your answers.



Keys and Venn diagrams

A **key** is a table or diagram used to classify a group into smaller groups. Most keys are **dichotomous** dividing a group into two smaller groups with different characteristics.

Here are two keys you used in the first module *The chemical earth.*

Compounds

- Covalent
  - nonconductors of electricity when liquid or dissolved in water
  - usually gas, solid or low melting point melting point solid.
- Ionic
  - conductors of electricity when liquid or dissolved in water
  - usually high melting point solid.

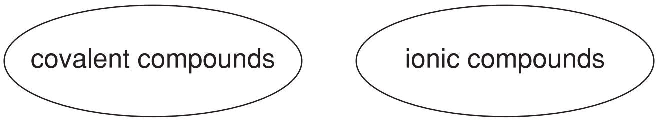
Chemicals

- mixtures ········ separated using physical properties ········▸ pure substances
  - mixtures: *properties vary*
  - pure substances: *properties fixed*
    - compounds ········ separated using chemical properties ········▸ elements
      - compounds: *different types of atoms combined*
      - elements: *one type of atom*

Combine these two dichotomous keys into one larger key. Do not list the characteristics – just use the headings for each group.

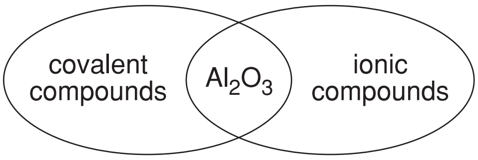
Check your answer.

Another useful way of showing connections between different sets of chemicals is to use a **Venn diagram.** A Venn diagram uses circles or ellipses.



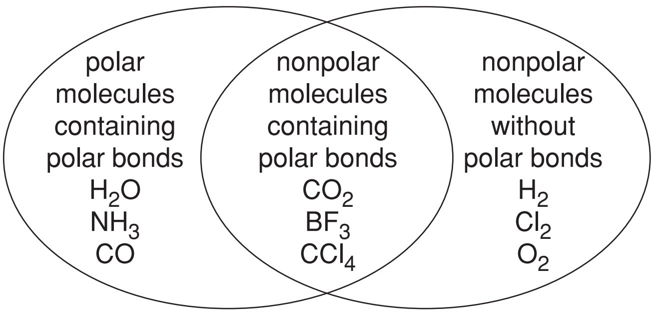
An advantage of a Venn diagram over a dichotomous key is that it can show where overlap occurs. For example in binary compounds where the electronegativity difference is about 1.7 the compound shows characteristics of covalent and ionic compounds. The conductivity of the compound is intermediate between that of covalent and ionic compounds. The bonds in these compounds are intermediate between covalent and ionic bonds.

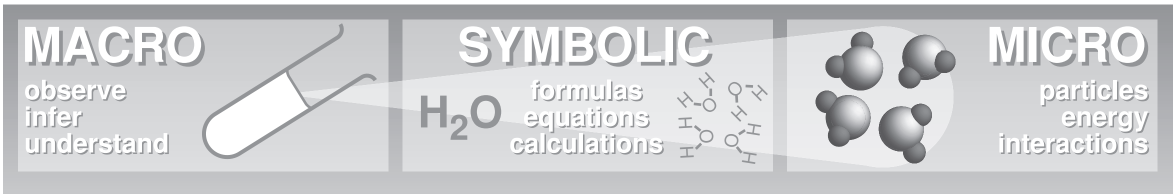
Aluminium oxide, Al_2O_3, with an electronegativity difference of 1.83 is an example. The high MP and BP and hardness are characteristic of continuous covalent compounds while the conductivity used in the electrolysis of Al_2O_3 to make aluminium is characteristic of ionic compounds.



In the previous section you saw how the shape and symmetry of a molecule that contains polar bonds could result in a nonpolar molecule. The polar bonds could cancel out eg. the two equal and opposite polar bonds at 180° in O=C=O, the three equal polar bonds arranged at 120° to one another in BF_3, the four equal polar bonds arranged tetrahedrally at 109.5° in CCl_4.

This can be shown as a Venn diagram:





Intermolecular forces

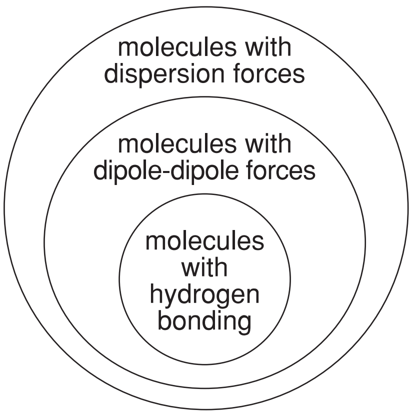
There are three main types of intermolecular forces:

- dispersion forces which occur between all molecules
- dipole-dipole forces between molecules with dipoles
- hydrogen bonding between molecules containing hydrogen covalently bonded to O, N or F.

Draw a trichotomous key diagram to show the three types of intermolecular force.

Check your answer.

This information can also be part of a Venn diagram.

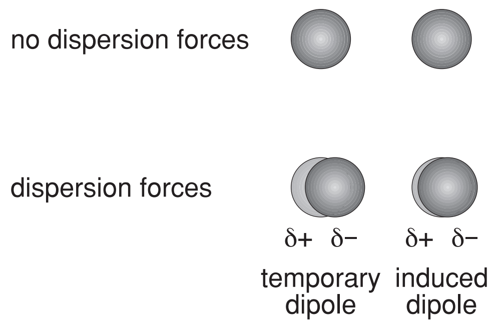


Advantages of the Venn diagram over the trichotomous key diagram are that it shows

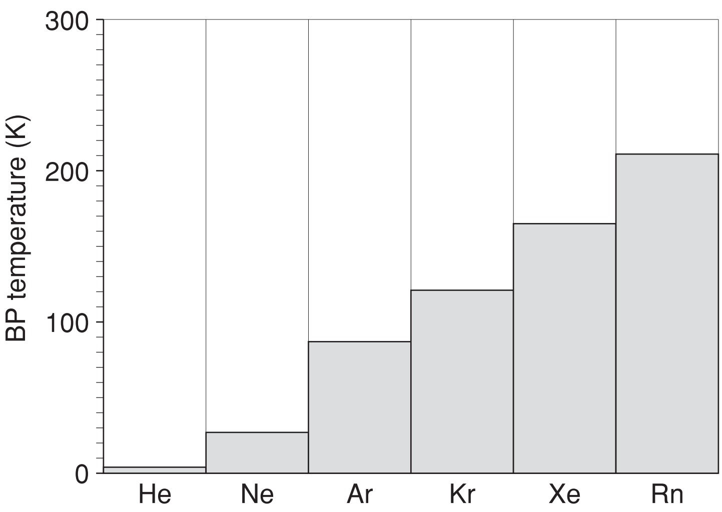
- all molecules have dispersion forces between them
- hydrogen bonding is between only some molecules with dipoles.

Dispersion forces

Dispersion forces are caused by movement of electron clouds about nuclei. The temporary distortion of an electron cloud, for example by collision with another molecule, produces a temporary dipole. One end of the molecule becomes temporarily positive and the other end temporarily negative. The temporary dipole can cause induced dipoles in nearby molecules.



Consider the noble gases. The BP of the noble gases He, Ne, Ar, Kr, Xe and Rn in K are shown in the graph below.



At 0 K, absolute zero, molecules cease moving and have no energy of motion. Liquid helium must be heated to 4 K before the molecules separate into the gas phase. This shows that there is a small force

keeping He molecules close together in the liquid phase until 4 K is reached. This small force is a dispersion force.

Notice that the dispersion force increases with the number of electrons. The increase in BP as the molecular mass of the noble gas increases is due to the greater number of electrons in the molecules.

Dispersion forces are the only type of force between:

- noble gas molecules
- non-polar molecules such as halogens (F_2, Cl_2, Br_2 and I_2) and most alkanes C_nH_{2n+2} (CH_4, C_2H_6, etc.).

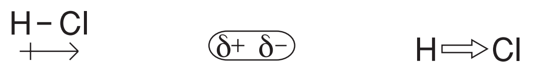
Always remember that dispersion forces occur between all molecules and the larger the molecule the stronger the dispersion forces.
Do not be mislead by the term dispersion which usually means scattering–dispersion forces bring molecules closer together.

Use dispersion forces to explain why the BPs of the halogens increase in the order F_2, Cl_2, Br_2, I_2.

Check your answer.

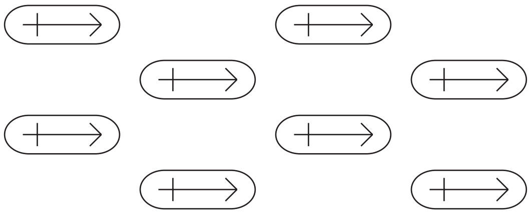
Dipole-dipole forces

A polar molecule has a slightly positive end δ+ and a slightly negative end δ–. The polar molecule is called a **dipole**. The strength of a dipole can be measured as a **dipole moment**. Here are different ways of representing the dipole in hydrogen chloride:

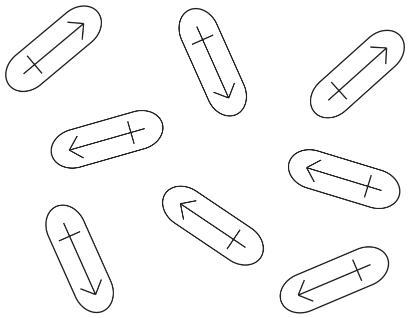


arrow with cross bar | δ+ and δ– | large arrow

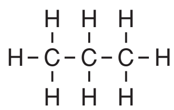
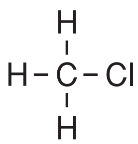
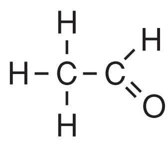
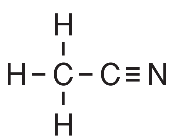
In the solid phase the dipoles are arranged in a regular pattern.



In the liquid phase, although the polar molecules are moving around, the dipoles affect the orientation of surrounding molecules.



Compare these polar molecules which have similar molecular masses. Molecules with similar molecular masses have similar dispersion forces.

Name	**Molecular formula**	**Structural formula**	**Molecular mass (g)**	**BP (kelvin)**	**Dipole moment (10^{-30}Cm)**
propane	$CH_3CH_2CH_3$		44.1	231	0.3
methyl chloride	CH_3Cl		50.5	249	6.2
ethanal	CH_3CHO		44.1	294	9.0
ethanenitrile	CH_3CN		41.1	355	13.1

1 Describe the connection between the dipole moment and the BP:

a) in words

b) by means of a labelled graph

2 The very small dipole for propane indicates that most of the intermolecular forces in propane are ____________ forces.

As the molecules become more polar as shown by the increasing ______ moment values, dipole-dipole forces between the polar molecules become more important.

___ increases as the dipole-dipole forces increase.

Check your answers.

Hydrogen bonding

A very strong type of intermolecular force occurs in molecules that have a hydrogen atom bonded to a small, highly electronegative atom with nonbonding pairs. This type of bonding is called hydrogen bonding and occurs when hydrogen is bonded to nitrogen, oxygen or fluorine.

The hydrogen bonding is represented as ….. between the H in one molecule and the N, O or F in another nearby molecule

H–F…..H–F   H–O–H…..OH_2 H–N–H…..NH_3
.
                                                          |
                                                          H

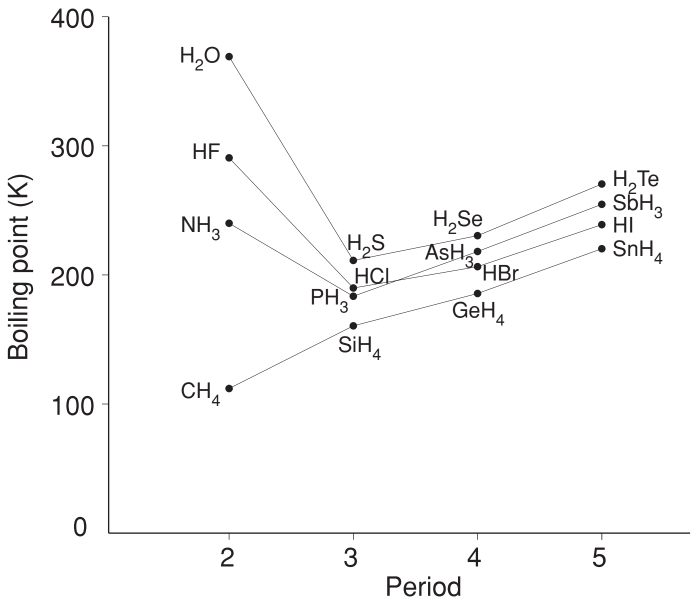
Study this graph showing the boiling point of hydrides of Group 14, 15, 16 and 17 elements. Answer the questions that follow.

1 a) For each group, which of the lines represents tetrahedral non-polar substances?

____

b) Only one type of intermolecular force keeps non-polar molecules together in the liquid phase. Name this type of force.

____



2 The other three group lines represent twelve polar substances. Nine of these twelve substances have dispersion and dipole-dipole forces between their molecules. The other three substances in these group lines have abnormally high BPs because of the presence of hydrogen bonding as well as dispersion and dipole-dipole forces. Give the formulas of the three substances with hydrogen bonding.

____

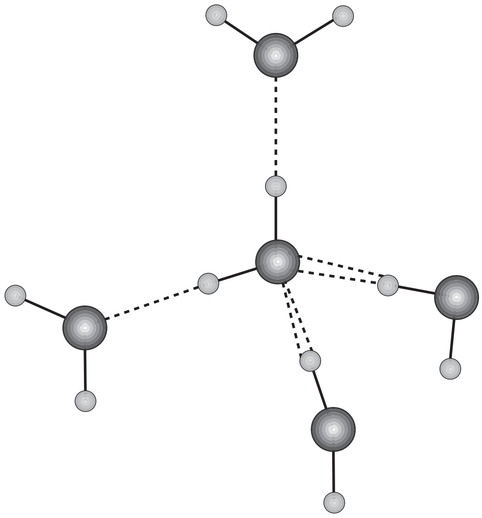
Check your answers.

The abnormally high MP and BP of water can be explained by the large amount of energy required to break the hydrogen bonds between water molecule. In solid water (ice):

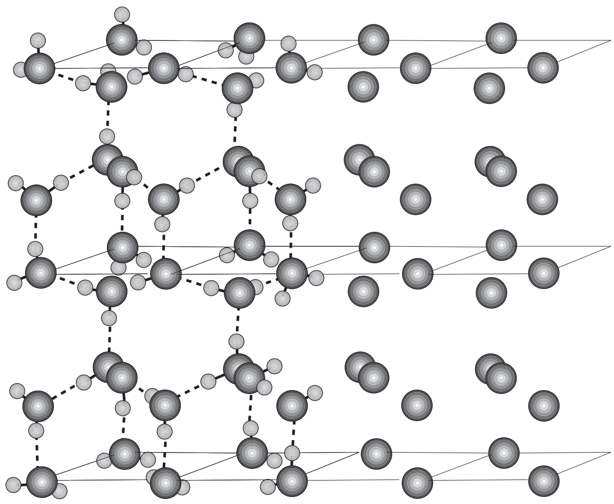
- each of the two H in a H_2O molecule can be hydrogen bonded to an O in other water molecules
- the single O in a H_2O molecule can be bonded to two H in two other separate water molecules.

Thus, each water molecule is bonded to four other water molecules. Each water molecule has four other water molecules arranged tetrahedrally around its O atom. As liquid water is cooled close to its freezing point, the oppositely charged poles of neighbouring molecules align themselves more and more. For hydrogen bonding to occur tetrahedrally the molecules must spread out. Thus, molecules are more widely spaced in solid water than in liquid water.

This tetrahedral arrangement produces the very unusual situation of a chemical with its solid phase less dense than the liquid phase. No other common chemical has a solid phase that floats on top of the liquid phase.



This tetrahedral arrangement around each H_2O gives an open hexagonal shaped crystal structure in ice. The RHS of the diagram below shows positioning of the O only; the LHS shows the H_2O positioning.

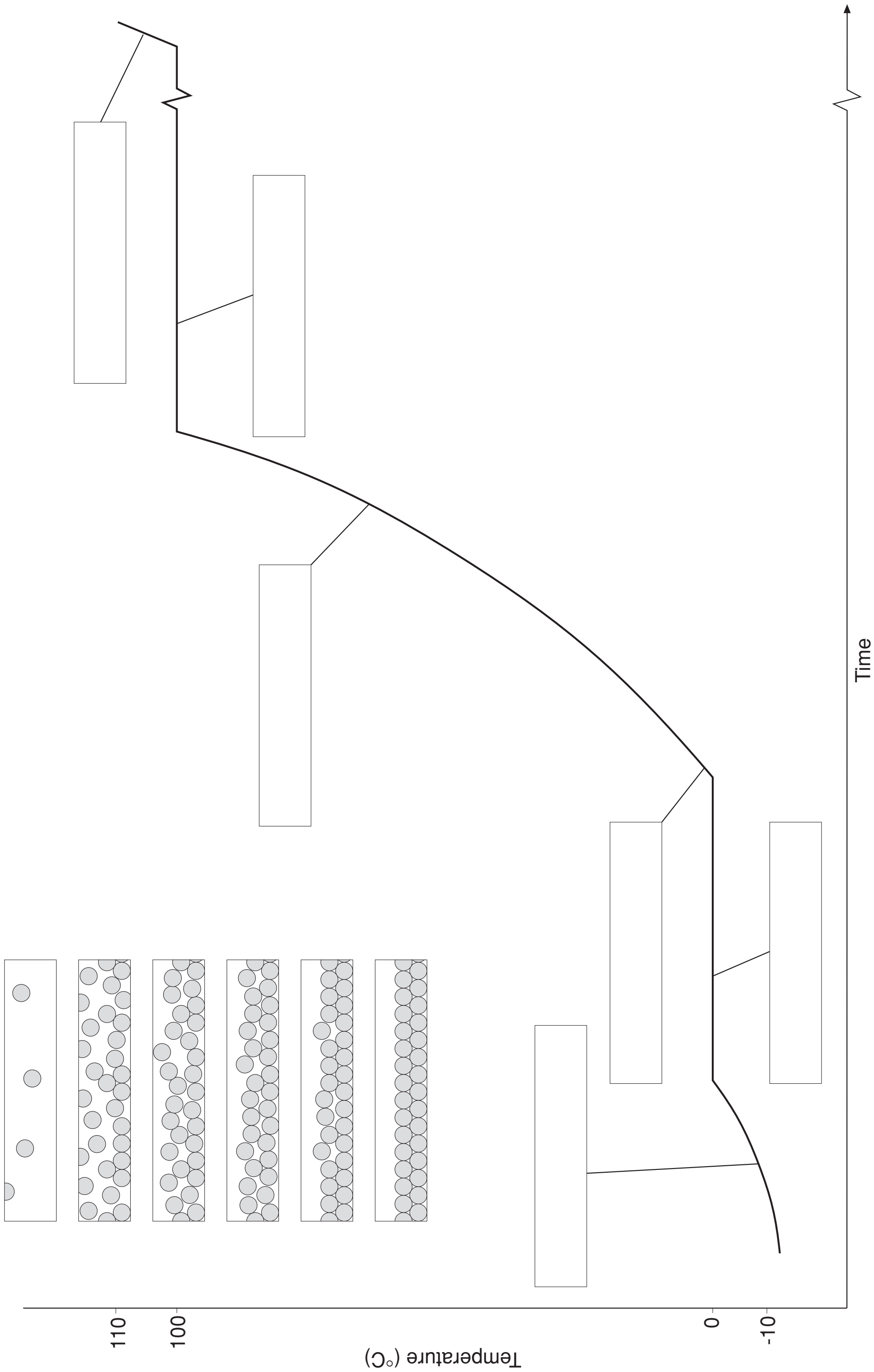


The large spaces inside the hexagonal structure of the solid show why the volume of water increases (and density decreases) when liquid water freezes. When ice melts at 0°C the tetrahedral arrangement breaks down and the loose molecules fill the open hexagonal spaces so that the volume decreases and density increases.

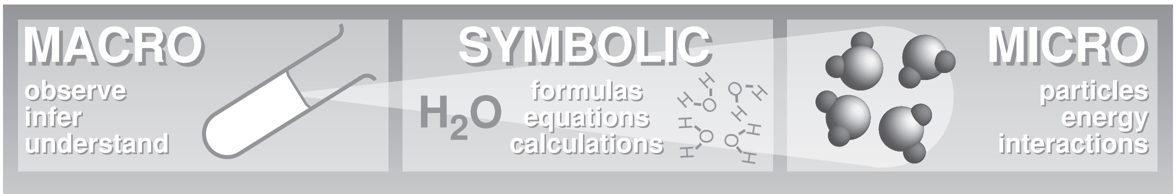
As the temperature rises the density increases to a maximum of 1.0000 g cm^{-3} at 4°C. Above 4°C the greater motion of the molecules starts to take up more space and the density decreases as the temperature rises.

At low temperatures most liquid H_2O molecules are part of a small cluster of water molecules associated together by hydrogen bonding. As the temperature rises the average size of these clusters decreases and more and more H_2O molecules move independently taking up more space.

The diagram following shows a water heating curve. Use the labels to show the appearance of H_2O molecules in six parts of the curve.



Check your answer.



Investigating properties of water

In this section you will choose equipment to perform first hand investigations to demonstrate the following properties of water:

- **surface tension** – cohesion in the surface layer of liquid
- **viscosity**.

You will be given suggestions on equipment that you could use. You may think of an alternative way of doing the activity. If so, make sure you thoroughly think through the activity, always keeping safety a high priority before carrying out the activity.

Surface tension

Surface tension results from the attractive forces between the molecules in the surface layer of a liquid.

- Take a clean dry paint brush and hold it in air. Observe how the hairs stick out like your own hair at the start of a bad hair day.
- Put the brush into a container of water. Observe how the hairs still stick out like human hair does under water.
- Lift the brush out of water. Observe how the hairs of the wet brush in air do not stick out – they stick together. The water surface tries to minimise its contact with the air and so draws the brush hairs together.

1 You go to the bathroom in the morning and notice your hair is all over the place. How could you use water to stop a bad hair day?

____

Water molecules minimise their contact with other molecules such as air molecules by coming close together.

2 Try to float a paper clip, dull razor blade or needle on water. With care you could succeed in floating a small steel object with a density of 8 g cm^{-3} on top of a liquid with density 1 g cm^{-3}!

Look at the floating object from the side. Does the surface of the water behave like a skin? A skin is a surface layer that does not break easily.

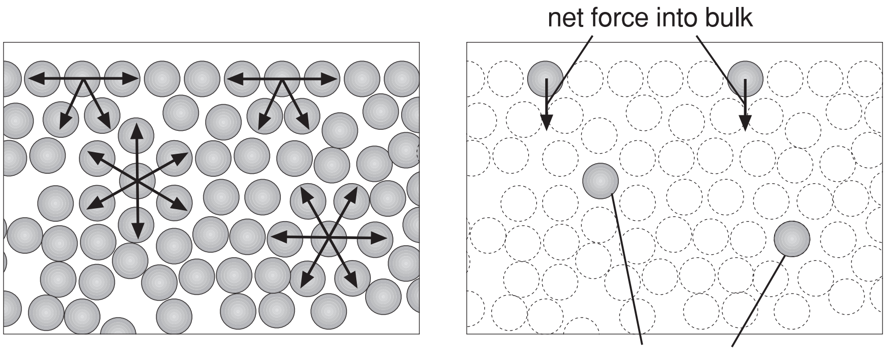
What happens if you push one end of the floating object under the skin of the water?

---

3 Have you observed water drops on the surface of flower petals, waxy leaves or a newly waxed car? Draw the appearance of a water drop on a waxy surface.

Check your answers.

The explanation for all these observations can be based on this simple diagram showing molecules in bulk liquid and in a surface layer.

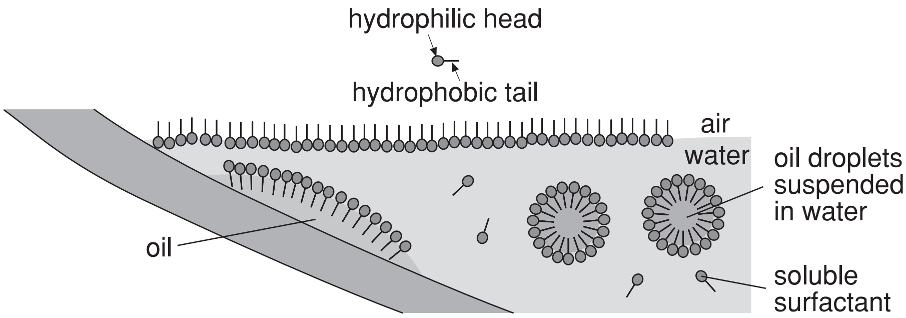


The molecules in the bulk liquid are attracted in all directions by surrounding molecules. However the molecules in the surface layer are attracted inwards towards the bulk liquid molecules. This causes a liquid surface to have the smallest possible area and so drops of liquid are spherical. The surface layer or skin is difficult to break so steel paper clips, razor blades and needles can be supported by the water skin and not sink in water.

The surface tension is the energy required to increase the surface area and is measured in joules per square metre, J/m^2. The stronger the force of attraction between the molecules the greater the surface tension. Water has a high surface tension because of its ability to form many hydrogen bonds.

The surface tension of water can be reduced by adding **surfactants** (surface-active agents) such as soap or detergent. Soap or detergent particles have a polar **hydrophilic** (water loving) end and a long thin **hydrophobic** (water fearing) chain. Surfactants collect at **interfaces** – the water/air surface and water/oil surface. The surfactants decrease the surface tension of water by disrupting the hydrogen bonding in the surface layer.

- Fill two clean identical containers with water.
- Sprinkle some water insoluble powder (eg baby powder, talcum powder or corn flour) on to the surface of the water.
- Add one drop of oil (eg cooking oil, lubricating oil) to one surface and note how much it spreads.
- Add one drop of surfactant (eg liquid detergent) to the other surface and compare.
- Do you agree that surfactant is a surface active agent?



Detergent and water in an oily dish.

You can observe detergent reducing the surface tension of water using three toothpicks, a clean container of water and detergent.

1. Place two toothpicks side by side on the surface in the centre of the water
2. Dip the tip of the third toothpick into detergent
3. Touch the detergent tip to the water surface between the floating toothpicks.

The detergent weakens the surface between the toothpicks. The tension or pulling force of the water outside the toothpicks is no longer balanced by the tension between the toothpicks.

Surfactants are added to **herbicide** spray solutions to lower the surface tension. This enables the solution to wet leaves and to enter the holes in the leaves. Surfactants produced inside human lungs help keep surfaces moist. Premature babies born before these surfactants have developed in their lungs are difficult to keep alive.

Viscosity

The viscosity of a liquid is a measure of its resistance to flow.

When a liquid flows the molecules slide over one another. If the molecules have strong intermolecular attractions and are long and thin and easily tangle the liquid will have a high viscosity eg tar, honey.
If the molecules have weak intermolecular attractions and are smaller the liquid has a low viscosity and flows easily eg. petrol.

The viscosity decreases as temperature rises. If you put some cooking oil into a pan it spreads slowly. When the pan is heated the cooking oil's viscosity decreases and it spreads quickly to form a thin layer.
What happens to the viscosity of honey that has been left in the refrigerator?

Relative viscosity is the viscosity of a liquid relative to another liquid such as water. Measurements of time for liquids to flow through a narrow tube are used to calculate relative viscosity.

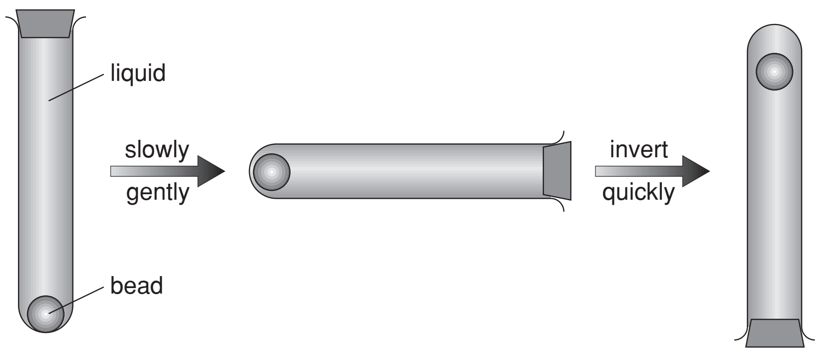
$$\text{relative viscosity of liquid} = \frac{\text{time for liquid flow}}{\text{time for water flow}}$$

Because viscosity varies with temperature all measurements must be carried out at the same temperature. For the measurements to be valid everything should be kept the same for the liquid and water.
Reliability is improved by making more than one measurement for each liquid and averaging the times.

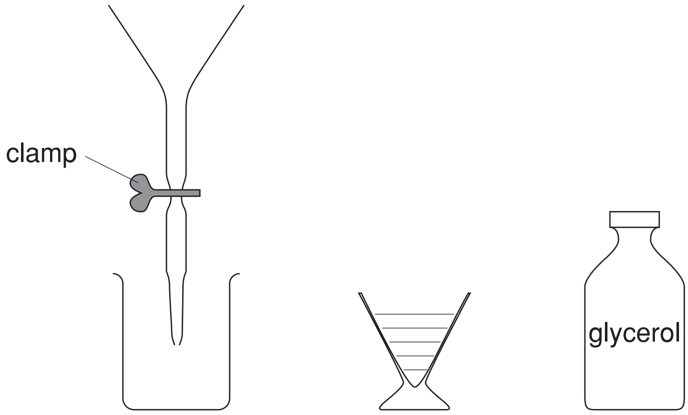
Relative viscosity of glycerol

Using simple apparatus as shown in Set A or Set B shown below.

Set A: Measuring time for a high density bead/ball to fall through liquid.



Set B: Measuring time for liquid to flow through a narrow tube.
Use a funnel, rubber or plastic tube, a clamp and the tip of an eyedropper.



Set out the steps you would carry out to measure the relative viscosity of glycerol compared with water.

After you have discussed your steps with another person carry out this activity. Then report your method, results, calculations and conclusions in Exercise 2.1.

Motor engine and gear oils have different viscosities. Special additives are added to a motor oil so that as the temperature rises the viscosity does not get too low. The additive molecules change from compact spheres to long thin molecules as the temperature rises. This shape change increases the surface area through which dispersion forces can act. Thus it is more difficult for the molecules to move past one another. This increase in viscosity counteracts any decrease due to raised temperature.

Capillarity and living things

Capillarity or capillary action is the ability of a liquid to rise up a narrow tube against the force of gravity. Capillarity occurs when the forces between the liquid and the tube walls are greater than the intermolecular forces within the liquid.

Capillarity affects the ability of plants to obtain water. The narrower the spaces between soil particles the higher the water will rise in soil.

Soil particles with stronger attraction for water enable the water to rise higher and reach more plant roots. Inside the plant, very thin tubes lined with water attracting substances help water move from the roots to the leaves. The columns of water molecules in the tubes are linked by attractive hydrogen bond forces. Water reaching the leaves can be used to carry out photosynthesis.

Some **pesticides** can enter animal bodies through capillaries (pores). The liquid carrying the pesticide can have its properties modified to increase the absorption of the pesticide. Addition of surfactant spreads the pesticide over the surface and increases the ability to enter pores.

The ability of water to move through capillaries increases with temperature as the viscosity of water halves from 0°C to 30°C. Water drains from and moves up through soil faster in summer than winter.

When you dry your skin with a towel you use capillarity to draw water molecules away from your skin. The water is attracted into the small spaces between the cellulose fibres of the towel. Before a person donates blood their skin is pricked and a capillary tube placed in the drop to draw up a blood sample for testing. The capillary tube is made of glass or hydrophilic plastic.

Activities you carried out in the first module used capillarity – the paper chromatography of food colourings and burning of a candle.

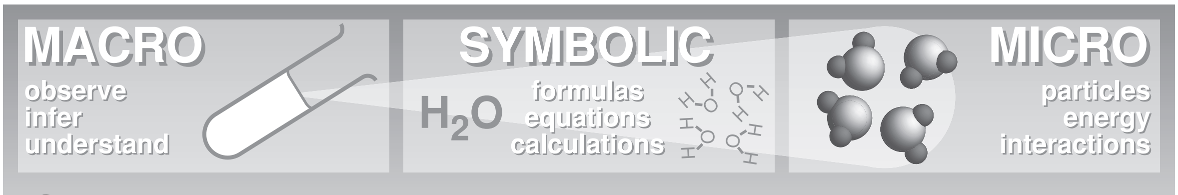
Explain, in words or by diagrams, the importance of capillarity in these two activities.

Check your answers.

Would you like to feel the energy released when hydrogen bonds form?

Place a drop of glycerol $C_3H_8(OH)_3$ on your tongue.

What you feel is the energy released as hydrogen bonds are formed between the OH (hydroxy) groups in glycerol and OH groups in water. (You will also experience a taste characteristic of molecules with many OH groups such as sugars. Glycerol is used in some toothpastes, icing and sugarless chewing gums.)



Suggested answers

Comparing similar sized molecules

1 MP and BP of water are greater than those of similar sized molecules.

2 The difference is not due to the number of atoms in isoelectronic molecules as the lowest values are for Ne and CH_4, the two molecules with the smallest number and largest number of atoms.

Constructing Lewis electron dot structures

Molecule	CH_4	NH_3	H_2O	HF
Central atom	C	N	O	F
Combining power of central atom	4	3	2	1
Number of valence electrons around the central atom **by itself**	4	5	6	7
Number of valence electrons around the central atom **in the molecule**	8	8	8	8
Number of valence electron pairs around the central atom **in the molecule**	4	4	4	4
Number of **bonding** electron pairs around central atom **in the molecule**	4	3	2	1
Number of **nonbonding** electron pairs around central atom **in the molecule**	0	1	2	3

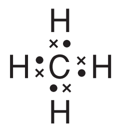
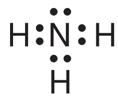
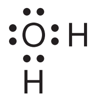
2 Combining power of an element is either the number of electrons in the valence shell of the atom (eg C) or the number of electrons required to give a valence shell of 8 (eg N, O, F).

Molecule	H_2O	H_2S
Central atom	O	S
Combining power of central atom	2	2
Number of valence electrons around the central atom **by itself**	6	6
Number of valence electrons around the central atom **in the molecule**	8	8
Number of valence electron pairs around the central atom **in the molecule**	4	4
Number of **bonding** electron pairs around the central atom **in the molecule**	2	2
Number of **nonbonding** electron pairs around the central atom **in the molecule**	2	2

3 The number of electrons around the central atoms in these molecules is eight.

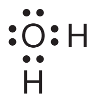
Modelling electron clouds

1

Molecule	CH_4	NH_3	OH_2
Electron dot formula			
Number of electron pairs around central atom	4	4	4
Number of nonbonding electron pairs around central atom	0	1	2
Number of bonding electron pairs around central atom	4	3	2
Number of electron clouds around the central atom	4	4	4
Bond angles expected around the central atom	109.5°	109.5°	109.5°
Actual bond angles	109.5°	107°	105°

2 Repulsion by the single large nonbonding electron cloud in ammonia NH_3 reduces the angles between N–H bonds from 109.5° to 107°. Repulsion by two large nonbonding clouds in H_2O reduces the angle between the O–H bonds from 109.5° to 105°. The four C–H bonds are all at 109.5° to one another because there are four equal bonding electron clouds around the central carbon atom, C.

1

Molecule	H_2O	H_2S
Electron dot formula		
Number of electron pairs around central atom	4	4
Number of nonbonding electron pairs around central atom	2	2
Number of bonding electron pairs around central atom	2	2
Electron clouds around central atom	4	4

2 Water and hydrogen sulfide have similar structures because the O and S are each surrounded by two bonding electron pairs and two nonbonding electron pairs.

Do water molecules have dipoles?

1 The stream of water should be attracted towards the charged object instead of going straight down.

2 Water molecules have dipoles because they are attracted towards charged objects. The end of the molecule that is attracted to the charged object will be closer than the oppositely charged repelled end.

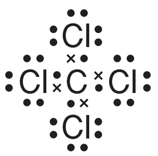
Modelling the electron clouds

1 Two balloons arrange themselves as far away as possible from one another (at 180° to one another).

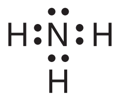
2 Three balloons arrange themselves in a plane at 120° to one another.

3 The photograph represents the electron clouds in BF_3.

4



Polar bonds in ammonia

1 

2 4

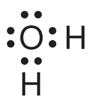
3 3

4 N has electronegativity (EN) of 3.04 while H has EN of 2.20. The difference of 0.84 indicates a polar bond

5 Three polar bonds at about 107° to one another are partly pointing in a common direction. Instead of cancelling they add together.

6 The polar N–H bonds add to produce a polar NH_3.

Polar bonds in water

1 

2 4

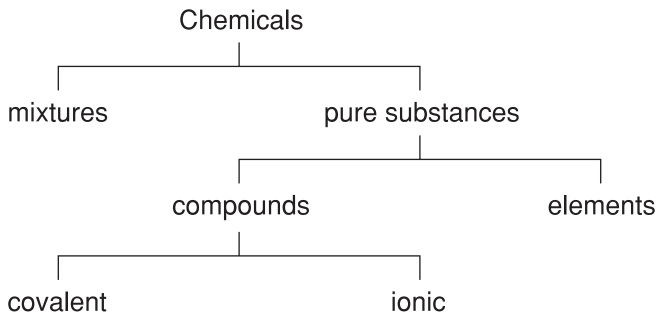
3 2

4 O has electronegativity (EN) of 3.44 while H has EN of 2.20. The difference of 1.24 indicates a very polar bond

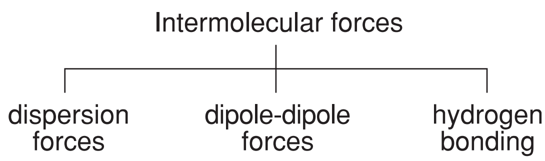
5 Two polar bonds at about 105° to one another are partly pointing in a common direction. Instead of cancelling they add together.

6 The polar O–H bonds add to produce a polar H_2O.

Keys and Venn diagrams



Intermolecular forces



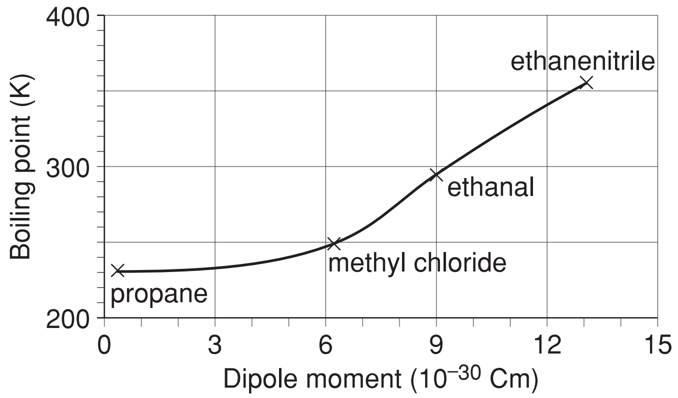
Dispersion forces

The BPs of the halogens increase in the order F_2, Cl_2, Br_2, I_2 because the number of electrons increases and this increases the dispersion forces.

Dipole-dipole forces

1) (a) As the dipole moment of the molecules with similar molecular masses and therefore similar numbers of electrons increases the BP increases.

b)



2 The very small dipole for propane indicates that most of the intermolecular forces in propane are *dispersion* forces.

3 As the molecules become more polar as shown by the increasing *dipole* moment values, dipole-dipole forces between the polar molecules become more important

BP increases as the dipole-dipole forces increase.

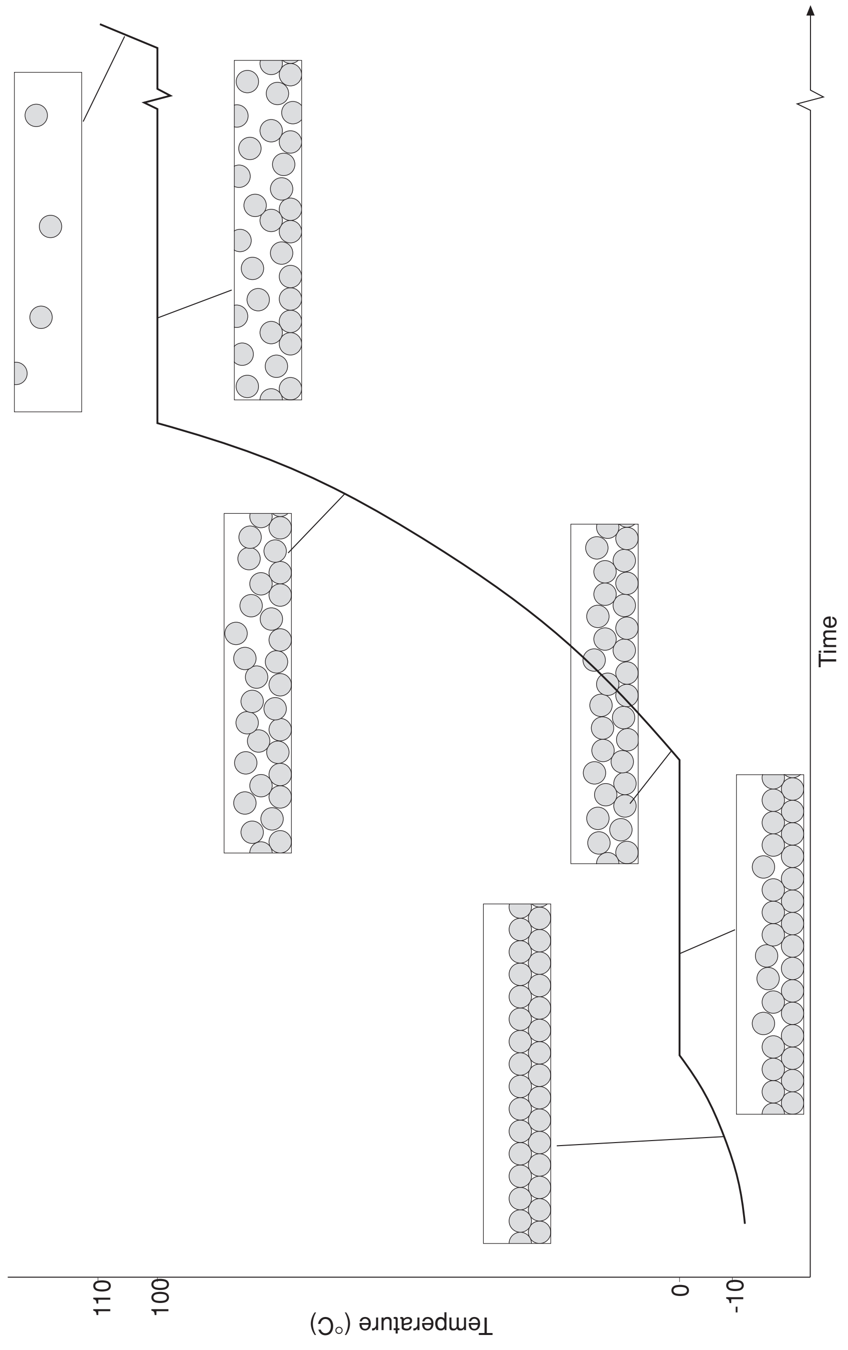
Hydrogen bonding

1 a) The bottom line for Group 14 hydrides (CH_4 etc.).

b) Dispersion.

2 H_2O HF NH_3 .

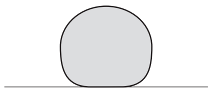
Water heating curve



Surface tension

1 Put water on your hair to stick hairs together and reposition them

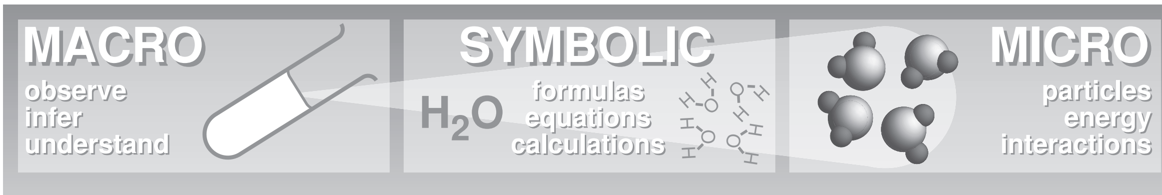
2 When one end of the floating object is pushed under the skin it sinks

3 

Capillarity and living things

In paper chromatography solvent and dissolved colourings moved up through spaces between the cellulose fibres of the paper. The weaker the attraction of the colouring to the cellulose the further that colouring travelled.

In burning a wax candle molten (liquid) wax moves up the spaces between the threads of the wick before vaporising and burning in the air.



Exercises - Part 2

Exercises 2.1 to 2.3 Name: ______________________

Exercise 2.1: Relative viscosity of glycerol and water

Method:

Results:

Calculations:

Conclusions:

Exercise 2.2: Designing another experiment

In the section on viscosity you considered two sets of equipment to measure the relative viscosity of glycerol compared with water.

In this exercise you are asked to explain how you would measure the relative viscosity of glycerol using other equipment. Instead of the high density bead/ball used in Set A you are asked to use an air bubble. The air bubble would have a low density compared with glycerol and water. Imagine you are supplied with a measuring cylinder to measure the volumes of any liquid used.

Using at least one labelled diagram outline in steps how you would measure the relative viscosity of glycerol compared with water using air bubbles if you were supplied with:

a measuring cylinder

two identical test tubes and stoppers

stopwatch

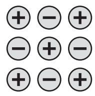
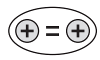
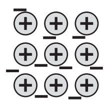
glycerol

water

paper towels.

Exercise 2.3: Summary of intermolecular forces

Here is a summary of the intramolecular forces that you studied in the first module, *The chemical Earth.*

Intramolecular force	**Attraction between:**	**Diagram of model**	**Energy range typical value (kJ/mol)**
ionic bond	cations and anions		400–4000 1000
covalent bond	two nuclei and shared electron pair		100–1000 300
metallic bond	cations and delocalised electrons		100–1000 300

Complete the summary below for the intermolecular forces you have studied in this part of the third module, *Water.*

Intermolecular force	**Attraction between:**	**Diagram of model**	**Energy range typical value (kJ/mol)**
	δ+ H and δ– N, O or F		10–40 30
	dipoles		5–25 10
	temporary dipole and induced dipole		0.1–40 3

Chemistry
Preliminary Course
Stage 6

NEW SOUTH WALES DEPARTMENT OF EDUCATION AND TRAINING
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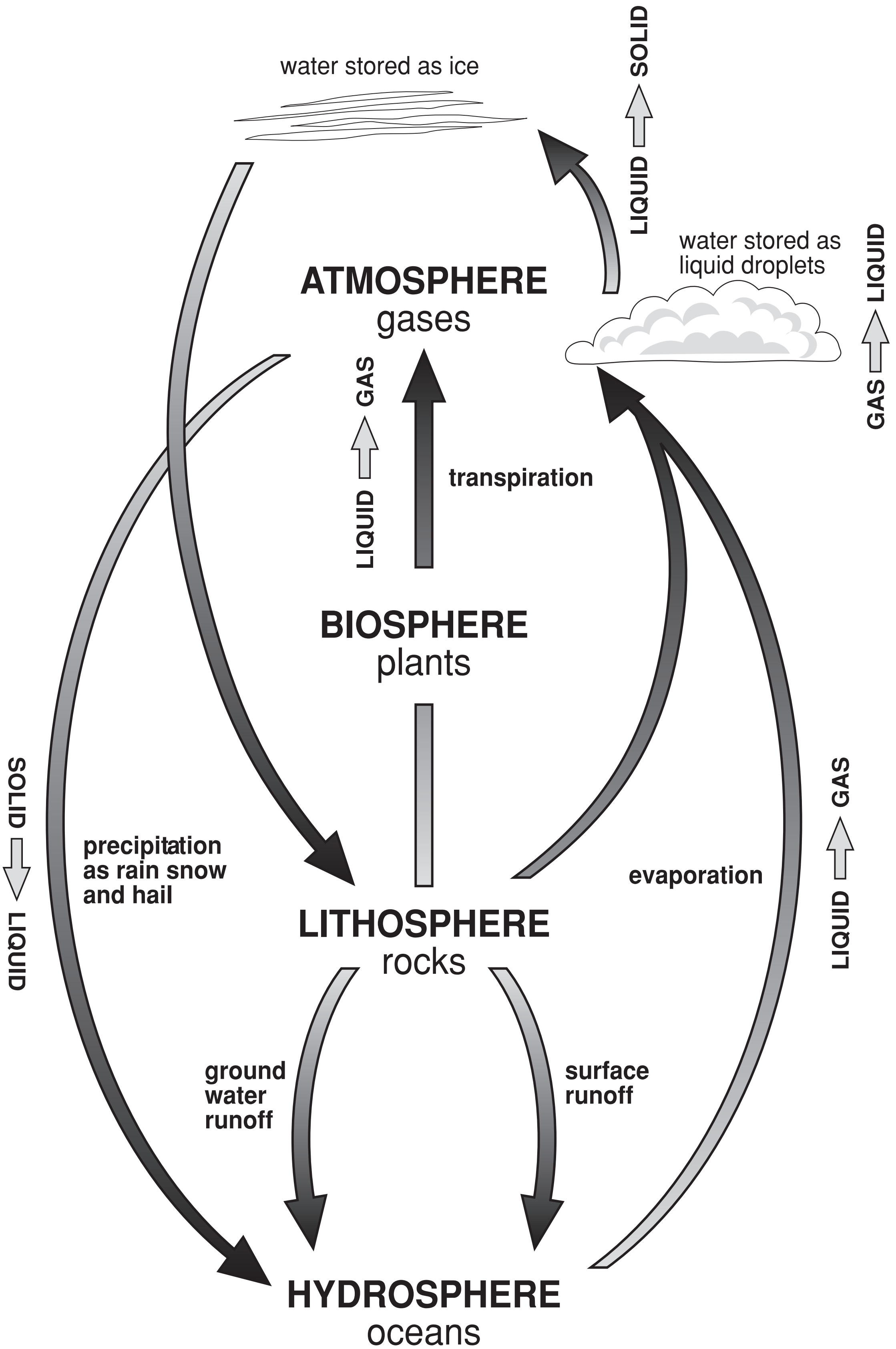
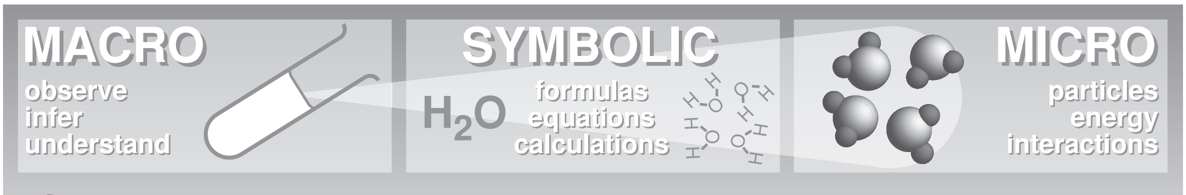
Water

Part 3: The important solvent

Incorporating October 2002 AMENDMENTS


water stored as ice
LIQUID → SOLID
water stored as liquid droplets
GAS → LIQUID
ATMOSPHERE
gases
LIQUID → GAS
transpiration
BIOSPHERE
plants
SOLID → LIQUID
precipitation as rain snow and hail
evaporation
LIQUID → GAS
LITHOSPHERE
rocks
ground water runoff
surface runoff
HYDROSPHERE
oceans


MACRO
observe
infer
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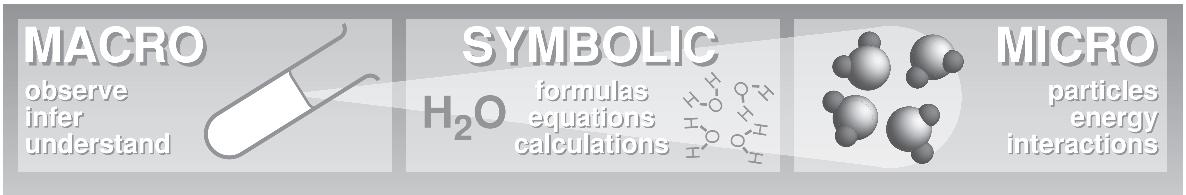
SYMBOLIC
H2O
formulas
equations
calculations
MICRO
particles
energy
interactions

Contents



Introduction

The polarity and hydrogen bonding ability of water have made it the most powerful solvent on Earth. Water is the solvent in oceans, run-off, lakes, and the fluids inside and outside living cells. It is the main transporter of gases, solids and liquids between different environments in the atmosphere, hydrosphere, lithosphere and biosphere. Most of the mass of life – biomass – is found in the oceans where animals depend on the solubility of oxygen and plants depend on the solubility of carbon dioxide.

Multicellular organisms depend on the movement of water in forms such as blood and sap to transport chemicals into and out of cells. The ability of water – nature's transport medium – to dissolve many substances enables water to supply cells with a range of **nutrients** and remove wastes.

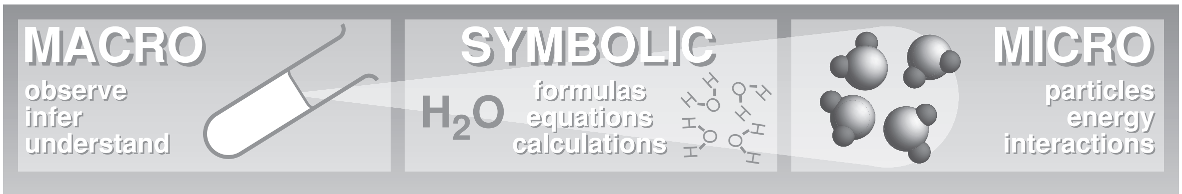
Sperm swim through aqueous solutions rich in dissolved sugars that provide energy to reach and fertilise the eggs of plants and animals. Animal embryos develop in and are protected from shock by aqueous solutions.

In Part 3 you will be given opportunities to learn to:

- explain changes, if any, to particles and account for those changes when the following types of chemicals interact with water
  - a soluble ionic compound such as sodium chloride
  - a soluble molecular compound such as sucrose
  - a soluble or partially soluble molecular element or compound such as iodine, oxygen or hydrogen chloride
  - a covalent network structure substance such as silicon dioxide
  - a substance with large molecules such as cellulose or polyethylene
- analyse the relationships between the solubility of substances in water and the polar nature of the water molecule.

In Part 3 you will be given opportunities to:

- perform a first-hand investigation to test the solubilities in water of a range of substances that include ionic, soluble molecular, insoluble molecular, covalent networks and large molecules
- process information from secondary sources to visualise the dissolution in water of various types of substances and solve problems by using models to show the changes that occur in particle arrangement as dissolution occurs.



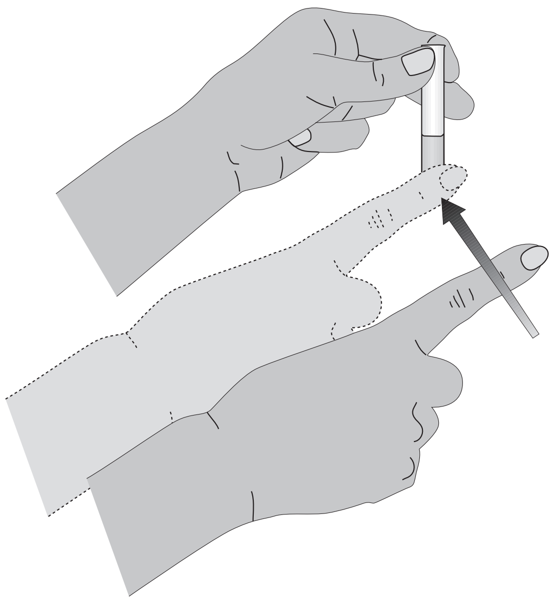
Investigating solubility

Different chemicals in water

To test the solubility of different substances in water:

- use about the same small amount of each substance
- shake each substance with the same amount of water, at the same temperature, for the same amount of time and in the same way.
- observe the volume of substance you added to the water and its volume after shaking and after letting the substance stand for a period of time before recording its solubility.

Suitable substances for testing include table salt, table sugar, kerosene, mineral turps, graphite, sand, cottonwool.



How to hold a test tube and shake its contents.

You will carry out this investigation qualitatively. You could report the solubility using words such as none, low, medium, high or by using symbols such as –, √, √√, √√√.

If the experiment was to be carried out quantitatively you would need:

- a balance for weighing small amounts
- a filter to filter off undissolved material
- a means of drying the undissolved substance collected as residue on the filter.

Describe how you would quantitatively measure the solubility of a salt in water. Set your procedure out as steps.

Check your answer.

For your qualitative investigation estimating solubility you will need:

- an ionic substance such as table salt NaCl
- a polar molecular substance such as table sugar sucrose $C_{12}H_{22}O_{11}$
- a non-polar molecular substance such as hydrocarbon molecules in liquid kerosene or mineral turps (eg $C_{12}H_{26}$) or solid candle wax (eg $C_{40}H_{82}$)
- a covalent network such as the element carbon C or compound silicon dioxide, SiO_2 (the main substance in sand grains)
- a large molecule substance from living things such as cellulose $(C_6H_{10}O_5)_n$ in cotton wool. The term $_n$ is a large number indicating that the large molecule (polymer) is made up of a large number of identical $C_6H_{10}O_5$ units (monomer).
- a large synthesised molecule substance such as polyethylene plastic $(C_2H_4)_n$ Recyclable polyethylene = polythene = polyethene = PE Plastic containers are marked with a triangle symbol containing a two for flexible high density HDPE or a four for rigid low density LDPE.

Explain what information is given by the formula $(C_2H_4)_n$ for polyethylene

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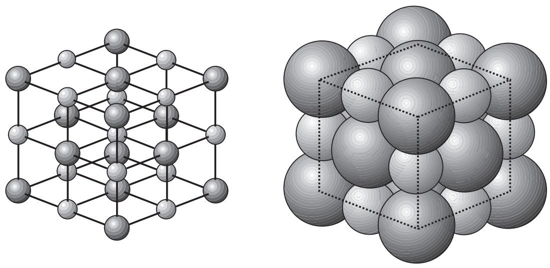
Check your answer.

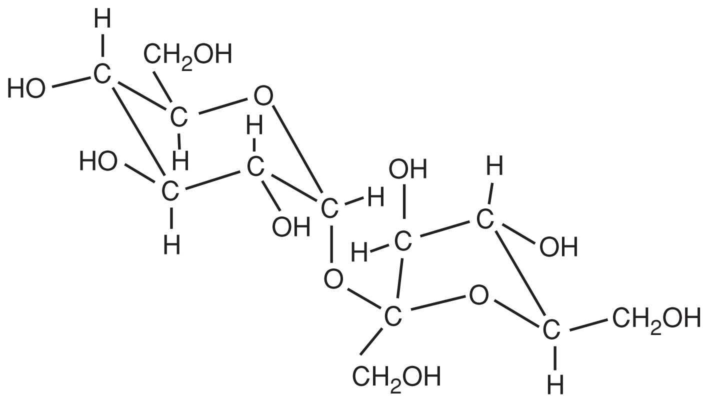
Method:

Using the advice given at the very beginning of this activity, investigate the solubility in water of these six different chemical structures.

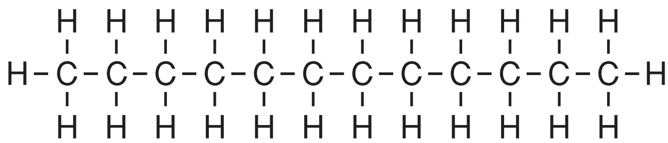
NaCl   Two representations shown. The smaller balls represent Na^+ while the larger balls represent Cl^-



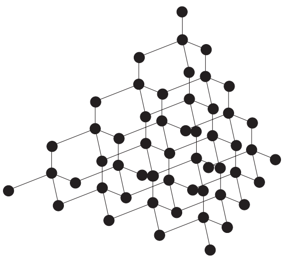
$C_{12}H_{22}O_{11}$



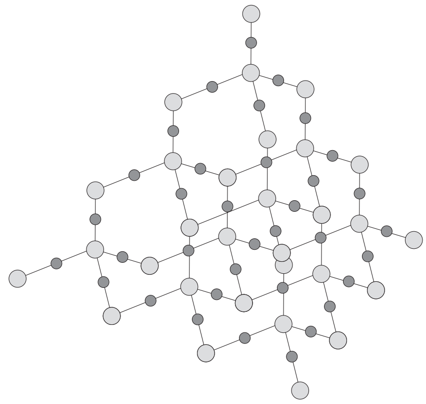
$C_{12}H_{26}$



C or SiO_2

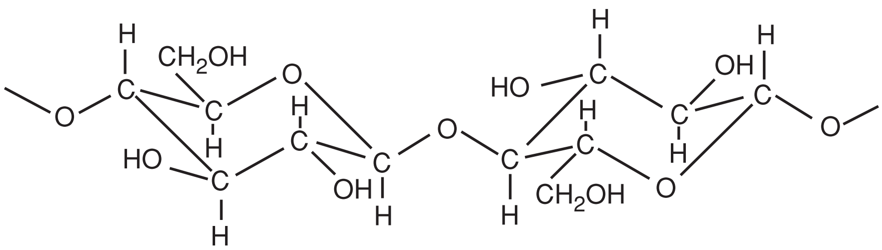


diamond C

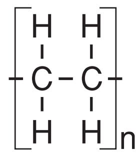


silicon dioxide SiO_2

cellulose $(C_6H_{10}O_5)_n$



polyethylene $(C_2H_4)_n$



Results:

Tabulate (organise into a table) your results for the solubility of NaCl, $C_{12}H_{22}O_{11}$, $C_{12}H_{26}$, C or SiO_2, $(C_6H_{10}O_5)_n$ and $(C_2H_4)_n$.

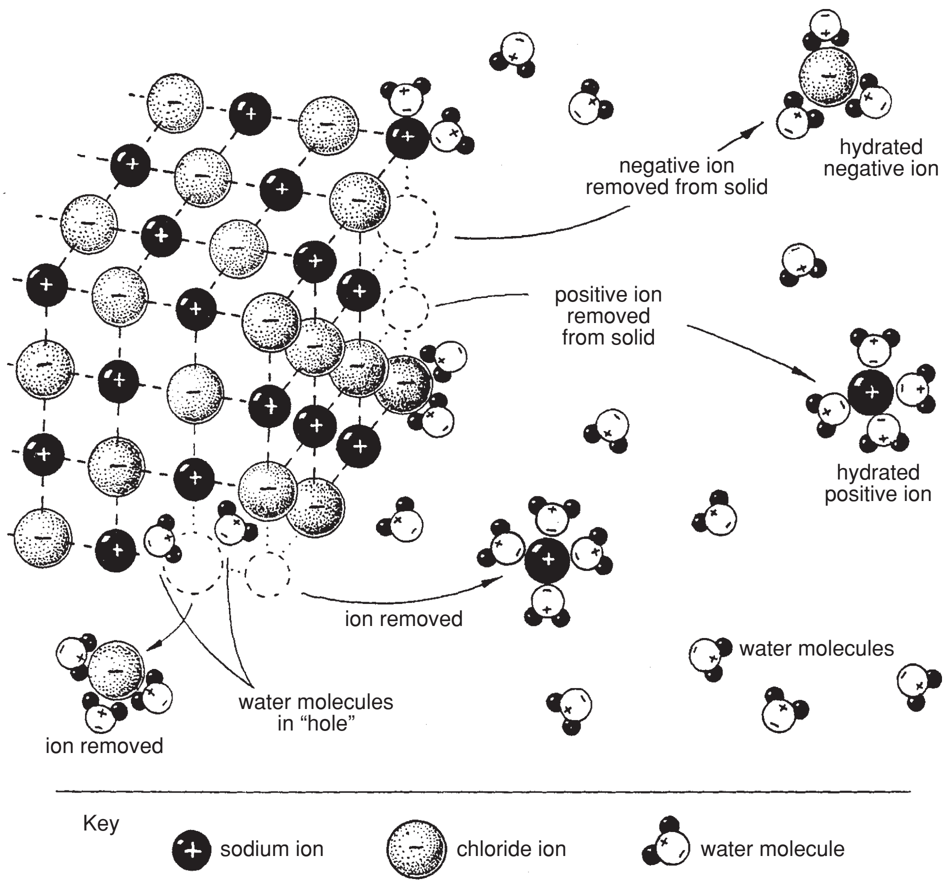
Check your answers.

Solubility and forces of attraction

To explain solubility of a solute in a solvent three sets of forces of attraction must be considered:

- between solute and solvent particles
- between solute particles
- between solvent particles.

For solubility to occur the interactions between the solute and solvent particles must be stronger than those between solute particles and those between solvent particles.



Source: Messel H (chair). (1964.) *Science for high school students.* The Foundation for Nuclear Energy. University of Sydney.

Ionic solids and water

Consider how a sodium chloride crystal dissolves in water. Refer to the diagram on the previous page.

The polar water molecules are attracted to both positive sodium ions and negative chloride ions. Note the different orientation of water molecules around Na^+ and Cl^-. The negative end of water dipoles are closest to Na^+ while the positive end of water dipoles surrounds Cl^-

Solute–solvent attractions > solute attractions + solvent attractions.

Three processes occur in the dissolving of an ionic solid in water.

1 The breaking down of the ionic lattice and separation of positive and negative ions absorbs considerable energy. This is aided by the ability of water molecules to reduce the **electrostatic forces** of attraction between the oppositely charged ions. Water molecules can reduce the force of attraction between charged ions to one eightyieth (1/80) of its original value. Water at 20°C is said to have a **dielectric constant** of 80 compared with air.

2 Some molecules of water separate from one another. This requires the breaking of hydrogen bonds and so absorbs energy. The hotter the water the more energy is available to break these hydrogen bonds. This is one reason why hot water is usually better at dissolving an ionic solid than cold water.

3 As the polar water molecules come closer to each ion a significant amount of energy is released to the surroundings. These forces of attraction between dissolved ions and polar water molecules are called ion-dipole forces. The attachment of polar water molecules to ions is called **hydration.**

If you have anhydrous copper(II) sulfate $CuSO_4$ available you can detect the heat of hydration released when water is added to it.

Read all the instructions below *before* you start this activity.

- Add cold water from a tap to quarter fill a small test tube. Measure the water temperature with a thermometer.
- Add about a rice grain volume of white anhydrous $CuSO_4$ to the water.
- Stir the solution with the thermometer until all the white anhydrous $CuSO_4$ is dissolved.
- Measure the temperature again.

If the heat energy released by formation of ion-dipole forces is greater than the total energy needed to separate ions in a lattice and some water molecules from one another then an ionic solid will probably be very soluble. Most ionic solids are soluble in water to some extent but a few are very insoluble. Temperature of the water, the particular ions in the solid and factors beyond the scope of this course (such as **entropy**) can have an effect on the solubility.

Soluble molecular compounds and water

Molecular compounds which are soluble in water are either polar or contain polar groups which attract water molecules.

Observing the solution of sucrose in water.

You will need:

- a clear glass container of water
- sucrose (table sugar or cane sugar). The sugar is best used in the form of a sugar cube or large crystals such as ‘coffee sugar’. If you cannot get a cube or large crystal use a pile of small crystals on a teaspoon.

What you will do:

1 Wait until any bubbles in the water in the container have come to the surface and the water looks clear.

2 Place a corner of the sugar cube or large crystal just below the surface of the water and observe. If you are using a pile of small crystals on a teaspoon place part of the sugar below the surface of the water and observe.

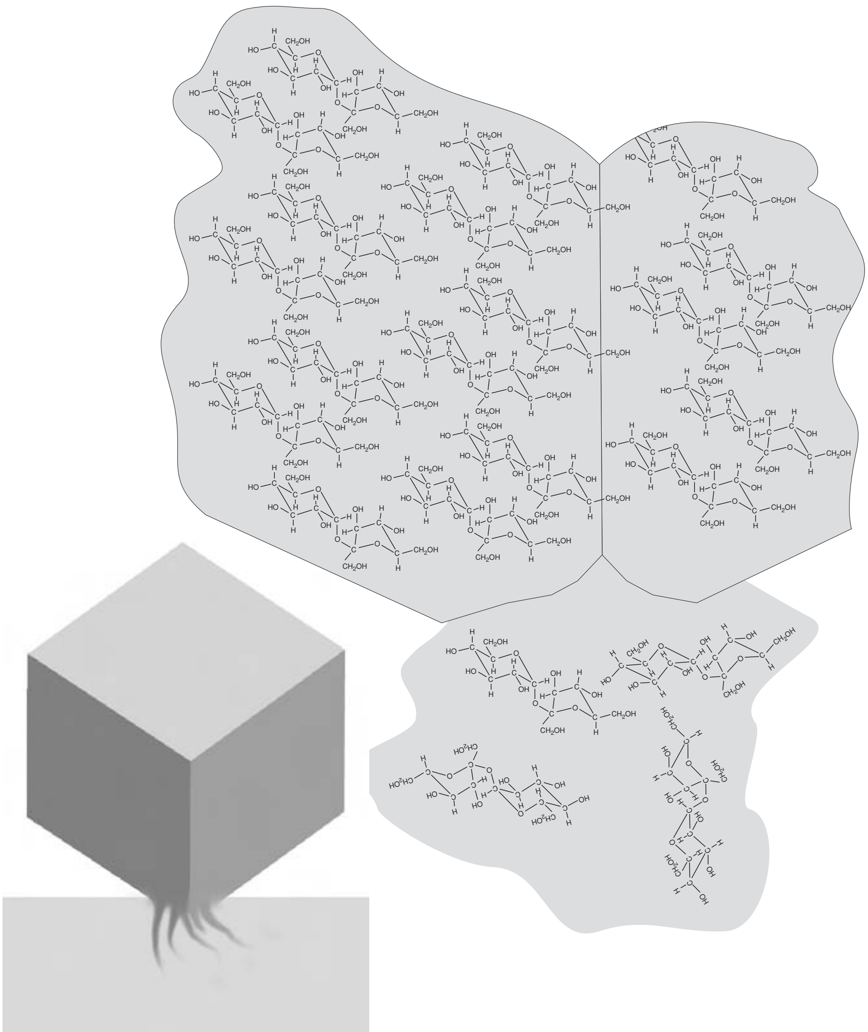
3 Use the factual statements below to explain your observations.

   - Sucrose has a density of 1.6 g cm^{-3}.
   - Water has a density of 1.0 g cm^{-3}.
   - Aqueous sucrose solution has a density greater than 1.0 g cm^{-3}.
   - Light is bent as it passes through different densities of liquid.
   - Water is a polar molecule.
   - Sucrose molecules contain polar hydroxy groups.
   - Hydroxy groups in different molecules can hydrogen bond to one another.

Check your answers.

Visualising the dissolving of sucrose

Complete the diagram below to show how a sucrose crystal dissolves in water by drawing in water molecules. Be careful in showing the orientations (directions in space) taken up by water molecules around the sucrose molecules. As in the previous diagram showing how a sodium chloride crystal dissolves in water, there is no need to show every water molecule.



Non-polar molecules and water

Most non-polar molecules and water do not mix – they are said to be **immiscible**. There is negligible attraction between the non-polar solute and the polar solvent molecules.

Immiscible liquids arrange themselves according to density.
Most non–polar carbon compounds such as $C_{12}H_{26}$ have a density less than 1.0 gcm^{-3} and so float on top of water. This is why petroleum based fires are not fought with water. The petroleum rises to the surface of the water and continues to burn.

Explain the low solubilities (at 25°C and 101.3 kPa) of:

- iodine I_2 0.34 g L^{-1}.
- nitrogen N_2 in water 0.02 g L^{-1}
- oxygen O_2 in water 0.04 g L^{-1}.

Check your answers.

You may have used iodine solution in testing for starch or as an antiseptic that kills bacteria. These aqueous iodine solutions contain some potassium iodide KI or ethanol C_2H_5OH to increase the solubility of the iodine.

The iodide ions in KI solution react with iodine molecules to form the water soluble tri-iodide ion $I_2 + I^- \rightarrow I_3^-$. As I_2 molecules react with starch or bacteria this reaction reverses making more iodine available. This is called a reversible reaction.

The ethanol molecule has a non-polar C_2H_5 group that is more attractive to the I_2 molecules. Ethanol is a very useful solvent for both polar and non-polar molecules. C_2H_5OH has a hydrophobic non-polar ethyl group C_2H_5 as well as a hydrophilic polar hydroxy group OH.

Covalent networks and large molecules in water

Covalent networks such as in carbon and silicon dioxide are not soluble in water. The many strong covalent bonds joining the atoms cannot be broken by attraction to water molecules.

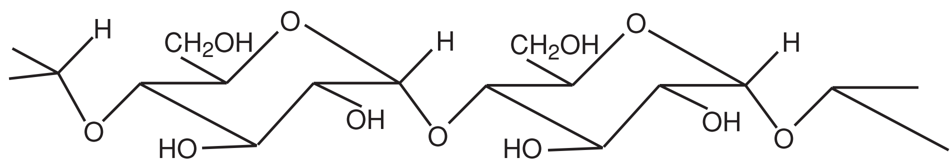
Most large molecules such as polyethylene and cellulose are insoluble in water. The larger the molecule, the less soluble it is likely to be. However, the main factor that determines the solubility of large molecules in water is the accessibility of polar groups such as hydroxy groups OH, amino groups NH_2 or acid groups COOH.

Polyethylene is non-polar so it does not dissolve in water. Polyethylene is used to make water buckets and plastic bags for food.

If you look at the structural formula for cellulose you will see polar OH groups. However in cellulose most of these are used to join the long flat cellulose chains to adjoining cellulose chains. The long chains lie side by side in bundles held together through hydrogen bonds between neighbouring OH groups. Consequently most of the OH groups are not accessible to the water molecules for hydrogen bonding. The moving water molecules cannot get between the cross-linked cellulose chains and separate them.

Some large molecules with polar groups on the outside such as the amylose form of starch and some proteins are water soluble.

Amylose starch, like cellulose, is a polymer made up of glucose units joined together. Notice below how the CH_2OH groups in amylose are on the same side while they alternate between the two sides in cellulose.



The hydroxy groups in amylose are not used to hydrogen bond to other amylose molecules the way cellulose molecules hydrogen bond to one another. Water molecules can hydrogen bond to these accessible hydroxy groups and separate amylose molecules from one another. Amylose is water soluble.

Polar groups are very important in determining the shapes of proteins in water. In the diagram below the **helices** represent water soluble human growth hormone (hGH).

The flat strands in the diagram below and to the side of the hGH represent the water soluble part of a receptor on the outside of a human cell membrane. The helices are called **α helices** (alpha helices) while the flat parts are called **β strands** (beta strands).

hGH (helices) fitting into the water soluble part of a receptor (flat strands) on the outside of a human cell membrane.
Diagram kindly supplied by Dr. Bret Church from the Garvan Institute of Medical Research, Science Week Committee, May 2000

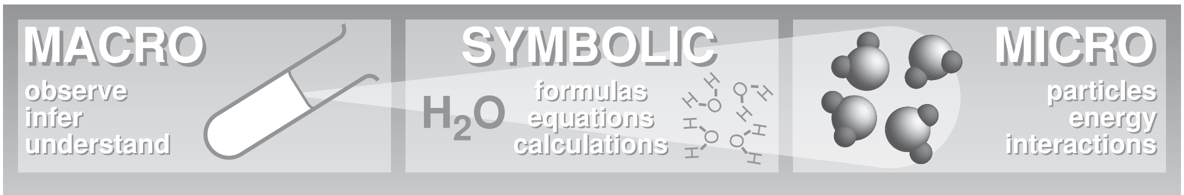
Both the hGH and its receptor have many polar groups which are attracted to and hydrogen bond with water molecules.

A water insoluble part of the receptor, although not shown in this diagram, extends below the water soluble part. The water insoluble part has mostly non-polar groups on its outside surface and so is found in non-polar regions of the cell membrane.

Computers are very important in helping scientists visualise the shape of large biological molecules in water. This enables them to design **antibiotics** and other chemical treatments of disease.

Websites showing how large molecules can be visualised are found at the chemistry section of http://www.lmpc.edu.au/science.

Any group on a large molecule that contains hydrogen bonded to O or N will be polar and able to hydrogen bond to water. This increases the water solubility of the large molecule.



‘Like dissolves like’

A generalisation, not an explanation

‘Like dissolves like’ is a useful generalisation or guide for solubility of solutes in solvents.

However it is important that you realise that this generalisation does not explain – it is not a chemical explanation.

An exam question asking you to explain is asking you to:

- relate cause and effect OR
- make the relationship between things evident OR
- provide why and/or how.

In your answer in a chemistry exam you would be expected to demonstrate knowledge and understanding of the concepts of the chemistry course.

If you were asked in an exam to ‘explain why sugars are soluble in water’ it is not likely that the answer ‘like dissolves like’ would receive a mark.

‘Polar substances dissolve in polar liquids’ is a slightly better answer but still so general that it is unlikely to receive a mark.

‘Polar substances dissolve in polar liquids, non-polar substances dissolve in non-polar liquids’ may give more information but it is not as good as the previous answer. There is ambiguity. The student has not indicated whether they think sugars are polar substances or non-polar substances.

‘Sugars are polar substances that dissolve in polar liquids like water’ is a better answer but may still not be sufficient to receive a mark.

Listed below are other answers for the question ‘Explain why sugars are soluble in water’. Number them in order from best answer (1) to poorest response(s).

‘Sugars consist of polar molecules that are attracted to polar water molecules’

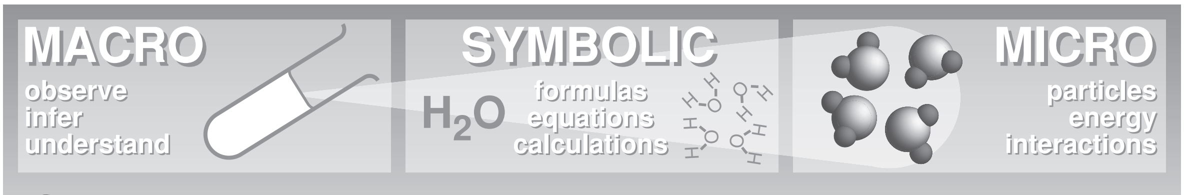
‘The attractive forces between sugar molecules and water are greater than the attractive forces between sugars molecules and the attractive forces between the polar water molecules’

‘Hydrogen bonding brings the polar sugar and polar water molecules together’

‘Hydrogen bonding between the hydrogen in sugar and the hydrogen in water is greater than the hydrogen bonding between polar sugar molecules and hydrogen bonding between polar water molecules’.

‘Hydroxy groups in sugars can hydrogen bond to polar water molecules. Water molecules move between and around sugar molecules, hydrogen bond to hydroxy groups and move the sugar molecules into solution’.

Check your answers.



Solubility of gases

Most solids and liquids become more soluble as temperature rises. However, the solubility of all gases decrease as temperature rises. As the temperature of the solvent rises, the extra and faster movement of particles makes it more likely that dissolved gas molecules will move towards the surface at high speed. If the dissolved gas molecule reaches the surface at high speed it will probably escape from the liquid phase.

Using the data in the table on solubility of gases, answer questions applying your knowledge of intermolecular bonding.

1 Six of these gases have high solubilities because their molecules react with water to form soluble ions. This ionisation increases the solubility of gases significantly. Give the formulas for the six highly soluble gases.

2 Explain why argon is more soluble than helium.

3 Why do you think hydrogen and helium are the least soluble gases?

4 A large fish, like a human, consumes a mole of oxygen each hour. At 20°C how much water would need to enter its gills?

Check your answers.

Gas formula and name	Solubility of gas (g per L) of water at different temperatures		
	20°C	**40°C**	**60°C**
Ar argon	0.059	0.042	0.030
CO carbon monoxide	0.028	0.021	0.015
CO_2 carbon dioxide	1.69	0.97	0.58
Cl_2 chlorine	7.29	4.59	3.30
H_2 hydrogen	0.0016	0.0014	0.0012
HCl hydrogen chloride	730	–	561
H_2S hydrogen sulfide	3.85	2.36	1.48
He helium	0.0015	0.0014	0.0013
N_2 nitrogen	0.019	0.014	0.010
NH_3 ammonia	529	316	168
O_2 oxygen	0.043	0.031	0.023
SO_2 sulfur dioxide	113	54	–

Gases that react with water

The gases carbon dioxide, chlorine, hydrogen chloride, hydrogen sulfide, ammonia and sulfur dioxide have high solubilities in water. They all react with water molecules forming soluble ions.

In reacting with water either H^+ or OH^- ions are produced. H^+ ions form acidic solutions while OH^- ions form alkaline solutions.

Hydrogen chloride dissolves forming an acidic solution called hydrochloric acid.

$$HCl \rightarrow H^+ + Cl^-$$

Ammonia dissolves forming an alkaline solution called ammonia water or ammonium hydroxide solution:

$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$

Note that this solution contains ammonium ion and hydroxide ions but a formula for ammonium hydroxide NH_4OH should not be used. NH_4OH does not exist. If the solution is evaporated the reaction reverses and the ammonia escapes into the air leaving water behind.

Sulfur dioxide reacts with water forming sulfurous acid solution. Sulfurous acid solution contains hydrogen sulfite ions HSO_3^- , sulfite ions SO_3^{2-} and hydrogen ions. If the solution is evaporated the reaction reverses and sulfurous acid H_2SO_3 cannot be obtained.

$SO_2 + H_2O \rightarrow H^+ + HSO_3^- \rightarrow 2\,H^+ + SO_3^{2-}$

Chlorine reacts forming a mixture of hydrochloric acid HCl and hypochlorous acid HOCl. The hypochlorous acid produced is an effective disinfectant; this is why chlorine is used in swimming pools.

$Cl_2 + H_2O \rightarrow H^+ + Cl^- + HOCl$

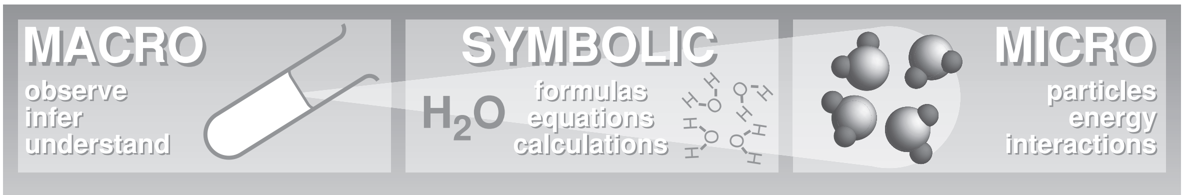
'Rotten egg gas' hydrogen sulfide molecules react with water molecules producing a solution containing hydrogen sulfide ions HS^-, sulfide ions S^{2-} and hydrogen ions.

$H_2S + H_2O \rightarrow H^+ + HS^- \rightarrow 2\,H^+ + S^{2-}$

Carbon dioxide dissolves forming hydrogen carbonate ions HCO_3^-, carbonate ions CO_3^{2-} and hydrogen ions.

$CO_2 + H_2O \rightarrow H^+ + HCO_3^- \rightarrow 2\,H^+ + CO_3^{2-}$

All these reactions are reversible reactions. All of the $\rightarrow$ could be replaced by a symbol $\rightleftharpoons$ showing the reaction goes in a forward and a reverse direction. When a solution of the gas in water is heated the reactions go in the reverse direction. As the temperature rises more and more of the gas comes out of solution. Eventually only water is left.



Diffusion

Diffusion is the movement of particles from a region of high concentration or pressure to regions of lower concentration or pressure. Diffusion occurs when a salt crystal or a sugar crystal dissolves in water. When a solid such as sodium chloride or sucrose dissolves in water their particles are considered to be in the liquid state because they are moving freely.

The membranes around cells are selectively permeable (also called semipermeable). This means small particles such as H_2O, O_2, CO_2, Na^+, Cl^- and $C_6H_{12}O_6$ can diffuse through while large particles such as starch and protein cannot. Small nutrient particles pass into and small waste particles pass out of the cell but larger particles making up the structure of the cell are retained.

If a human body cell has a low concentration of oxygen then oxygen will diffuse from the higher concentration of oxygen in blood into the cell. Carbon dioxide waste increasing in the cell diffuses out of the cell to the blood. Blood carries carbon dioxide to the lungs and out of the body.

Diffusion and scale

Diffusion is very important at the microlevel in moving substances in and out of cells. Movement of particles over small distances is mostly by diffusion – random movement from high concentrations to lower concentrations. Movement of an oxygen molecule from a blood capillary into a cell could take 10^{-2} second.

At the macrolevel diffusion is not the main reason why particles spread. Bulk movements such as convection currents – movement of gas and liquid due to temperature differences – are much more important in moving particles.

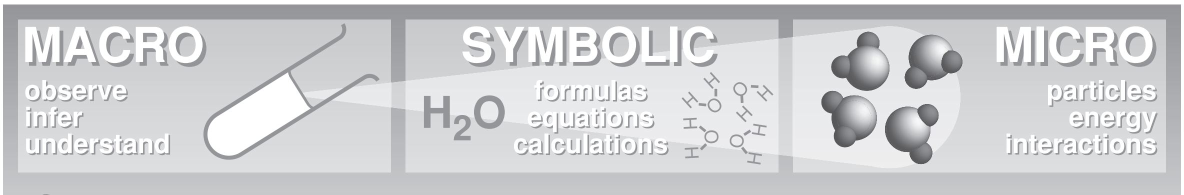
A perfume molecule evaporated from a human body could move at 100 ms^{-1} and collide with air molecules about 10^9 times per second.

In perfectly still air the perfume molecule could take a month to travel just one metre away from its original position while the total zigzag path covered in that month could be over 10^9 m. Yet if someone wearing perfume walks one metre distance from you then you would probably smell perfume molecules in your nose only seconds later.

The movement of that person's body pushing air, air draughts due to pressure differences, convection currents due to temperature differences and the drawing of air into your lungs would have all speeded up the movement of perfume molecules into your nose.

On an even larger scale when poisonous gas or radioactive particles escape the particles do not move from the source evenly. The direction of winds is very important in determining who will be harmed.
The Chernobyl Nuclear accident in Ukraine, in 1986 sent radioactive clouds North West to Scandinavia then south and east to reach Greece in ten days.

Diffusion may not contribute much to large scale movement of particles but it is the main means of transportation of chemicals in and out of cells in biological systems.



Suggested answers

Different chemicals in water

- Weigh out an initial mass of salt
- Add the salt to a known volume of water in a bottle
- Shake the salt with the water in the bottle vigorously
- Leave the bottle for 24 hours
- Filter the solution to separate undissolved salt.
- Dry the filtered undissolved salt and weigh
- Dissolved salt = Initial mass of salt–undissolved salt
- Solubility = dissolved salt/volume of water

$(C_2H_4)_n$ shows that poly(ethene) is made up of many ($_n$) ethene (C_2H_4) monomer units joined to form a polymer unit with molecular mass $_n$ times the molecular mass of ethene.

Sample results for qualitative solubility.

Substance	Formula	Solubility in water
sodium chloride	NaCl	√√√
sucrose	$C_{12}H_{22}O_{11}$	√√√
mineral turps	$C_{12}H_{26}$	–
graphite	C	–
silica sand	SiO_2	–
cotton wool cellulose	$(C_6H_{10}O_5)_n$	–
LowDensityPolyEthylene	$(C_2H_4)_n$	–

Soluble molecular compounds and water

Sucrose molecules contain polar hydroxy groups which attract and can hydrogen bond to water molecules. The moving water molecules move the sucrose molecules into solution. Light is bent as it passes from water through the denser aqueous sucrose solution. The denser sucrose solution sinks to the bottom of the container.

Non-polar molecules and water

Non-polar molecules only have weak dispersion forces between their molecules and water. Consequently their solubilities are low. The more electrons in the molecule and the larger the mass of the molecule then the larger the dispersion forces. The solubilities of these molecules increase with molecular mass.

'Like dissolves like'

4 'Sugars consist of polar molecules that are attracted to polar water molecules'. *No mention of hydrogen bonding or hydroxy groups.*

2 'The attractive forces between sugar molecules and water are greater than the attractive forces between sugar molecules and the attractive forces between the polar water molecules'. *Good description of balance of forces but no information about the type of forces involved.*

3 'Hydrogen bonding brings the polar sugar and polar water molecules together'. *No indication that hydrogen bonding is between hydroxy groups.*

5 'Hydrogen bonding between the hydrogen in sugar and the hydrogen in water is greater than the hydrogen bonding between polar sugar molecules and hydrogen bonding between polar water molecules'. *Hydrogen bonding does not occur between hydrogen atoms – it can occur between the hydrogen of one molecule and an oxygen in another molecule*

1 'Hydroxy groups in sugars can hydrogen bond to polar water molecules. Water molecules move between and around sugar molecules, hydrogen bond to hydroxy groups and move the sugar molecules into solution'. *Most correct and complete answer.*

Solubility of gases

1 HCl NH_3 SO_2 Cl_2 H_2S CO_2.

2 Argon and helium are both non-polar monatomic molecules but Ar has about ten times the mass of He and therefore greater dispersion forces with the water molecules.

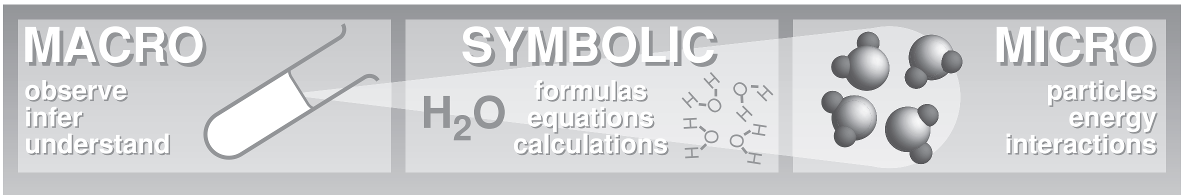
3 Only dispersion forces attract these non-polar molecules to water. Because these are the smallest molecules their dispersion forces with water are the smallest.

A fairer way to compare gases would be to convert solubilities to moles per litre. This gives a direct comparison of solute particles per litre.

For example at 20°C the solubilities of non-polar gases are similar when expressed in mol L^{-1}:

Gas	H_2	He	N_2	O_2	Ar
10^{-4} mol L^{-1}	8	4	7	13	15

4 One mole of oxygen gas O_2 = 32 g. 0.043 g of oxygen dissolves per L of water at 20°C. Therefore minimum of 32 g/0.043 gL^{-1} = 744 L of water needs to enter the gills (assuming all dissolved oxygen is extracted by the gills and consumed).

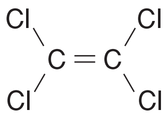


Exercises – Part 3

Exercises 3.1 to 3.2 Name: ______________________________

Exercise 3.1: 'Dry' cleaning

In dry cleaning organic solvents are used instead of water.
Organic solvents dissolve nonpolar substances such as grease or oil out of clothing. Perchlorethylene is the trivial (common) name and tetrachloroethene the systematic name of the most commonly used drycleaning liquid.



a) The C–Cl bonds in this molecule are polar. The C=C bond is between identical atoms and therefore nonpolar. Explain why this symmetrical molecule is nonpolar.

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b) A small amount of water is added to dry cleaning fluid to dissolve any polar substances in food stains. Name two polar substances, soluble in water, you could find in food stains.

____

c) Surfactants are also added to the dry cleaning fluid. What does adding surfactant do to the water in the dry cleaning fluid so that it spreads more readily, dissolving food stains?

____

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Exercise 3.2: Soluble organic compounds

Most organic compounds are insoluble in water. Listed below are four organic compounds which can dissolve in water.

You may find it useful to revise pages 10 and 11 as well as pages 15 and 16 before answering the questions below.

Use their structures to explain why each of these organic compounds is water soluble:

a) Ethanol C_2H_5OH (drinking alcohol, the only alcohol which is poisonous in moderate quantities rather than small quantities)

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____

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b) Glucose $C_6H_{12}O_6$ (the main sugar used in the cells of living things to release energy in respiration)

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c) Propylamine $CH_3CH_2CH_2NH_2$ (used as a sedative)

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d) Ethanoic acid CH_3COOH (acetic acid, the food preservative in vinegar which is a 5% solution of acetic acid in water).

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Chemistry
Preliminary Course
Stage 6

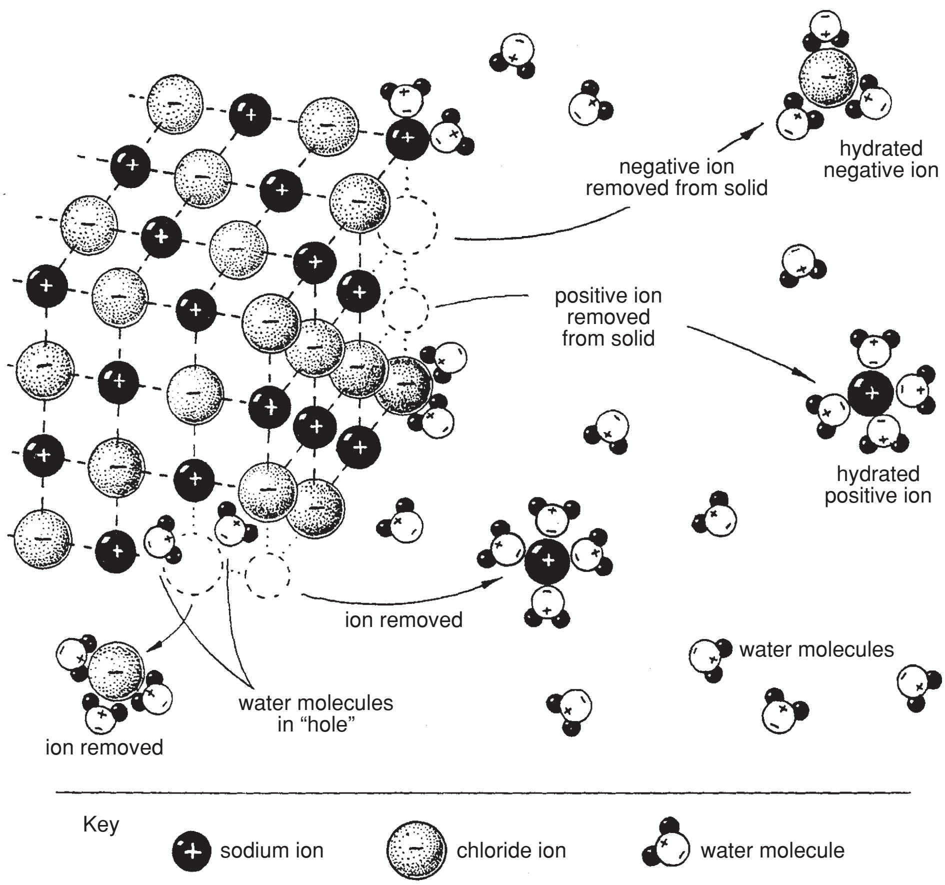
NEW SOUTH WALES DEPARTMENT OF EDUCATION AND TRAINING
TAFE NSW
OPEN TRAINING AND EDUCATION NETWORK - DISTANCE EDUCATION

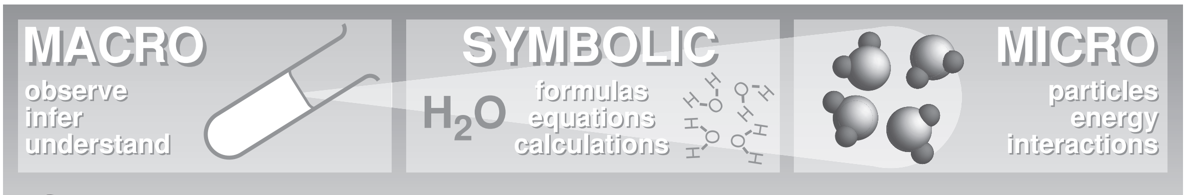
Water

Part 4: Salts in water

Incorporating October 2002 AMENDMENTS

Anion	Cation	
	soluble (no reaction)	insoluble (precipitate forms)
nitrates NO_3^-	all	
acetates CH_3COO^-	all	
chlorides Cl^-	most	Ag^+, (Pb^{2+})
sulfates SO_4^{2-}	most	Ba^{2+},(Ca^{2+}),Pb^{2+},(Ag^+)
sulfides S^{2-}	Group 1, NH_4^+, Group 2	most
hydroxides OH^-	Group 1, NH_4^+, Ba^{2+}	most
carbonates CO_3^{2-}	Group 1, NH_4^+	most




MACRO
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H2O
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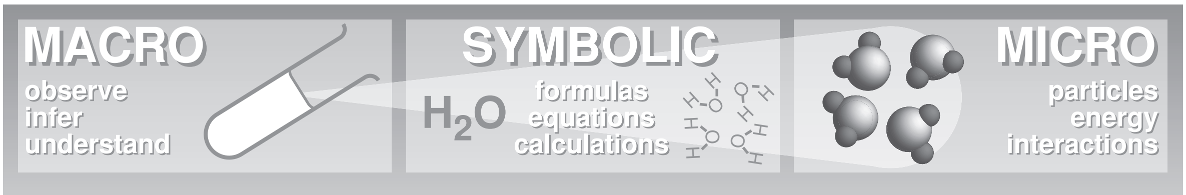
Contents



Introduction

All cations and anions undergo ion-dipole interactions with water molecules. The strength of these interactions and the combinations of ions present in a solution determine whether the ions are likely to remain in solution or precipitate out as insoluble solid.

In this part you will focus on saline solutions. Saline solutions are solutions of salt(s) in water. To an irrigation specialist or water supply expert, saline water is fresh water contaminated with salt so that it may not be suitable for consumption by plants or animals. Oceanographers measure the salinity of sea water which can vary with depth and location of sea water near supplies of fresh water and ice.

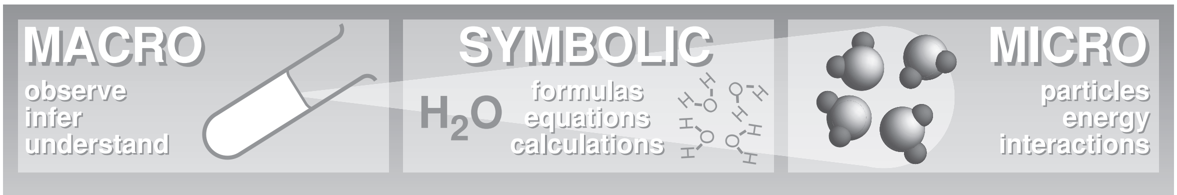
In Part 4 you will be given opportunities to learn to:

- identify some combinations of solutions which will produce precipitates, using solubility data
- describe a model that traces the movement of ions when solution and precipitation occur
- identify the dynamic nature of ion movement in a saturated dissolution
- describe the molarity of a solution as the number of moles of solute per litre of solution using $c = \dfrac{n}{V}$
- explain why different measurements of concentration are important.

In Part 4 you will be given opportunities to:

- construct ionic equations to represent the dissolution and precipitation of ionic compounds in water
- present information in balanced chemical equations and identify the appropriate phase descriptors – (s), (l), (g) and (aq) – for all chemical species

- perform a first-hand investigation, using micro-techniques, to compare the solubility of appropriate salts in solution through precipitation reactions
- carry out simple calculations to describe the concentration of given solutions, given masses of solute and volumes of solution
- perform a first-hand investigation to make solutions to specified volume-to-volume and mass-to-volume specifications and dilute them to specified concentrations $cV = constant$
- calculate mass and concentration relationships in precipitation reactions as they are encountered.



Dissolution and precipitation

Dissolution is the dissolving of a substance in a liquid. The dissolution process is also called the solution process. **Precipitation** is the reverse of dissolution as the substance comes out of solution as a solid. Precipitation means falling out because the solid is usually more dense than the liquid and its crystals fall out of solution to the bottom of the container.

When water containing H^+ passes over limestone rock dissolution of the calcium carbonate $CaCO_3$ in limestone occurs. When this solution evaporates inside limestone caves precipitation of $CaCO_3$ forms **stalactites** and **stalagmites**. Stalactites grow from the ceiling of the cave while stalagmites grow from the ground or floor of the cave.

In this activity you are going to carry out dissolution and precipitation of Epsom salts, (magnesium sulfate) to **simulate** the formation of stalactites and stalagmites.

You will need:

- Epsom salts, magnesium sulfate crystals, from a supermarket, health food shop or pharmacy
- two containers of 200 to 1000 mL volume
- about 20 cm of drawstring cord (eg from an old swimming costume or pyjama pants) or rope or thick string
- a place where you can leave your equipment undisturbed for a week eg. a low shelf away from pets or top of a flat toilet cistern.

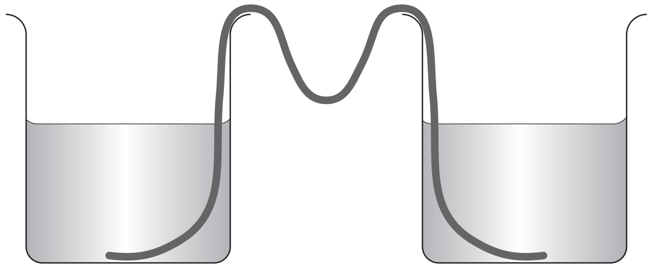
What you will do:

1 Make a concentrated solution of Epsom salt by dissolution of four tablespoons of Epsom salt in 200 mL of water

2 Divide the solution between the two containers

3 Place the containers with a gap of about 10 cm between them

4 Put the cord/rope/string between the two containers as show below:



5 Leave the equipment for about a week.

Results:

Did you grow a 'stalactite', a 'stalagmite' or a connecting column between a stalactite and stalagmite?

Describe what you think happened using the terms dissolution, capillary action, evaporation, concentration and precipitation (in that order).

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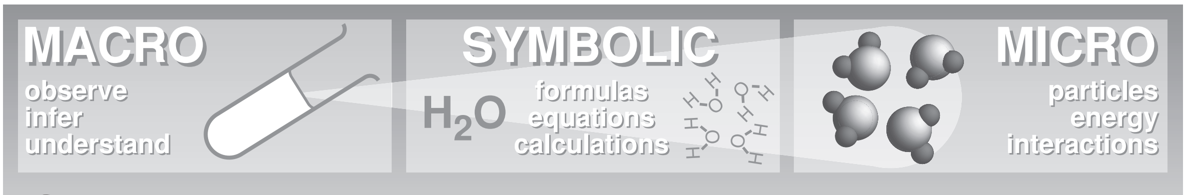
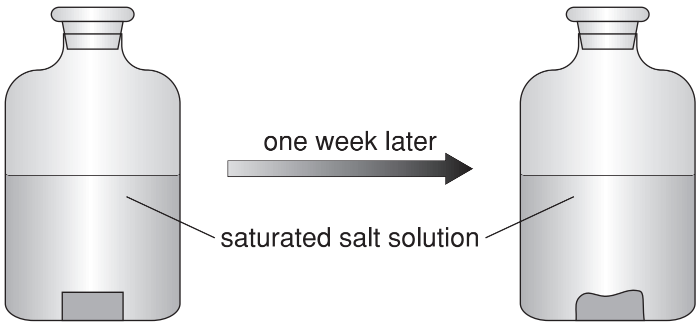
Check your answer.



Saturated solutions

A **saturated solution** cannot dissolve any more solute. If you add a salt to water with shaking and stirring and undissolved salt is still in contact with the solution after a day then the solution is probably saturated.

In a saturated solution still in contact with undissolved solute both dissolution and precipitation are occurring at the micro level. It may be difficult to see dissolution and precipitation occurring at the macro level unless you look at the shape of the undissolved salt crystals. Over a week you should be able to observe that the amount of undissolved salt in an undisturbed container will not change but its shape will.



Two processes are occurring in the container. Salt is going into solution (dissolution) and salt is coming out of solution (precipitation).
Overall Rate of dissolution = Rate of precipitation

When the forward reaction and the reverse reaction are equal in rate the system is said to be in chemical **equilibrium.**

However on one side of the crystals dissolution may be greater than precipitation and that part of the crystals decreases in size. In another part of the crystals precipitation may be greater than dissolution and so that part of the solid increases in size.

The result is a change in solid shape but not the amount of solid in crystal form.

Consider $MgSO_4$ solid in equilibrium with a saturated solution of $MgSO_4$

dissolution $MgSO_4(s) \rightarrow Mg^{2+}(aq) + SO_4^{2-}(aq)$

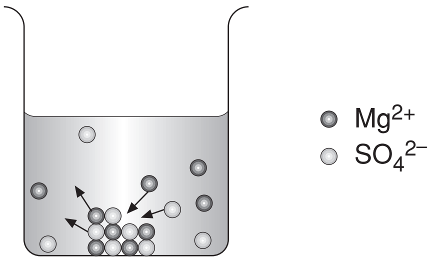
precipitation $Mg^{2+}(aq) + SO_4^{2-}(aq) \rightarrow MgSO_4(s)$

Because this is a **reversible reaction** it can be represented as

$$MgSO_4(s) \rightleftharpoons Mg^{2+}(aq) + SO_4^{2-}(aq)$$

Equal sized arrows going in opposite directions show that the rate of dissolution = rate of precipitation and so the system is said to be in equilibrium.

The simplified diagram below shows solid $MgSO_4$ in equilibrium with its ions in a saturated solution.



Write a dissolution equation, a precipitation equation and a reversible equation for calcium carbonate $CaCO_3(s)$ in contact with a saturated solution of $CaCO_3$.

Check your answer.

If you had $CuSO_4(s)$ in equilibrium with a saturated solution of blue $CuSO_4(aq)$ no change would be seen in the intensity of colour of the solution. This indicates that the concentration of ions in the saturated solution is not changing. At equilibrium all the properties of the system observed at the macro level, except for shape of undissolved solid, are constant.

However at the micro level, if you were able to see the ions, you would see that things are not constant. Continuous change is occurring, that is, the system is **dynamic.**

If **radioactive** copper was present in $CuSO_4(s)$ crystals placed in a saturated solution of $CuSO_4(aq)$ then radioactivity would start to appear in the solution. This indicates dissolution of radioactive $CuSO_4(s)$.
On the other hand, if $CuSO_4(s)$ crystals were placed in radioactive $CuSO_4(aq)$ then radioactivity would start to appear in the crystals.
This indicates precipitation from radioactive $CuSO_4(aq)$.

Although macro properties are constant, change is always occurring at the microlevel and the equilibrium is called dynamic.

Video animations (videos you can watch on a computer screen) of equilibrium situations can be accessed through the chemistry section, water module of http://www.lmpc.edu.au/science.

Observing a reversible change

Take two ice blocks from a freezer. Leave them at room temperature for about a minute to start melting. Place one ice block on top of the other. If they are awkward shapes try to put two flat surfaces together and keep the two blocks in contact using a rubber band around them. Leave them for three minutes for more melting $H_2O(s) \rightarrow H_2O(l)$ to occur.

1 After three minutes hold the top ice block and lift it up (remove the rubber band first if you used one). Is there any evidence that the reverse process, freezing of water, has occurred?

2 Write an equation to represent this reverse process at the symbolic level.

3 Combine the equations for melting and freezing into a single equation using reversible arrows.

4 Although what is happening can be represented using reverse arrows equilibrium for this system only occurs at a certain temperature. What is the temperature at which the rate of melting = rate of freezing?

5 If the ice blocks are kept at room temperature for 30 minutes

a) what happens to the rate of melting?

b) what happens to the rate of freezing?

c) is this now an equilibrium situation?

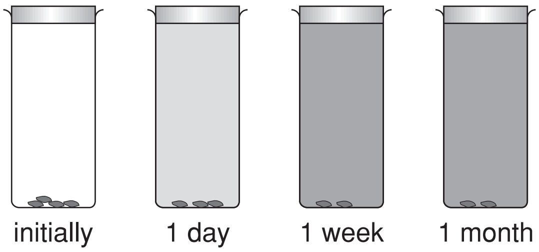
Check your answers.

Reversible reactions that eventually reach equilibrium

Iodine sublimes

Dark crystals of the element iodine $I_2(s)$ are placed in a glass tube with a plastic lid. At room temperature iodine sublimes, that is, changes directly from solid to gas without going through the liquid phase:

$I_2(s) \rightarrow I_2(g)$



Equilibrium is reached when the concentration of violet iodine vapour $I_2(g)$ is constant. This is when the intensity of colour is constant.

1 Complete this equation with reversed arrows to show the reversible reaction $I_2(s) \rightleftharpoons$

2 Using the diagram estimate how long the system took to reach equilibrium. Explain why you came to this conclusion.

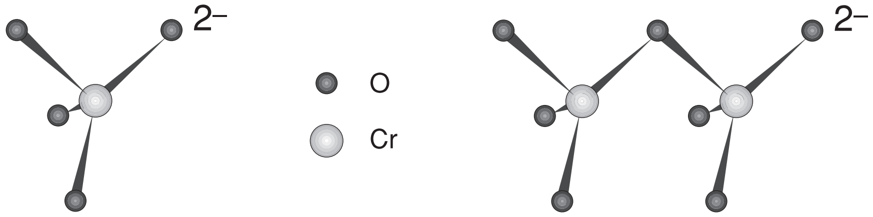
Check your answers.

The sealed container is called a **closed system** – a system that does not allow matter to enter or leave. A closed system is required for equilibrium to be established.

Chromate/dichromate salts

Chromate salts such as potassium chromate K_2CrO_4 are yellow while dichromate salts such as potassium dichromate $K_2Cr_2O_7$ are orange.

Because most potassium salts are colourless (their crystals look white because of reflected light from interfaces) potassium ions K^+ must be colourless. Thus chromate ions CrO_4^{2-} must be yellow and dichromate ions $Cr_2O_7^{2-}$ must be orange. In these ions the oxygens atoms are covalently bonded to the chromium in a tetrahedral arrangement.



chromate ion

dichromate ion

If a chromate solution is prepared (for example by dissolving K_2CrO_4 in water) it is yellow. If some acid solution (a solution containing hydrogen ions) is added the yellow dichromate solution turns orange-yellow.
If more acid is added, it turns orange indicating the presence of dichromate ions.

$$\underset{\text{yellow}}{CrO_4^{2-}(aq)} + H^+(aq) \rightarrow \underset{\text{orange}}{Cr_2O_7^{2-}(aq)} + H_2O(l)$$

Balance the ionic equation above by placing numbers, where needed, in front of the formulas. The equation is balanced when:

- there are the same number of each type of atom on both sides
- the total charge on the LHS equals the total charge on the RHS.

(Hint: you only need to add numbers to the LHS)

Check your answer.

If a dichromate solution is prepared (for example by dissolving $K_2Cr_2O_7$ in water) it is orange. If some alkaline solution (a solution containing hydroxide ions) is added the solution turns yellowish-orange. If more alkaline solution is added it turns yellow indicating the presence of chromate ions.

$$\underset{\text{orange}}{Cr_2O_7^{2-}(aq)} + 2OH^-(aq) \rightarrow \underset{\text{yellow}}{2CrO_4^{2-}(aq)} + H_2O(l)$$

The changes that occur can be summarised in a single equation.

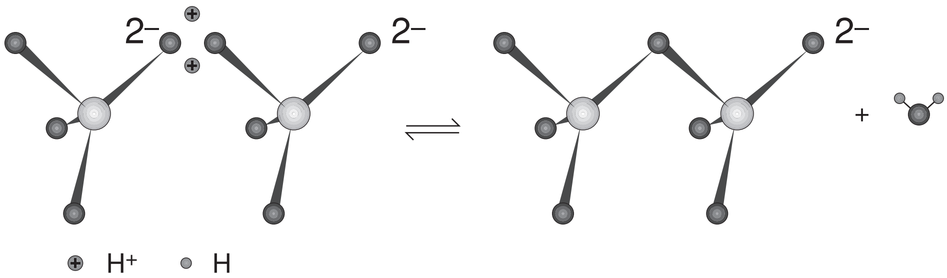
Acidic solution (presence of H^+) drives the reaction to the right of this equation.

Alkaline solution (contains OH^- which react with and remove H^+) drives the reaction to the left of this equation.

acidic → ← alkaline

$$2CrO_4^{2-}(aq) + 2H^+(aq) \rightleftharpoons Cr_2O_7^{2-}(aq) + H_2O(l)$$

yellow orange



Ions dissolved in water solution can be in a stable equilibrium because they are in a closed system. The ions do not escape into the air because they are attracted to and held in the aqueous liquid phase by the polar water molecules. Water provides a closed system for reactions involving ions. The motion of the water molecules brings ions together for reaction.

Characteristics of a system at equilibrium

1 The system is closed – it does not exchange matter with its surroundings. For example a sealed container for an equilibrium involving gas or water for an equilibrium involving mobile ions.

2 The system is dynamic. Macro level properties (what you can observe) are not changing (constant) but change is occurring at the micro level (what you can visualise is happening to ions and molecules).

3 The system can be approached from either direction. For example a yellow-orange solution of chromate and dichromate ions can be produce either by starting with yellow chromate solution and adding acidic solution or by starting with orange dichromate solution and adding alkaline solution.

Explain why a beaker containing magnesium chloride $MgCl_2$ crystals in contact with a saturated solution of magnesium chloride is a system at equilibrium.

Note that writing down that this system is closed, dynamic and can be approached from either direction is not explaining – it is demonstrating knowledge. Explain means relate cause and effect; make the relationships between things evident; provide why and/or how – it is demonstrating understanding as well as knowledge.

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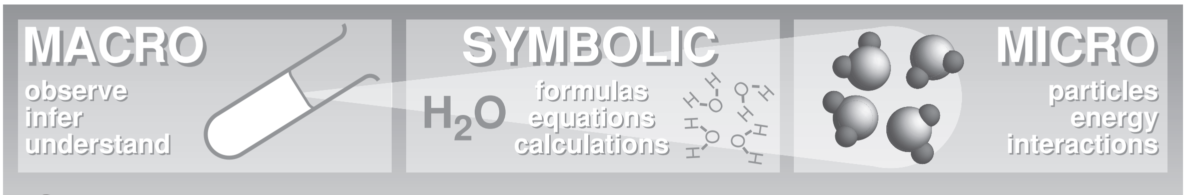
Compare your answer to the suggested answer.



Salt solutions

When chemists prepare a solution containing a particular anion (negative ion) they look in their chemical store for an ammonium (NH_4^+) or group one (Li, Na, K, Rb or Cs) salt containing that anion. All ammonium and group one salts are soluble.

Similarly when a chemist wants to prepare a solution containing a particular cation (positive ion) they look for a nitrate or acetate (also known as ethanoate) salt containing that cation. All nitrate and acetate salts are soluble.

Comparing solubilities using precipitation reactions

In this activity you will use micro-techniques involving small quantities of salt solutions. You will compare the water solubilities of various salts through precipitation reactions. If a cation and an anion precipitate, the salt formed by them is insoluble. For example whenever two separate clear solutions, containing silver ions Ag^+ and chloride ions Cl^- are mixed, a white precipitate forms; therefore AgCl is insoluble.

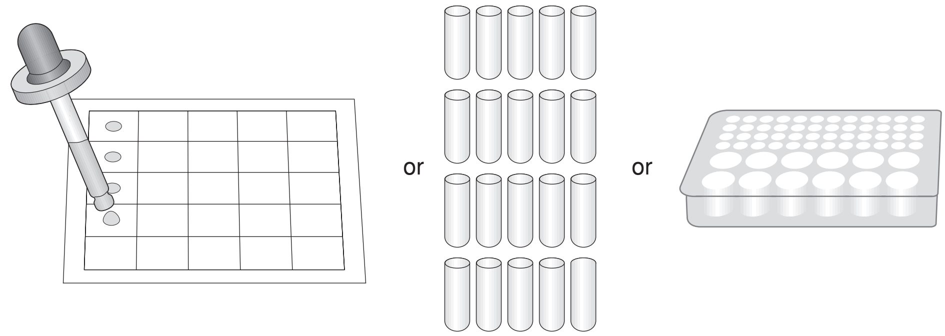
What you will need:

- 9 dropper solution bottles containing soluble salt solutions (5% salt by mass) separately containing the cations Ba^{2+}, Ca^{2+}, Fe^{2+} Cu^{2+}, Pb^{2+} and the anions Cl^-, SO_4^{2-}, CO_3^{2-}, OH^-.

Make sure you wash your hands well after handling these solutions especially the lead solution. All the ions used are soluble and therefore easily absorbed.

- A flat plastic sheet such as an overhead projector sheet or 20 small test tubes or a plastic well plate. This will be used to mix a drop of cation solution and a drop of anion solution.

If you use a flat plastic sheet remove the copy of the cation/anion table from the *Appendix* at the back of this part. Place the table underneath the plastic sheet.



- a dark surface on which the sheet/test tubes/well plate can be placed for viewing.

What you will do:

1. Record the colours of each cation and anion under their formula in the results table on the next page
2. Use the results table to systematically arrange four drops of each cation in rows. Then five drops of each anion in columns so that each drop of anion solution. is added to a drop of cation solution.
3. Be careful not to contaminate the original solutions or the droppers. Make sure the dropper tips do not come in contact with any other solution and that the droppers are returned to their correct containers. Use one dropper at a time and return it to its bottle before taking another dropper.
4. Record the colour of any precipitate that forms in the appropriate square. Salts producing precipitates are insoluble.
5. If no precipitate forms record – in the square for the soluble salt.

Results:

	chloride Cl^-	sulfate SO_4^{2-}	carbonate CO_3^{2-}	hydroxide OH^-
barium Ba^{2+}				
calcium Ca^{2+}				
iron(II) Fe^{2+}				
copper Cu^{2+}				
lead Pb^{2+}				

Conclusions:

Use soluble or insoluble to complete these sentences:

Most chlorides are ____________.

Most carbonates are ____________.

Predicting reactions from solubility data

If a larger number of cation and anion combinations are tested for precipitation the following generalisations about water solubilities can be reached:

Anion	**Cation**	
	soluble (no reaction)	insoluble (precipitate forms)
nitrates NO_3^-	all	
acetates CH_3COO^-	all	
chlorides Cl^-	most	Ag^+ (Pb^{2+})*
sulfates SO_4^{2-}	most	Ba^{2+},(Ca^{2+})#,Pb^{2+},(Ag^+)#
sulfides S^{2-}	Group 1, NH_4^+, Group 2	most
hydroxides OH^-	Group 1, NH_4^+, Ba^{2+}	most
carbonates CO_3^{2-}	Group 1, NH_4^+	most

Table of salt solubilities.

Notes: * $PbCl_2$ is insoluble in cold water, but soluble in hot water

$CaSO_4$ and Ag_2SO_4 *are borderline;* precipitate will probably form if concentrated solutions at low temperature are used.

If two salt solutions are mixed together and no precipitate can form then no reaction occurs.

eg. sodium nitrate solution + calcium chloride solution $\rightarrow$ no reaction

$$Na^+(aq) + NO_3^-(aq) + Ca^{2+}(aq) + 2\,Cl^-(aq) \rightarrow \text{mixture of ions}$$

The solutions that are started with are solutions of soluble salts.

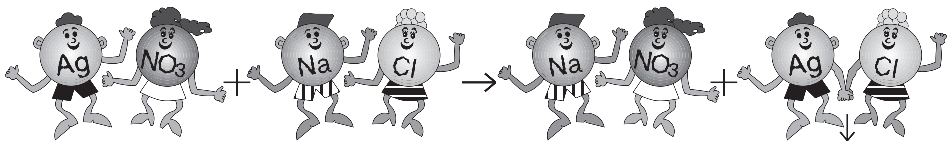
The sodium ions and chloride ions coming in contact in the mixture will not precipitate nor will the calcium ions and nitrate ions as both sodium chloride and calcium nitrate are soluble salts.

If two salt solutions are mixed together and any cation and anion combination present forms a precipitate then reaction occurs.

For example, silver nitrate solution and sodium chloride solution will react because silver ions and chloride ions form a precipitate. As soon as the two solutions are mixed a white precipitate falls out of the mixture. In some textbooks the precipitate is followed by a ↓ to indicate that it falls out of solution.

$Ag^{+}(aq) + NO_3^{-}(aq) + Na^{+}(aq) + Cl^{-}(aq) \rightarrow Na^{+}(aq) + NO_3^{-}(aq) + AgCl(s)\downarrow$

The rapid formation of an insoluble solid from clear solutions causes a change that is easily seen. The diagram below shows a student's **analogy** for this precipitation reaction:



Use the table of salt solubilities to predict if a precipitate forms when the following solutions are mixed. Write the name of any precipitate you think forms.

1 barium chloride + sodium sulfate ___________________________

2 barium chloride + potassium hydroxide ______________________

3 calcium nitrate + ammonium sulfate_________________________

4 copper sulfate + ammonium sulfide _________________________

5 lead acetate + sodium carbonate ____________________________

Check your answers.

When you write an equation for a reaction that forms a precipitate there is no need to show spectator ions. Spectator ions are ions that take no part in the reaction. Spectator ions float around in solution before mixing and they keep on floating around after mixing.

Look at the diagram of the student's analogy of a precipitation reaction. Name the spectator ions in this reaction.

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Check your answers.

The ions that come together on mixing and form ionic bonds are the reaction ions. They are the only ions that must be shown in a **net ionic equation**. For example, the net ionic equation below summarises the reaction between silver nitrate and sodium chloride.

eg. $Ag^{+}(aq) + Cl^{-}(aq) \rightarrow AgCl(s)$

A disadvantage of using net ionic equations is that they do not show that nitrate ions and sodium ions were present.

However, a major advantage is that this type of equation summarises that any silver salt solution will react with any chloride salt solution. Knowing that AgCl is insoluble enables you to predict, for example, that silver acetate solution will react with potassium chloride solution. The attraction between Ag^{+} and Cl^{-} is stronger than between these ions and water molecules. Thus a precipitate forms.

Summarise the reactions in the previous exercise by writing net ionic equations (showing only reaction ions). Do not show any spectator ions.

1 ____

2 ____

3 ____

4 ____

5 ____

Check your answers.

Double precipitate reactions

Sometimes two different precipitates will form when two solutions are mixed. For example, when barium hydroxide solution and magnesium sulfate solution are mixed barium sulfate and magnesium hydroxide precipitate.

$$Ba^{2+}(aq) + 2OH^{-}(aq) + Mg^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s) + Mg(OH)_2(s)$$

Concentration and solubility

Whether a precipitate forms or not will depend on the concentration of ions that are mixed.

The practical activity *Comparing solubilities using precipitation reactions* recommended that you use 5% by mass salt solutions. If you had used 1% by mass solutions you probably would not have seen a precipitate of $PbCl_2$. Salts vary enormously in their solubilities.
For example, $AgNO_3$ is about one hundred million (10^8) times more water soluble than AgCl.

Concentration units

There are many different ways of expressing concentration. Always note and record the concentration unit as well as the numerical value of a concentration measurement.

A mouthful of sodium chloride solution of 5 mol L^{-1} concentration would probably make you sick if you drank it while a mouthful of sodium chloride of concentration 5% would taste like sea water. You would not be able to taste the salt in a solution of 5 ppm concentration.

In this section you will learn more about mol L^{-1}, %, ppm and other concentration units.

Percentage by mass

Percentage by mass is extensively used in industry.

$$\text{Percentage by mass} = \frac{\text{mass}_{\text{solute}}}{\text{mass}_{\text{solution}}} \times \frac{100}{1}\% = \frac{m}{m}\% = \%\frac{m}{m}$$

For chemists and industry this is the same as:

$$\text{Percentage by weight} = \frac{\text{weight}_{\text{solute}}}{\text{weight}_{\text{solution}}} \times \frac{100}{1}\% = \frac{w}{w}\% = \%\frac{w}{w}$$

For example 5 g of NaCl dissolved in 95 g of water to give 100 g of NaCl solution is 5% (m/m) or 5% (w/w) NaCl solution.

What would you mix together to prepare 100 g of a 10% (m/m) solution of aqueous magnesium sulfate?

Check your answer.

If a solution concentration is % written by itself it usually means % (m/m) for a solid dissolved in a liquid. If however the solution is of two liquids eg. a 5% alcohol drink the % usually refers to volumes (v/v).

Percentage volume

Percentage by volume is used for mixtures of liquids eg alcoholic drinks

$$\text{Percentage by volume} = \frac{\text{volume}_{\text{solute}}}{\text{volume}_{\text{solution}}} \times \frac{100}{1}\% = \frac{v}{v}\% = \%\frac{v}{v}$$

For example 5% (v/v) aqueous ethanol consists of 5 mL of ethanol per 100 mL of aqueous ethanol solution. Such a solution is made from about 5 mL of ethanol and about 95 mL of water.

A wine is labelled as 12% (v/v) alcohol. What volume of alcohol is present in each 750 mL bottle of the wine?

Check your answer.

Mass per volume

In pharmacy and medicine, mass per volume is normally used eg mass per volume of solution (m/v) is used to measure the blood alcohol level of drivers.

A blood alcohol level of 0.02 refers to 0.020 g/100 mL of blood.

An 18 year old has about 5 L of blood in her body. If she had a blood alcohol level of 0.020 g/100mL

a) What total mass of alcohol is in her blood?

b) Alcohol has a density of 0.78 g cm^{-3}. Change the total mass of alcohol in her blood to a volume. What is her blood alcohol level in mL/100mL?

Use density = $\frac{\text{mass}}{\text{volume}}$ $\quad D = \frac{m}{v}$ $\quad$ in your calculation

a)

b)

Check your answers.

Parts per million

Parts per million (ppm) or parts per billion (ppb) are often used to measure environmental pollutants.

The ppm or ppb figures can be m/m, comparing masses (usually used for solids in solids) or v/v, comparing volumes (usually used for gases in gases). These ppm or ppb figures are ratios because they compare the same mass units or the same volume units.

Water pollution measurements are often reported in milligrams per litre (mg/L) or micrograms per litre μg/L. mg/L is usually considered equivalent to ppm and μg/L equivalent to ppb. In water one litre has a mass close to one kilogram thus:

$$1 \text{ mg/L} = 1\text{mg/kg} = 10^{-3} \text{ g}/10^{3} \text{ g} = 10^{-3}/10^{3} = 1/10^{6} = \frac{1\,part}{1\,million} = 1 \text{ ppm}$$

1 ppm is about the concentration of one table salt crystal dissolved in a cup of water.

$$1 \text{ μg/L} = 1\text{μg/kg} = 10^{-6} \text{ g}/10^{3} \text{ g} = 10^{-6}/10^{3} = 1/10^{9} = \frac{1\,part}{1\,billion} = 1 \text{ ppb}$$

1 ppb is about the concentration of one table salt crystal dissolved in a bath of water.

$$1\,ppm = \frac{1}{1000000} = \frac{1000}{1000000000} = \frac{1\,thousand}{1\,billion} = 1000\,ppb$$

At 15°C fresh water can contain up to 10 mg/L = 10 ppm of dissolved oxygen gas. Most people find 10 ppm easier to remember than 10 mg/L or 0.010 g/L.

The NHMRC Australian guidelines recommend less than 300 micrograms of iron per litre in drinking water. Change 300 μg/L to ppm and ppb figures.

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Check your answers.

Molarity

Molarity is Chemistry's main unit of concentration.

When chemists are using solutions measured in molarity it is easier to relate the quantities used (macro level) to number of particles (micro level) and reaction equations (symbolic level). The molarity of a solution is the number of moles of solute per litre of solution.

$$c = \frac{n}{V}$$

c = concentration in moles per litre (M or mol L^{-1})

n = number of moles of solute

V = volume of the solution in litres.

Chemistry textbooks of different ages and countries of origin show molarity units as **molar** (M), moles per litre (mol L^{-1}), moles per cubic decimetre (mol dm^{-3}) or kilomoles per cubic metre (kmol m^{-3}).
All these units are equivalent.

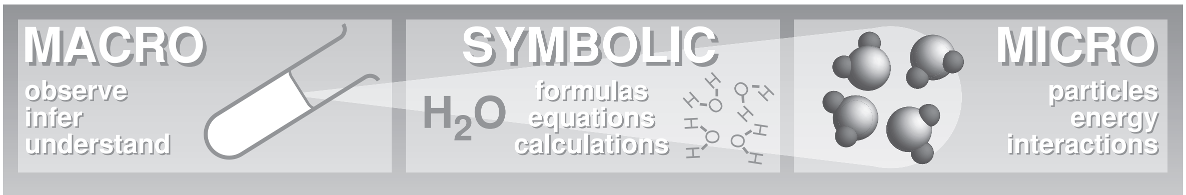
$1\ M = 1\ mol\ L^{-1} = 1\ mol\ dm^{-3} = 1\ kmol\ m^{-3}$

Pathology measurements of chemicals in blood, plasma, serum, urine and other body fluids are often in millimoles per litre (mmol L^{-1}) or micromoles per litre (μmol L^{-1}).

If a person is taking medication to reduce blood cholesterol below 5 it means that the aim is to get the concentration of cholesterol below 5 millimoles per litre of blood.

Blood glucose levels are also typically about 5 millimoles per litre (unless a person is diabetic and the blood glucose level varies a lot).

Concentrations of vitamins in the blood are much lower and normally measured in micromoles per litre. Vitamin A blood concentration is normally about 1 μmol L^{-1}.



Preparing solutions

Procedure to prepare 1.00L of a 2.00 mol L^{-1} solution of sodium chloride is as follows:

1 Calculation

   $n = c \times V = 2.00 \text{ mol } L^{-1} \times 1.00L = 2.00 \text{ mol}$
   mol NaCl = 2.00 x (22.99+35.45) = 116.88 g

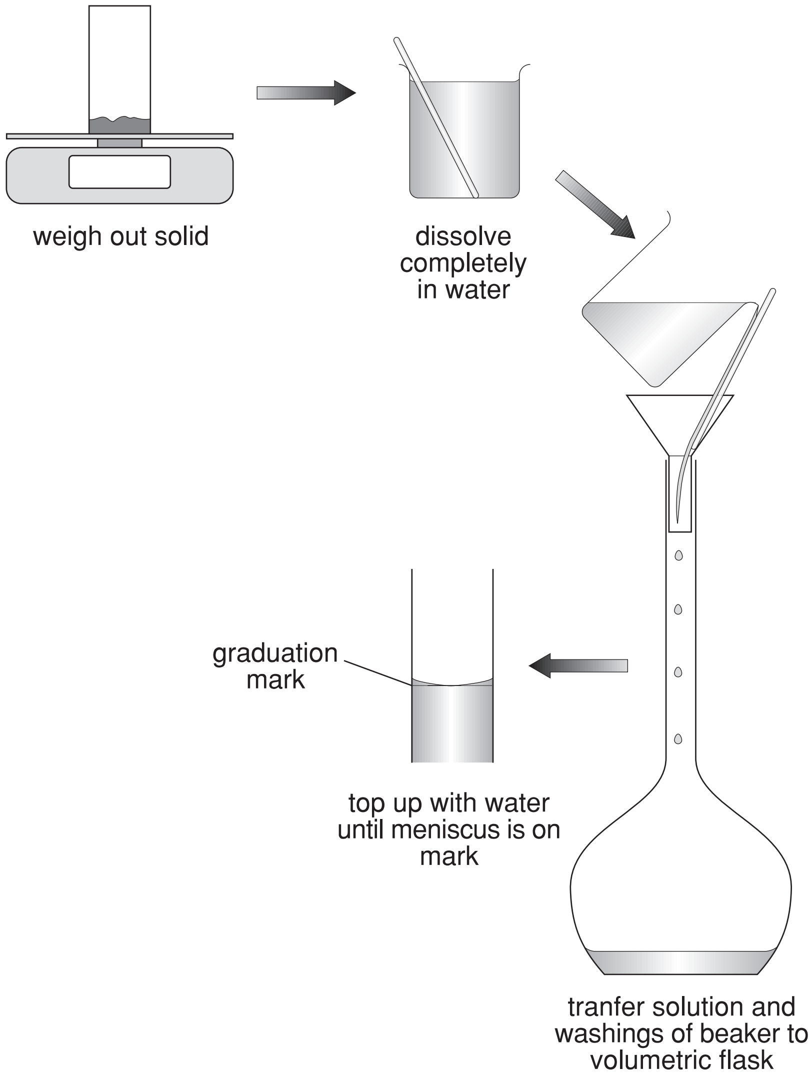
2 Weigh out 117 g of NaCl

3 Dissolve NaCl in some water

4 Make the solution up to a total solution volume of 2.00 L.

Note that in preparing solutions measurements are usually to three significant figures. Later in your course when carrying out quantitative analysis using volumes of known concentration (**volumetric analysis**) you will make all measurements to at least three significant figures.

Note also that all concentration units are for a certain volume of a *solution* not a certain volume of solvent. When a solution is prepared the solute is dissolved in a small amount of solvent and then more solvent added until the total volume of solution is obtained. The volume of solution prepared can be accurately measured using a volumetric flask.



The last few drops of water are added by a dropper so that the bottom of the meniscus lies on the level line marked on the neck of the volumetric flask.

Describe how you would prepare 0.500 L of 4.00 mol L^{-1} of $CaCl_2$.

Molar mass of $CaCl_2$ is 111.0 g.

____

____

____

Check your answer.

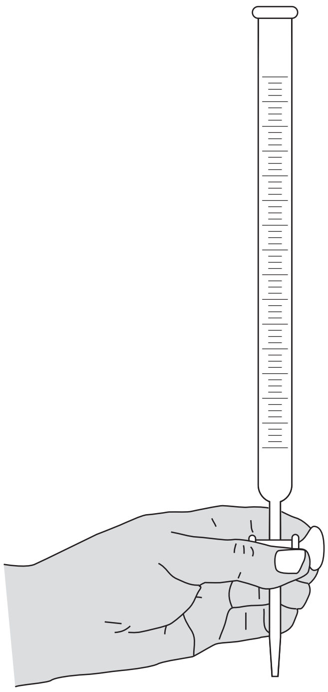
Prepared solutions are a convenient and quick way of handling different amounts of a chemical. It is much quicker to measure out a volume of a prepared solution of a chemical than to have to weigh out the chemical.

Suppose you needed to measure out an amount of NaCl, twice that amount, three times that amount and then four times that amount. If you weighed out the amount four weighings would be needed.

The advantages of using a prepared solution of known concentration are:

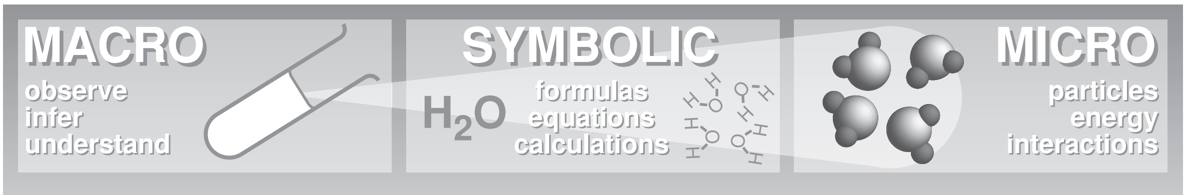
- pouring out a calculated volume, followed by twice, three times and four times that volume would be much quicker than weighing four times
- a chemical in solution has particles that are moved around by the solvent and so the chemical solution will mix and react more readily.

A burette is used to transfer different volumes of a solution of known concentration.



Note that a burette tap is operated with the left thumb in front and fingers behind.

Before use the burette is washed with some of the solution of known concentration, drained and then filled with solution. Also before using the burette to measure volumes the tap needs to be opened to clear any air bubbles trapped around the tap or in the tip.



Diluting solutions

Diluting a solution involves:

- taking an initial volume (V_i) of a solution containing a fixed number of moles of solute *n*
- adding water until a final, larger volume V_f is obtained containing the same number of moles of solute *n*.

  initial number of moles n = final number of moles n

Because $c = \frac{n}{V}$, then n = c x V

initial n = final n

Thus, initial $c_i \times V_i$ = final $c_f \times V_f$

This can be expressed as cV = constant or $c_1V_1 = c_2V_2 = c_3V_3$ and so on for a series of dilutions.

Example: Calculate the concentration of a solution formed by the dilution of 50 mL of 2.0 mol L^{-1} to a new volume of 400 mL.

$c_1V_2 = c_2V_2$

$2.0 \times 50 = c_2 \times 400$

$$c_2 = 2.0 \times \frac{50}{400} = \frac{100}{400} = 0.25$$

Concentration of the diluted solution = 0.25 M

Note that any units can be used for *c* and *V* provided the same unit is used on both sides of the equation.

cV is only equal to the number of moles *n* when *c* is in mol L^{-1} and *V* is in litres (L).

1 Calculate the volume of solution prepared when 250 mL of 0.4 mol L^{-1} NaCl is diluted to produce a concentration of 0.1 mol L^{-1}.

___

___

2 If 2.0 L of 0.100 mol L^{-1} $MgSO_4$ was prepared from 220 mL of a bulk solution what was the concentration of the bulk solution?

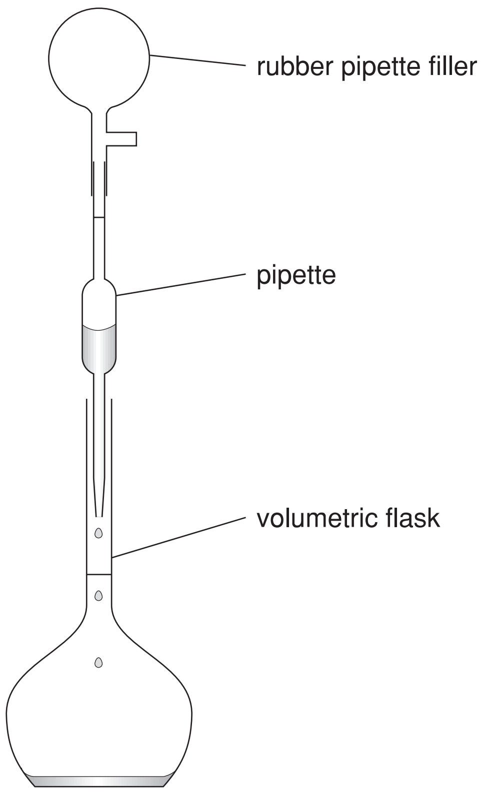
___

___

___

Check your answers.

A pipette and a volumetric flask can be used to dilute a solution.



If a 25 mL pipette and a 250 mL volumetric flask are used correctly the diluted solution will have 1/10 the concentration of the original solution.

Preparing solutions in the kitchen

In this activity you will prepare solutions to specified v/v and m/v specifications and dilute them to specified concentrations. Data you record in this activity will be used in Exercise 4.2.

You will need:

- cordial or fruit juice for dilution
- sugar cubes in a packet with information on the weight of each cube or the total weight and the number of sugar cubes
- containers and volume measuring equipment in the kitchen
- iceblocks, preferably cube shaped; keep these frozen until ready for the fifth part of this activity.

Method and results:

1 Most cordials or fruit juices are concentrated and prepared for drinking by adding 1 volume of cordial/fruit juice to 4 volumes of water. Prepare a drink in this way and you have prepared a solution to a v/v specification. Taste some of the original fruit juice/cordial and your diluted cordial/fruit juice and record their tastes:

____

____

Keep the remaining v/v solution.

2 Measure the edge of a sugar cube(l): ____cm

Calculate the volume of the cube(l^3): ____cm^3

Measure or calculate (from total weight and number of cubes in packet) the mass of a sugar cube: ____g

Measure a fixed volume (say 50 mL = 50 cm^3) of water into a transparent container. Mark the top of the water level with a marker pen. Volume of water used:____cm^3 = ____mL

Add a sugar cube and describe what happens as it dissolves:

____

____

____

____

Noting the number used, add sugar cubes with stirring until no more will dissolve. Notice how the viscosity of the solution increases.

Number of sugar cubes required to saturate ____ cm^3 of water = ___

Measure the total volume of the saturated solution: ____ mL

Concentration of saturated sugar solution = ____ g/____mL

Convert this concentration to g/L:____________________________

Keep this m/v solution.

3 Diluting the diluted cordial/fruit juice solution to one tenth of its concentration.

Estimate the volume of water ____mL to be added to a fixed volume of cordial/fruit juice solution____mL. Add these volumes and measure the total volume of diluted solution____mL.

Taste some of this product and record its taste:_________________

4 Diluting the sugar solution to half its original concentration

Estimate the volume of water ____mL to be added to a fixed volume of sugar solution____mL. Add these volumes and measure the total volume of diluted solution____mL.

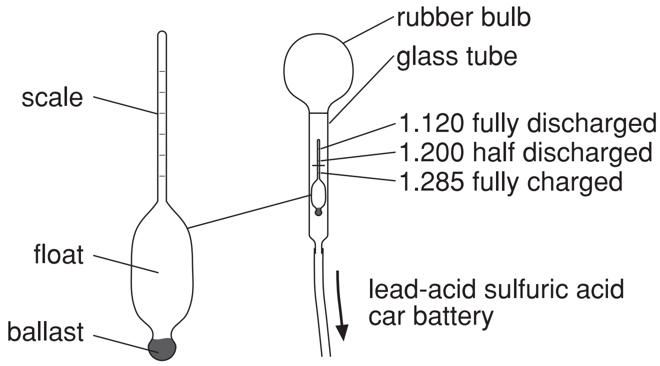
Do Exercise 4.1: *Using flotation to estimate density* after carrying out the activity below.

Add an iceblock to each of the solutions you prepared in the four parts of this activity. Estimate the percentage of each iceblock that is below the liquid level for:

- 20% v/v cordial/fruit juice
- saturated sugar solution
- 2% v/v cordial/fruit juice
- diluted sugar solution.

Comparing the densities of solutions

In industry the densities of different solutions can be estimated by floating objects in them. Instruments called hydrometers draw a small amount of liquid into their body and then measure the density of the liquid by floating an object with scales in the liquid. The density and concentration can be read off the scales.



Density reading of battery acid can show the extent of discharge of the battery.

Hydrometers can be used to measure the concentrations of:

- sulfuric acid solution in lead-acid car batteries
- sugar solutions
- alcohol solutions.

MACRO
observe
infer
understand

SYMBOLIC
H_2O
formulas
equations
calculations

MICRO
particles
energy
interactions

Calculating mass and concentration

A group of students were asked to determine the solubility of lead chloride $PbCl_2$ in water. You will be asked to do some of the calculations they did and answer questions about the experiment.

This is the procedure followed and measurements made. Complete any calculations and sentences where spaces exist.

100 mL solution of 1.00 mol L^{-1} NaCl solution was prepared.

1 Calculate the mass of NaCl required.

number of moles required n = c x V = 1.00 mol L^{-1} x ______ L

= 0.100 mol NaCl

= 0.100 mol x (22.99 + 35.45) g/mol

= ______ g

100 mL solution of 0.500 mol L^{-1} $Pb(NO_3)_2$ solution was prepared

$Pb(NO_3)_2(s) \rightarrow Pb^{2+}(aq) + 2NO_3^-(aq)$

Each unit of $Pb(NO_3)_2$ releases one Pb^{2+} and two NO_3^-. The concentration of Pb^{2+} will be the same as $Pb(NO_3)_2$, that is _____ mol L^{-1} while the concentration of NO_3^- will be twice as great, at 1.00 mol L^{-1}.

2 Calculate the mass of $Pb(NO_3)_2$ required.

number of moles required, n = c x V = _______ mol L^{-1} x ______ L

= 0.0500 mol $Pb(NO_3)_2$

= ________ mol x (207.2 + 2 x 14.01 + 6 x 16.00) g/mol

= _______ g

The two clear solutions were mixed and immediately went cloudy white. A white precipitate of lead chloride had formed.

lead nitrate(aq) + sodium chloride(aq) → lead chloride(s) + sodium nitrate(aq)

$Pb(NO_3)_2 + 2NaCl \rightarrow PbCl_2(s) + 2\,NaNO_3$

The net ionic equation is: _________________________________

3 Brackets, [ ] are used to represent mol L^{-1} concentration

In the 1.00 mol L^{-1} NaCl solution $[Na^+] = 1.00$ $[Cl^-] =$ _____

In the 0.500 mol L^{-1} $Pb(NO_3)_2$ solution $[Pb^{2+}] = 0.500$ $[NO_3^-] =$ _____

The mixture consisted of solid $PbCl_2$ precipitate and 200 mL of solution containing $[Na^+] = 0.500$ and $[NO_3^-] = 0.500$.

4 Explain why the $[Na^+]$ and $[NO_3^-]$ are now half of their original concentrations

____

____

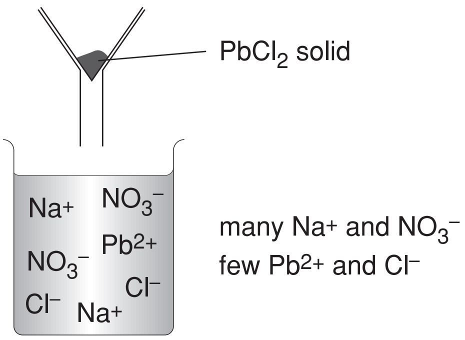
____

____

____

____

The mixture was stirred and filtered through filter paper. The temperature of the filtered solution (filtrate) was measured and found to be 25°C. Because this solution had been in contact with solid lead chloride it was saturated with Pb^{2+} ions and Cl^- ions. As well as being a solution of $NaNO_3$ made up of the spectator ions that had not reacted the solution also contained low concentrations of Pb^{2+} and Cl^- ions.



After drying the precipitate of $PbCl_2$ was weighed and found to weigh 11.25g.

The original 100 mL of 0.500 mol L^{-1} $Pb(NO_3)_2$ solution contained 0.100 L x 0.500 mol L^{-1} = 0.0500 mol of Pb^{2+}.

The original 100 mL of 1.00 mol L^{-1} NaCl solution contained 0.100 L x 1.00 mol L^{-1} = 0.100 mol of Cl^-.

Applying the mole concept:

Pb^{2+}	+	$2\ Cl^{-}$	$\rightarrow$	$PbCl_2$
1 mol	+	2 mol	$\rightarrow$	1 mol
0.0500 mol	+	0.100 mol	$\rightarrow$	0.0500 mol

5 Calculate the mass of 0.0500 mol of $PbCl_2$

The students compared the mass of 0.0500 mol of $PbCl_2$ expected with the dry mass collected from the filter paper. The difference was equal to the lead chloride in the 200 mL of saturated solution in the filtrate. The difference was changed from g to mol and the mol L^{-1} concentration of the saturated $PbCl_2$ calculated.

6 Using the answer to question 5 and 11.25 g for the weight of dry $PbCl_2$ precipitate calculate the mol L^{-1} concentration of saturated $PbCl_2$ solution.

Check your answers.

MACRO observe infer understand | SYMBOLIC H_2O formulas equations calculations | MICRO particles energy interactions

Suggested answers

Dissolution and precipitation

After dissolution of the magnesium sulfate in water the magnesium sulfate solutions were drawn up the cord/rope/thick string by capillary action between the strands. Evaporation of water from the solution increased the concentration until the solution was saturated. The solute precipitated out of the saturated solution forming a stalactite. Solution falling to the ground evaporated forming a stalagmite. Sometimes the stalactite and stalagmite joined forming a connecting column.

Saturated solutions

$CaCO_3(s) \rightleftharpoons Ca^{2+}(aq) + CO_3^{2-}(aq)$

Observing a reversible change

1 The bottom ice block is joined to the top one showing that freezing of melted water occurred between the two surfaces.

Write an equation to represent this reverse process at the symbolic level.

2 $H_2O(l) \rightarrow H_2O(g)$

Combine the equations for melting and freezing into a single equation using reversible arrows.

3 $H_2O(s) \rightleftharpoons H_2O(l)$

4 0°C

5 a) The rate of melting increases then decreases.

b) The rate of freezing increases then decreases.

c) This is probably not an equilibrium at 30 minutes but at some time before this the rates were probably equal.

Iodine sublimes

1) $I_2(s) \rightleftharpoons I_2(g)$

2) The system appears to have reached equilibrium between one day and one week. The intensity of colour of $I_2(g)$ in the container is the same for one week and one month indicating that equilibrium was reached between one day and one week.

$$2CrO_4^{2-}(aq) + 2H^+(aq) \rightarrow Cr_2O_7^{2-}(aq) + H_2O(l)$$

Characteristics of a system at equilibrium

Some aqueous magnesium ions and chloride ions coming in contact with the crystals will precipitate onto the crystal surface. At the same time, at another part of the crystal surface magnesium ions and chloride ions will dissolve in the saturated solution. The rate at which precipitation occurs will be the same as the rate at which dissolution is occurring.
This maintains the concentration of the saturated solution and keeps the amount of solid constant. Equilibrium exists because the rates of forward and reverse reactions are equal.

Precipitation reactions

	chloride Cl^- colourless	sulfate SO_4^{2-} colourless	carbonate CO_3^{2-} colourless	hydroxide OH^- colourless
barium Ba^{2+} colourless	–	white	white	–
calcium Ca^{2+} colourless	–	white	white	white
iron(II) Fe^{2+} light green	–	–	green grey	green
copper Cu^{2+} blue	–	–	green	blue
lead Pb^{2+} colourless	white	white	white	white

Most chlorides are *soluble.*

Most carbonates are *insoluble*.

Predicting reactions from solubility data

1 barium sulfate

2 no precipitate

3 calcium sulfate

4 copper sulfide

5 lead carbonate.

The spectator ions are sodium ions and nitrate ions.

1 $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

2 –

3 $Ca^{2+}(aq) + SO_4^{2}(aq) \rightarrow CaSO_4(s)$

4 $Cu^{2+}(aq) + S^{2-}(aq) \rightarrow CuS(s)$

5 $Pb^{2+}(aq) + CO_3^{2-}(aq) \rightarrow PbCO_3(s)$

Percentage by mass

10 g of magnesium sulfate salt plus 90 g of water gives 100 g of 10% m/m magnesium sulfate solution.

Percentage by volume

750 mL bottle of wine labelled as 12%(v/v) wine contains 12/100 of 750 = 90 mL of alcohol.

Mass per volume

a) 5 L of blood containing 0.020 g/100 mL contains

(5/0.100) x 0.020 = 1.0 g

b) $Volume = \dfrac{mass}{density}$ = 1.0/ 0.78 = 1.3 mL

v/v concentration = 1.3 mL/ 5 L = 0.26 mL/L =0.026 mL/100 mL.

Parts per million

300 μg/L = 300 ppb = 300/1000 ppm = 0.300 ppm.

Preparing solutions

1 Calculation $n = c \times V = 4.00 \times 0.500 = 2.00$ mol

2.00 mol $CaCl_2$ = 2.00 x 111.0 g = 222 g

2 Weigh out 222 g of $CaCl_2$

3 Dissolve $CaCl_2$ in some water

4 Make the solution up to a total solution volume of 0.500 L

Diluting solutions

1 $0.4 \times 250 = 0.1 \times V_2$ $V_2 = (0.4 \times 250)/0.1 = 1000$ mL

2 220 mL = 0.220 L $c_1 \times 0.220 = 0.100 \times 2.0$

$c_1 = (0.100 \times 2.0)/0.220 = 0.91$ mol L^{-1}

Calculating mass and concentration in a precipitation reaction

1 0.100 mol x (22.99 + 35.45) g/mol = <u>5.84</u> g

Each unit of $Pb(NO_3)_2$ releases one Pb^{2+} and two NO_3^- the concentration of Pb^{2+} will be the same as $Pb(NO_3)_2$, that is <u>0.500</u> mol L^{-1} while the concentration of NO_3^- will be twice as great, at 1.00 mol L^{-1}.

2 number of moles required n = c x V = <u>0.500</u> mol L^{-1} x <u>0.100</u> L

= 0.0500 mol $Pb(NO_3)_2$

<u>= 0.0500</u> mol x (207.2 + 2 x 14.01 + 6 x 16.00) g/mol

= <u>16.56</u> g

The net ionic equation is: <u>$Pb^{2+}(aq) + 2Cl^-(aq) \rightarrow PbCl_2(s)$</u>

3 In the 1.00 mol L^{-1} NaCl solution $[Na^+] = 1.00$ $[Cl^-] =$ <u>1.00</u>

In the 0.500 mol L^{-1} $Pb(NO_3)_2$ solution $[Pb^{2+}] = 0.500$ $[NO_3^-] =$ <u>1.00</u>

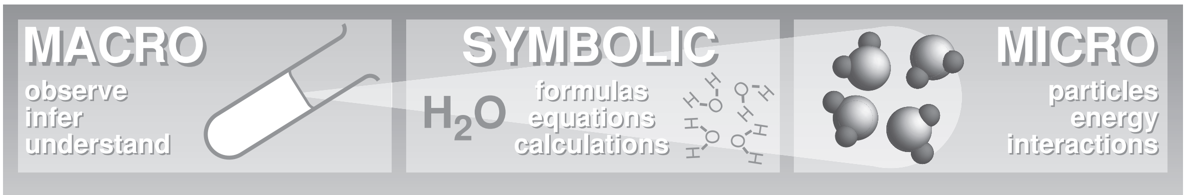
4 100 mL of solution containing Na^+ and 100 mL of solution containing NO_3^- were mixed so that each of the original concentrations was halved.

5 0.0500 mol x (207.2 + 2 x 35.45) g/mol = 13.91 g

6 13.91 – 11.25 = 2.66 g of $PbCl_2$ is dissolved in the 200 mL of saturated solution.

2.66 g = 2.66/(207.2 + 2 x 35.45) = 0.00956 mol

Molarity = 0.00956 mol/0.200 L = 0.0478 mol L^{-1}



Exercises Part 4

Exercises 4.1 to 4.2 Name: ________________________________

Exercise 4.1: Using flotation to estimate the density of solutions

In the activity on preparing solutions you were asked to add an iceblock to each of the solutions you prepared in the four parts of this activity. Estimate the percentage of each iceblock that is below the liquid level for:

- 20% v/v cordial/fruit juice
- saturated sugar solution
- 2% v/v cordial/fruit juice
- diluted sugar solution.

Draw a labelled diagram showing the appearance of an iceblock floating in each of these solutions. Include your percentage estimate of the amount of each iceblock that is below the liquid level.

If 50% of a uniform object is below the liquid level its density is 50% of the liquid's density. If 75% of a uniform object is below the liquid level its density is 75% of the liquid's density. If 90% of a uniform object is below the liquid level its density is 90% of the liquid's density. If 100% of the object is just below the liquid level its density is about the same as the liquid density. If the object sinks away from the liquid surface its density is greater than the liquid density.

The density of ice is 0.92 g cm^{-3}

a) Select appropriate **terminology** and **reporting style** to organise the information on the previous page. In other words, translate the information given in the diagrams you drew into wrting.

b) Calculate the density of each solution showing relevant working.

Exercise 4.2: Density, concentration and BP of an aqueous solution of sodium chloride

1.00 g of sodium chloride is dissolved in water to make up a solution of 100 mL. The weight of this solution is measured at 101 g.

a) Calculate the density of the solution

____

____

____

b) Calculate the concentration of sodium chloride in:

i) %m/m

____

____

____

ii) %m/v

____

____

____

iii) g/100 g water

____

____

____

iv) mol L^{-1}

____

____

____

v) ppm

____

____

____

c) If this solution was diluted to a total volume of 4.0 L what would be its new molar concentration?

____

____

____

d) What additional information would you need to calculate %v/v of the sodium chloride solution?

____

____

____

e) The temperature of boiling water is about 100°C and stays at this temperature until all the water boils away.

The temperature of a boiling sodium chloride solution is about 102°C initially but slowly increases as the boiling proceeds. Explain why the boiling temperature of the mixture rises as the water but not the sodium chloride boils off.

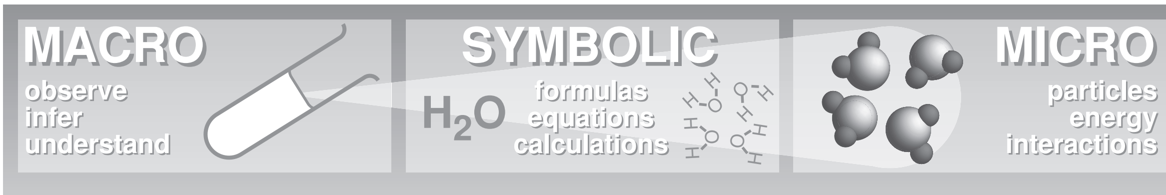
____

____

____

____

____



Appendix

	chloride Cl^-	sulfate SO_4^{2-}	carbonate CO_3^{2-}	hydroxide OH^-
barium Ba^{2+}				
calcium Ca^{2+}				
iron(II) Fe^{2+}				
copper Cu^{2+}				
lead Pb^{2+}				

Chemistry

Preliminary Course
Stage 6

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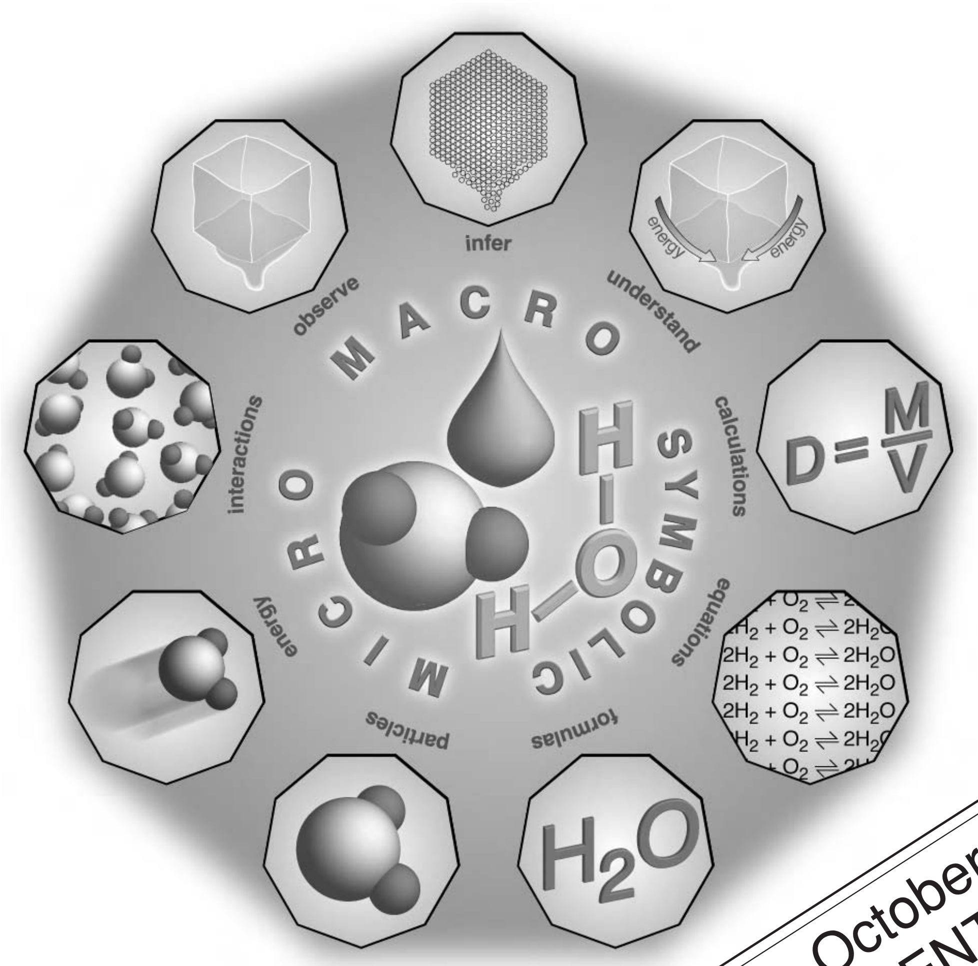
Water

Part 5: Water and heat



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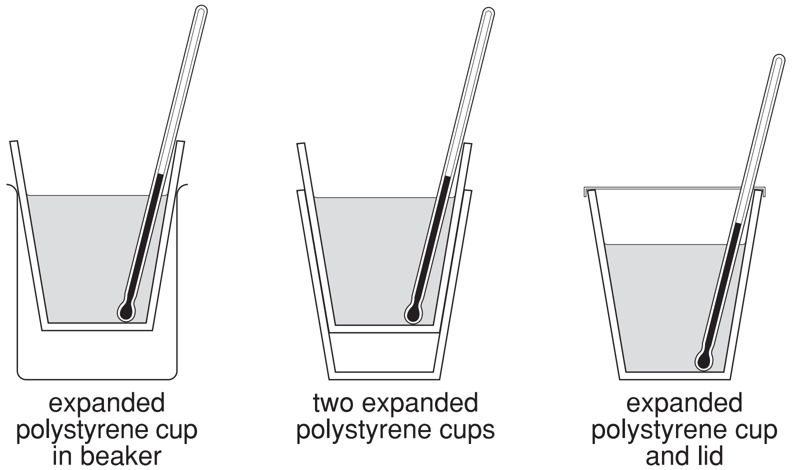
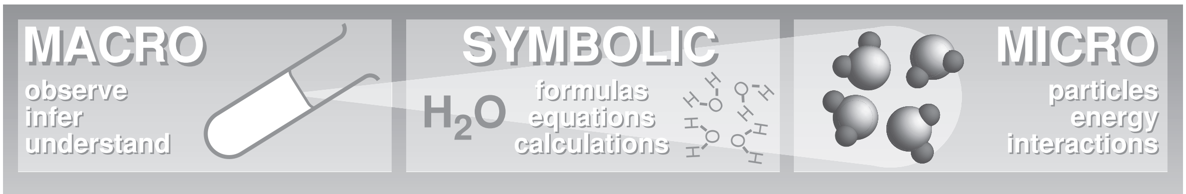

expanded
polystyrene cup
in beaker
two expanded
polystyrene cups
expanded
polystyrene cup
and lid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infer
understand
SYMBOLIC
H_2O
formulas
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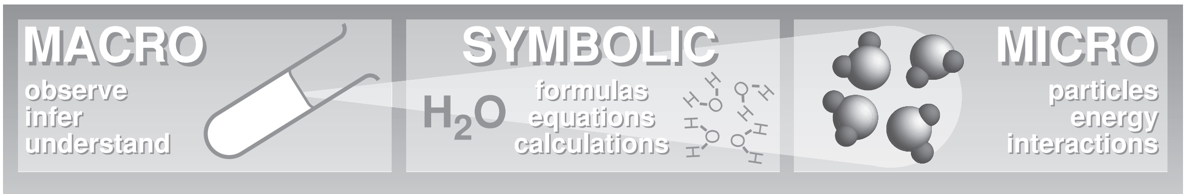
calculations
MICRO
particles
energy
interactions

Contents



Introduction

Water's ability to form many strong hydrogen bonds is the basis of its ability to absorb and store heat. This ability moderates the temperature range experienced by life.

A human body releases about 10 000 kJ of heat a day from chemical reactions. Most of this energy goes into breaking water molecules apart while still in the liquid phase rather than increasing the movement and hence temperature of the water molecules. This ability of water to absorb heat produces a stable human body temperature and minimises loss of the main chemical, water, from the body.

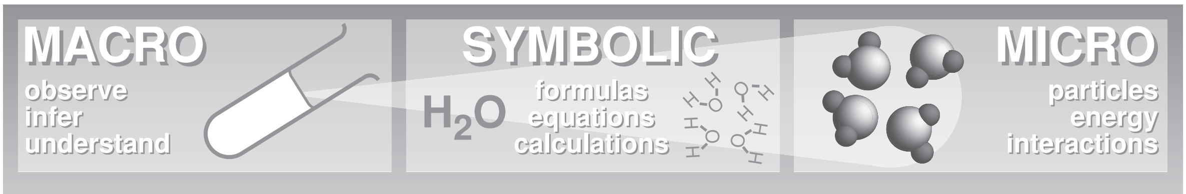
The narrow ranges of planet temperature and body temperature provided on Earth by water's ability to absorb heat have created conditions suitable for life on earth.

In Part 5 you will be given opportunities to learn to:

- explain what is meant by the specific heat of a substance
- compare the specific heat of water with a range of other solvents
- explain and use the equation $\Delta H = -m\,C\,\Delta T$
- explain how water's ability to absorb heat is used to measure energy changes in chemical reactions
- describe dissolutions which release heat as exothermic and give examples
- describe dissolutions which absorb heat as endothermic and give examples
- explain why water's ability to absorb heat is important to aquatic organisms and to life on earth generally
- explain what is meant by thermal pollution and discuss the implications for life if a body of water is affected by thermal pollution.

In Part 5 you will be given opportunities to:

- choose resources and perform a first-hand investigation to measure the change in temperature when substances dissolve in water and calculate the molar heat of solution
- process and present information from secondary sources to assess the limitations of calorimetry experiments and design modifications to equipment used.



Heat of vaporisation

The **heat of vaporisation** is the heat required to vaporise a liquid to gas. It can be expressed in kJ g^{-1} or kJ mol^{-1} and is usually measured at the BP of the liquid and 100 kPa pressure.

In this activity you will be measuring the heat of vaporisation of water in the kitchen.

You will need:

- an electric kettle
- a watch timing to seconds
- equipment to weigh water or measure the volume of water.

Method:

1 Empty the electric kettle of water. Lift up or remove the lid (this prevents a power cut-off device from turning the electric kettle off as soon as the water boils).

Record the power rating in watts (W) for the kettle (this could be on the bottom or back of the kettle). ________________ 1 W = 1J s^{-1}

2 Weigh 1000 g or measure out a volume of 1000 mL of water

3 Add 1000 g = 1000 mL of water to the electric kettle. Make sure the water level is at least 5 cm above the minimum level of water specified for using the kettle. If not then measure an additional amount of water and add it to the kettle. The water level must always cover the heating element.

4 Turn on the electric power. If the power rating is x watts then x joules of electrical energy changes to x joules of heat energy per second.

As the temperature rises, this energy is going into breaking hydrogen bonds between water molecules and increasing the kinetic energy of movement of the molecules. As the kinetic energy of water molecules increases, so does the temperature.

5 When you first see steam come out of the top of the kettle start timing with the watch for five minutes, then turn off the power. The energy supplied to the water during this five minutes is used to overcome forces of attraction in the liquid state and separate molecules so they are in the gaseous state. The energy is going into vaporisation, not raising the temperature of water.

6 Let the remaining water cool then measure the mass or volume of water remaining. Record the mass or volume __________.

Calculations:

1 Mass of water which vaporised from liquid to gas is equal to:
original mass – final mass = _____ – _____ = ____ g

2 Joules of energy used to vaporise ____ g of water

= ____ joules per second x (5 x 60) seconds = ________ J

3 Heat of vaporisation of water = _______ J / ____ g
= ____ $J\ g^{-1}$ = ____ $kJ\ g^{-1}$

4 Molar heat of vaporisation of water = ____ $J\ mol^{-1}$
= ____ $kJ\ mol^{-1}$

Discussion

1 How does your calculation of heat of vaporisation of water compare with the published value of 40.6 $kJ\ mol^{-1}$?

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2 Suggest reasons why your calculated value is different from the published value.

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3 When water is boiling at about 100°C the temperature does not rise until all the liquid water has changed to gaseous water. Why is this so?

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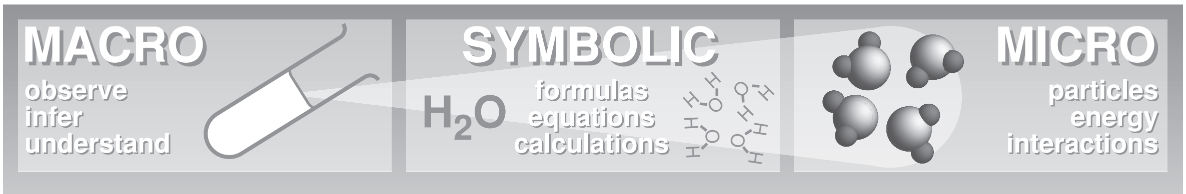
Check your answers.



Specific heat (capacity) C

The activity below requires two similar balloons and a box of matches. It should be carried out away from flammable materials and at a metal or porcelain sink with a water supply. You must not wear anything made of flammable material. Cotton and wool clothing or protective clothing such as a laboratory coat are best. If a balloon catches alight place it in the sink and spray it gently with water from the tap.

Water is the only liquid you may use in this experiment. Using any other liquid could cause a dangerous explosion. Read the whole activity through and take appropriate precautions before starting this activity.

1. Take two similar balloons.
2. Add about 25 mL of water to one balloon and blow it up with air. Tie it off.
3. Blow the other balloon up to the same extent with air. Tie it off.
4. Ask someone to hold both balloons about one metre above the sink. Make sure there are no flammable materials such as methylated spirits or curtains are nearby.
5. Light two matches and hold the top of the flames about 10 cm under the bottom of the balloons. Which balloon burst first?

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6. If neither balloon burst then light another two matches and use them about 5 cm below each balloon.

Comment on the ability of water to absorb heat energy and limit rise in temperature.

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Check your answer.

Substances differ in their ability to absorb heat energy. This ability is called heat capacity.

The **specific heat (capacity) C** of a substance is measured per gram of the substance. The specific heat is the number of joules of heat energy required to raise the temperature of one gram of the substance by one celsius degree°C or one kelvin (K). [Remember that one celsius degree is the same size as one kelvin; one celsius degree is a change in temperature of one kelvin. It is not 1°C which is a particular temperature just above the MP of ice].

Specific heat, C is measured in $J\ g^{-1}\ K^{-1} = J / g\ K$
or $kJ\ kg^{-1}\ K^{-1} = kJ / kg\ K$

Substance	**C ($J\ g^{-1}\ K^{-1}$ = $kJ\ kg^{-1}\ K^{-1}$)**
expanded polystyrene	0.3
glass	0.7
solid water $H_2O(s)$	2.10
octane C_8H_{18}	2.22
glycerol $C_3H_5(OH)_3$	2.38
1,2-ethanediol $C_2H_4(OH)_2$	2.39
ethanol C_2H_5OH	2.44
methanol CH_3OH	2.53
liquid water $H_2O(l)$	4.18

Specific heats, C, of some common substances.

All of the values for chemical compounds are given to two decimal places. Why should the values for expanded polystyrene (polystyrene plastic expanded with gas to leave air spaces in the plastic) and glass only be given to one decimal place?

Check your answer.

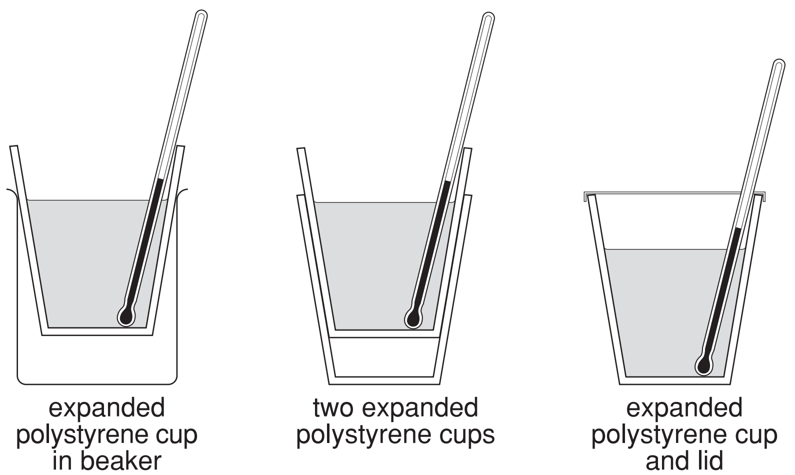
Estimating specific heats of metals

You will need:

- small compact objects made of metal
- a means of weighing the small compact metal objects or a knowledge of their weights (metal masses with their masses marked are useful)
- a saucepan in which to heat water and metal objects
- a simple calorimeter with thermometer readable to the nearest degree Celsius. Examples are shown below.
- a means of measuring the mass or volume of water used in the calorimeter.
- fork or tongs or tweezers or attached string to transfer metal at 100°C from hot water to cold water

If you are unable to measure the mass of metal objects you could use a number of coins. ‘Silver’ coins such as Australian 5 to 50 cents and ‘gold’ coins such as Australian $1 or $2 could be used. The ‘silver’ coins are 75% Cu + 25% Ni while the ‘gold’ coins are 92% Cu + 6% Al + 2% Ni. Typical masses of these Australian coins are shown in the table following.

Coin	5 cents	10 cents	20 cents	50 cents	1 dollar	2 dollar
Mass (g)	2.8	5.6	11.3	15.7	9.0	6.6



Different calorimeter designs.

Method:

1 Measure out 200 mL (= 200 g) of water into the calorimeter. Leave the thermometer in the water for at least two minutes before measuring the initial temperature T_i.

Make sure the thermometer bulb is still in the middle of the water as you read the initial temperature. Remove the thermometer after measuring the initial temperature.

2 Place the metal objects in boiling water in the saucepan. Leave for at least two minutes. The metal objects will then be at about 100°C.

3 Remove a metal object from the boiling water, quickly shake it dry then place it in the water of the calorimeter. Stir the water with the thermometer being very careful to not hit the metal object. Measure the maximum temperature reached due to transfer of heat from the metal to the water. Wait until the temperature reached is constant for two minutes or decreases over two minutes before recording it.

4 Repeat steps 1 to 3 for at least one other metal.

Results:

Metal	Object	Mass (g)	Initial water temperature T_i (°C)	Maximum temperature T_{max} (°C)	$\Delta T = T_{max} - T_i$ (°C = K)	$\frac{\Delta T}{\text{mass}}$ (K g^{-1})

Conclusions:

Compare the abilities of the different metals to store heat.

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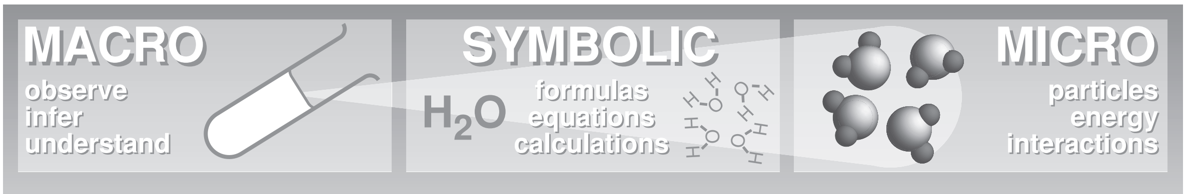
Compare your results with the published values for specific heat capacity at constant pressure and 25°C (J g^{-1} K^{-1}).

Metal and symbol	**C (J g^{-1} K^{-1})**	**λ (J s^{-1} m^{-1} K^{-1})**
aluminium Al	0.90	237
copper Cu	0.39	401
gold Au	0.13	317
iron Fe	0.45	80
magnesium Mg	1.02	156
nickel Ni	0.44	91
silver Ag	0.23	429

Is there any **correlation** between heat capacity, C and thermal conductivity, λ?

Check your answer.



Calorimetry

Calorimetry is the measurement of heat changes in a system.
The heat change could be for a physical change or a chemical change.
A **calorimeter** is equipment used to measure the heat change.

A calorimeter consists of a thermally insulated container containing a known mass of water. Heat is released or absorbed from a change and then absorbed or released by the water. This causes a change in the temperature of the water.

heat released by exothermic reaction = heat absorbed by calorimeter

heat absorbed by endothermic reaction = heat released by calorimeter

The water of the calorimeter absorbs and releases most of the heat. The other substances used in the calorimeter are insulators with low heat capacities.

Knowing the mass of water m, the specific heat C of water and the change in temperature ΔT then the amount of heat released or absorbed ΔH can be calculated using the following formula.

$$\Delta H = -m\,C\,\Delta T$$

Using C measured in J g^{-1} K^{-1}, ΔT measured in Celsius degrees (°C = K), and m measured in g then the heat absorbed or released is in J.

If the reaction is exothermic ΔT is positive so the ΔH value is negative.
If the reaction is endothermic ΔT is negative so the ΔH value is positive.

Water is used as the liquid in most calorimeters because it has such a high specific heat capacity of 4.18 J g^{-1} K^{-1}. If a question about using a calorimeter does not mention the liquid used assume the liquid is water and use 4.18 J K^{-1} g^{-1} in your calculations.

Measuring and calculating molar heat of solution

When a solute dissolves in a solvent the heat change is called the heat of solution.

Measurements of temperature change for a given mass of water in a calorimeter and a known mass of solute together with the equation $\Delta H = -m\,C\,\Delta T$ can be used to calculate ΔH. If the formula of the solute is known the heat of solution for a mole of solute, that is the molar heat of solution can be calculated.

If the temperature of the water increases on solution due to release of heat the process is exothermic. This means heat is given out to the surroundings and ΔH is reported as a negative value.

If the temperature of the calorimeter water decreases due to absorption of heat the process is endothermic. ΔH is reported as a positive value.

You will need:

- a calorimeter with thermometer readable to the nearest degree Celsius
- a means of weighing the solute to be dissolved
- a spoon for handling solute and stirring solution
- a means of measuring the volume of water; in this activity the water behaves as both a reactant and calorimeter liquid
- at least two suitable solutes such as:
  - sodium chloride $NaCl$ (table salt),
  - ammonium nitrate NH_4NO_3 (fertiliser),
  - sodium hydrogen carbonate $NaHCO_3$ (baking soda or bicarb of soda).

Method:

1 Measure 200 mL (= 200 g) of water into the calorimeter. Leave the thermometer in the water for at least two minutes before measuring initial temperature. Remove the thermometer after measuring the temperature.

2 Weigh out approximately 0.1 mol of the solute. Weigh or estimate the weight of the solute as accurately as you can. The table below gives the molar mass of the solutes suggested above.

Name	Formula	Molar mass (g)
sodium chloride	$NaCl$	58.4
ammonium nitrate	NH_4NO_3	80.0
sodium hydrogen carbonate	$NaHCO_3$	84.0

Molar mass of a number of substances.

3 Add all the solute to the water then stir vigorously until all solute is dissolved.

4 Measure the maximum or minimum temperature reached due to the release or absorption of heat. Wait until the maximum or minimum temperature reached is constant for one minute before recording it.

5 All solutions prepared in this activity can be safely disposed of as waste water down a sink.

Results:

Mass of water (g)	Solute	Solute mass (g)	Initial water temperature T_I (°C)	Maximum or minimum temperature T_{max}/T_{min} (°C)	ΔT (C degrees)

Calculations:

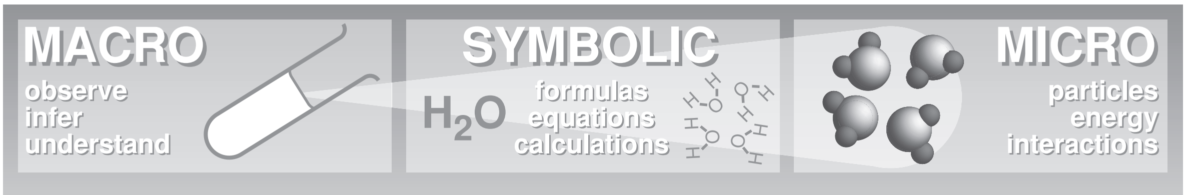
$\Delta H = -m\,C\,\Delta T$ where C = specific heat of water = 4.18 J g^{-1} K^{-1}

Calculate the molar heat of solution.

Suggestions for improvement:

Suggest three ways of improving this activity by modifications to the equipment used. Use data from the table of C values on page 7 to justify (that is, support) one of your suggestions.

Check your answers.



Dissolution and bond energies

When a solute dissolves in a solvent energy is absorbed to break bonds and energy is released when bonds form:

reactant particles + energy for breaking bonds $\rightarrow$
product particles + energy from forming bonds.

If $E_{\text{breaking bonds}} > E_{\text{forming bonds}}$ the solution process is endothermic. Energy is absorbed from the surroundings and the temperature of the calorimeter water decreases.

If $E_{\text{breaking bonds}} < E_{\text{forming bonds}}$ the solution process is exothermic. Energy is released to the surroundings and the temperature of the calorimeter water increases.

Cold packs used to reduce the pain and swelling of sports injuries usually contain a salt which undergoes endothermic solution. A single use cold pack contains a salt such as ammonium nitrate in a thick outer bag and water in a thin inner bag. When the outer bag is punched the thin inner bag breaks and water mixes with the salt. Solution of the salt forms a very cold solution.

The overall change can be regarded as made up of two main steps:

1 An endothermic step in which forces of attraction between particles are overcome eg overcoming the attraction between cations and anions in an ionic solid

$NH_4NO_3(s) + \textbf{heat} \rightarrow NH_4^+ + NO_3^-$

2 An exothermic step in which ion-dipole bonds are formed between the solute ions and solvent molecules

$NH_4^+ + NO_3^- + (x + y)\ H_2O \rightarrow NH_4^+(OH_2)_x + NO_3^-(H_2O)_y + \text{heat}$

In the solution of ammonium nitrate the **heat** absorbed in step 1 is greater than the heat released in step 2 and so the overall change is endothermic.

Explain why the water molecules attracted to ammonium ions are written as $NH_4^+(OH_2)_x$ while the water molecules attracted to nitrate ions are shown as $NO_3^-(H_2O)_y$.

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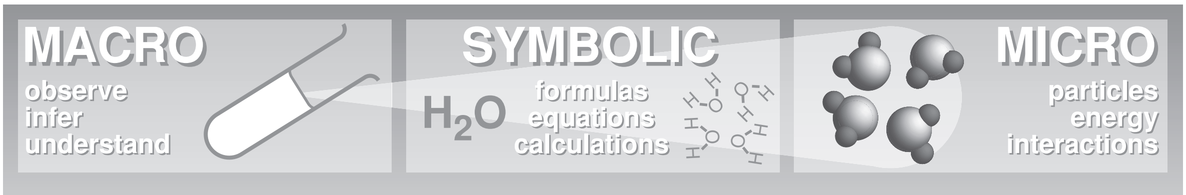
Check your answers.



Exothermic solutions

When acids or alkalis are dissolved in water considerable heat is released. Great caution is required when preparing acidic or alkaline solutions. If too much heat is released the solvent can boil and spray acidic or alkaline solutions from the container. The harm that acidic solutions can do to human skin and eyes is well known.
Alkaline solutions are even more harmful to the eyes. Alkaline solution on the skin reacts with skin fats producing soap and the slippery feel of alkaline solution.

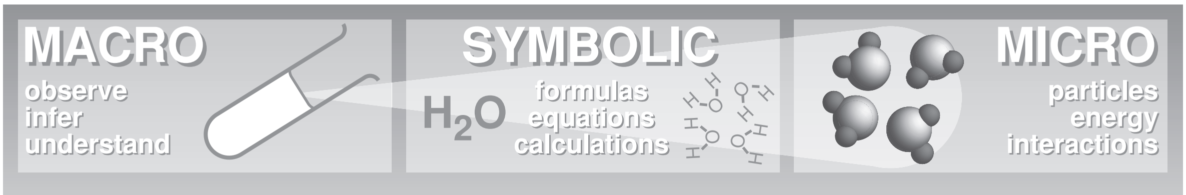
Eye protection must be worn and protective clothing such as a laboratory coat should be worn when preparing an acidic or alkaline solution.
This activity must only be undertaken near a plentiful supply of cold water so that body and clothes can be immediately flooded with water for minutes in the event of an accident.

The acid or alkali must be added in small amounts to a much larger volume of water with plentiful stirring. Make sure the portion that is first added has dissolved before further acid or alkali is added.

The solution prepared must be stored in a suitable container. This should be labelled with:

- the written name of the chemical (chemical formula is an optional extra – most people can read a name, few understand formulas)
- concentration (approximate or accurate, include units)
- date the solution was prepared
- name of person who prepared the solution.

Some salts which have exothermic heats of solution are anhydrous copper sulfate $CuSO_4$, calcium chloride $CaCl_2$, sodium acetate $NaCH_3COO$ and most lithium salts.



The human body, heat and water

The average human body contains about 40 kg of water and the chemical reactions in its cells release about 10 000 kJ of heat each day.

Using $\Delta H = -m\,C\,\Delta T$

10 000 000 J released in a day = 40 000 g x 4.18 J K^{-1} g^{-1} x ΔT

ΔT = 10 000 000 / 40 000 x 4.18 = 59.8 K = 59.8 Celsius degrees increase in temperature.

In other words, the heat released inside a human body in a day could raise the body temperature from 37.8°C by 59.8 Celsius degrees to 97.8°C – close to the temperature of boiling water!

Clearly this does not happen as heat escapes from the human body.

One of the main ways in which the human body can lose heat is by evaporation of sweat which is mostly water.

1 Calculate how much water would need to be vaporised to remove 10 000 000 J = 10 000 kJ from a human body. Use the heat of vaporisation of water = 2.25 kJ/g in your calculation.

2 The average person sweats 0.5 kg a day. (A physically active person might sweat 4 kg a day while someone active in a hot desert region could sweat more than 15 kg per day) How much heat energy would be removed by the vaporisation of 0.5 kg of water?

3 List ways other than sweating that the human body can use to lose heat energy.

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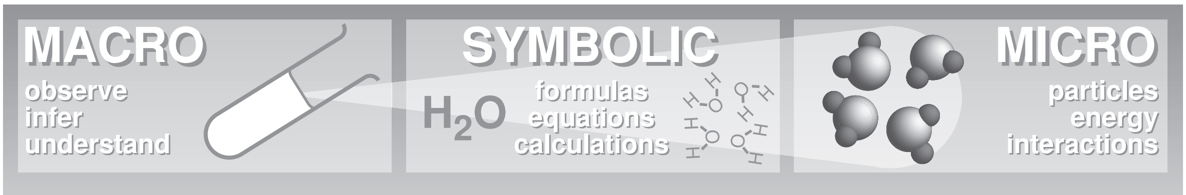
Check your answers.



Life and water’s heat capacity

Chemically all living things can be pictured as aqueous solutions surrounded by membranes. The chemical reactions in the aqueous solutions and on the surface of the membranes are catalysed by enzyme proteins which function in a narrow temperature range.

It is critical that the narrow temperature range be maintained otherwise the structure of enzymes can change so much that they cannot function. A temperature rise in a human can be fatal as essential enzymes stop functioning. Normal body temperature is 37°C while at 40.5°C the person can become delirious. Permanent body damage or death can occur at 44°C.

The following properties of water all help water to minimise temperature change and act as a thermal regulator. Water has:

- high heat capacity
- low viscosity and relatively high heat conductivity which support movement of heat away from a source
- large heat of vaporisation.

In an aquatic environment the temperature is held in a narrow range. The high heat capacity of water minimises fluctuations in temperature. Most aquatic animals’ body temperature depends on the temperature of their surroundings. Large fish such as tuna which chase other animals have ways of reducing heat loss. This ensures that chemical reactions releasing energy can proceed at an adequate rate to supply the bursts of energy needed to catch prey.

A higher proportion of terrestrial animals maintain a constant body temperature. Sweating to cool the body and behaviours like seeking shade, hunting at night rather than day, help minimise fluctuations in temperature. Dogs pant by breathing with their moist mouth open and moist tongue hanging out. Evaporation of this moisture cools the dog’s body.

Ocean currents probably carry more heat energy from the equatorial regions of the earth to the poles than the more rapid movement of air in the atmosphere. The Gulf Stream in the Atlantic Ocean provides the British Isles with sufficient heat for a moderate climate. The heat energy transferred from the Gulf of Mexico to the Arctic Ocean per day by the Gulf Stream is about equal to all the heat energy released in a year by the burning of coal throughout the world.

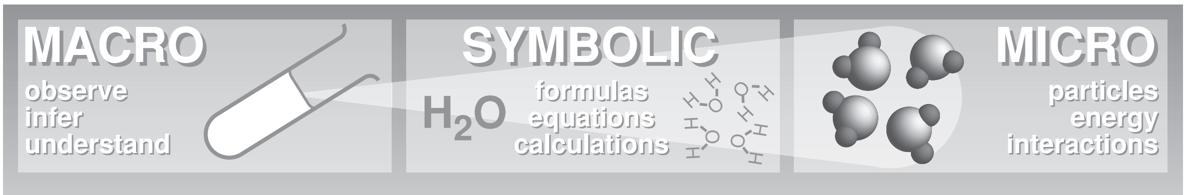
The global climate changes monitored by Southern Oscillation Index measurements (El Nino and La Nina effects) are driven by the movement of warm water in the Pacific Ocean.

The specific heat capacity of ice is half that of liquid water but is still significant in the Earth's climate. For information on how liquid water and solid water affect climate look at the web sites at:
http://www.lmpc.edu.au/science

If the body of a mammal starts to freeze between 2 and 5% ice formation kills the mammal. Growing ice crystals slash and stab cell membranes turning tissue mushy. Antarctic fish produced a sugar–coated protein (glycoprotein), frogs produce glucose and many insects produce glycerol as anti-freeze. These substances make it difficult for small ice crystals to grow.

Have you completed your open-ended investigation on the effect of salt on the boiling point of water? You need to write a report on this for Part 6.

One type of genetically modified strawberry contains a gene from the Arctic flounder fish. This gene produces an anti-freeze that protects the strawberry against frosts. Can this type of strawberry be classified as food for vegetarians?



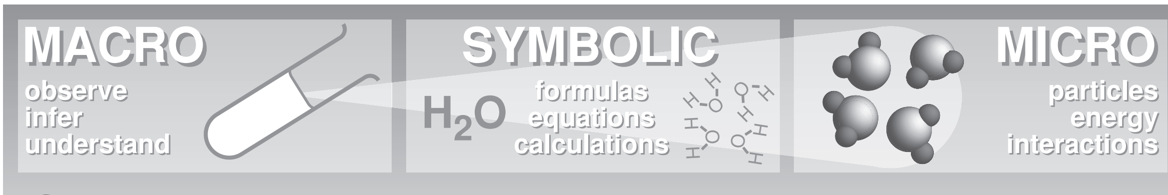
Thermal pollution

Thermal pollution occurs when heat energy is added to the environment causing unwanted or harmful effects. Human activity producing increased temperature in waterways is the main source of thermal pollution.

Because of water's large heat capacity it is often used as a cooling fluid for industry or electrical energy generation in power stations. When hot water is released into the environment such as in a river or shallow lake the high heat capacity of water results in release of a large amount of heat energy. This large amount of energy may take a long time to disperse.

Heat released into a lake or river immediately lowers the concentration of dissolved gases such as oxygen. Large fish may not have enough oxygen available and suffocate. Other living things may not survive the sudden change in their body temperature. The higher temperature of the water may be outside the normal temperature range for many organisms.
The death of one type of organism may threaten the survival of other organisms higher in a food chain.

Do you understand the difference between heat energy and temperature? Suppose the same quantity of heat energy was released into a lake on two separate occasions. Would it matter if the water released was at 40°C or 80°C if the same amount of energy was released?
Discuss this with another person.



Suggested answers

Measuring the heat of vaporisation of water

1 Your answer will probably be greater than the published 40.6 kJ mol^{-1}.

2 Some of the heat supplied to your equipment did not go into changing liquid water into gaseous water. Instead that heat was used to heat the electric kettle or escaped to the surroundings. The published value would have been calculated after allowing for heat required to heat the calorimeter and heat losses to the surroundings.

3 Up to 100°C the energy is increasing the kinetic energy of the water molecules. The temperature stays at 100°C while heat energy is used to overcome attractive forces between molecules close together in the liquid state. This energy is stored in the gaseous molecules. When steam comes in contact with you it condenses to liquid releasing this stored heat and scalding your skin.

Specific heat (capacity) C

The water in one balloon absorbs heat energy from the match flame. This balloon will either not burst or take longer heating before it bursts.

The specific heats C for polystyrene and glass are only given to one decimal place because these materials are mixtures.
Expanded polystyrene is a mixture of polystyrene and chemical residues from any process used to produce gases to expand the polystyrene.
The density of expanded polystyrene can be adjusted by using different amounts of polystyrene and gas producing chemicals. Glasses consist of varying proportions of silica, sodium silicate, calcium silicate and other chemicals depending on the type of glass. The specific heat value will vary according to the different proportions of components in each mixture. The other substances in the table are pure compounds.
Their specific heats values are fixed.

Estimating specific heats of metals

There is little if any correlation between heat capacity and thermal conductivity eg gold with the lowest heat capacity listed has one of the highest thermal conductivities. Heat capacity is the ability to store heat whereas thermal conductivity is the ability to allow passage of heat.

Measuring and calculating molar heat of solution

Some published values for molar heats of solution are:
NaCl 222 J/mol; KCl 984 J/mol; NH_4NO_3 1 469 J/mol
Values are positive because they are endothermic heats of solution.

1 Use a thermometer that can be read to better than the nearest degree.

2 Use a well insulated calorimeter eg two expanded polystyrene cups with a lid for the inner cup. The lid could have a small hole for the thermometer. (but remember the thermometer bulb must still be in the middle of the well stirred solution when the temperature is read).

3 Carry out the experiment once and determine the temperature change eg four degrees decrease. Repeat the experiment but start with water that is half the temperature change (two degrees for this example) above room temperature. The temperature change for the equipment should then be from two degrees above room temperature to two degrees below room temperature. The heat lost from the equipment to the surroundings when it is above room temperature should then balance the heat gained by the equipment when it is below room temperature.

Dissolution and bond energies

$NH_4^+(OH_2)_x$ indicates the slightly negative oxygen ends of *x* water molecules are attracted to a positive ammonium ion while in $NO_3^-(H_2O)_y$ the slightly positive hydrogen ends of *y* water molecules are attracted to a negative nitrate ion.

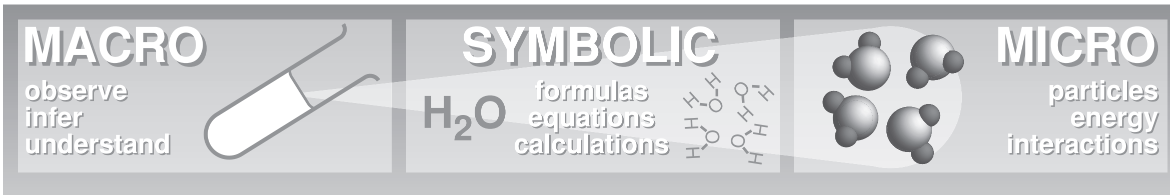
The human body, heat and water

1 ΔH = 10 000 kJ   2.25 kJ vaporises 1 g of water

Therefore 10 000 kJ vaporises 10 000/2.25 g of water = 4.44 kg

2 0.5 kg = 500g requires 500 x 2.25 kJ = 1125 kJ to vaporise

3 Loss of heat in air, urine and faeces passing out of the body; radiation, convection and conduction from the skin.



Exercises - Part 5

Exercises 5.1 to 5.2 Name: ________________________________

Exercise 5.1 Heat packs

Chemical reactions involving water are used in heat packs for warming cold parts of the human body and for producing hot meals.

Heat packs purchased from camping and mountaineering stores consist of an inner paper bag in an outer plastic bag. The inner paper bag contains iron powder, activated charcoal, table salt and sawdust, all dampened with water. The inner paper bag has small holes that allow air to enter. When the inner bag is shaken in air, oxygen enters the small holes and reacts with the iron to form iron(III) oxide and release heat. If the inner paper bag is placed back in the outer plastic bag and this sealed a limited amount of heat is released. The heat pack can be used a number of times by doing this until all the iron is changed to iron(III) oxide.

a) Draw a labelled diagram showing the parts of a heat pack.

b) Write a word equation and a balanced equation in symbols for the reaction that releases heat. Show the states of matter for the two reactants and the product.

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c) $\Delta H = -47$ kJ/mol of product. How much energy is released per mol of iron that reacts? Note that 1 mol of Fe gives 0.5 mol of Fe_2O_3.

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d) Heat packs used by military personnel to produce hot meals consist of a bag into which a meal container, a heater and some water are placed. The heater unit contains magnesium and a small amount of iron and salt. The small amount of iron and salt help remove the surface oxide layer on the magnesium so the metal is exposed and can react readily with the water. The exothermic reaction is:

magnesium + water → magnesium hydroxide + hydrogen

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Under the word equation write the balanced equation in symbols including states of matter.

e) Explain why the bag used cannot be tightly sealed and must be kept away from flames and sources of electrical sparks.

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f) $\Delta H = -355$ kJ/mol of magnesium. How much heat would be released by reaction of 18 g of water with excess magnesium? Note the mole ratio between magnesium and water in the balanced equation before commencing your calculation.

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Exercise 5.2: Comparing solvent data

Compound	acetone	methanol	ethanol	1-propanol	octane	water
Molecular formula	C_3H_6O	CH_3OH	C_2H_5OH	C_3H_7OH	C_8H_{18}	H_2O
Molar mass (g)	58	32	46	60	114	18
Dipole moment (D)	2.9	1.70	1.69	1.55	–	1.86
MP (°C)	–95	–98	–114	–126	–57	0
BP (°C)	56	65	78	97	126	100
C ($Jg^{-1}K^{-1}$)	2.17	2.53	2.44	2.39	2.22	4.18
Density (g cm^{-3})	0.79	0.79	0.79	0.80	0.70	1.00
Flashpoint (°C)	–19	11	12	15	13	–
Ignition temperature (°C)	538	385	363	370	206	–
TLV (ppm)	750	200	1000	200	300	–

Flashpoint is the lowest temperature at which a small flame causes vapour above a flammable liquid to ignite.

Ignition temperature is the minimum temperature at which the substance can be ignited.

TLV = Threshold Limit Value = maximum concentration to be tolerated in working environments.

Use the information in this table to answer the questions that follow.

a) Explain why water does not have a flashpoint or ignition temperature.

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b) Which property shows that acetone is more polar than water?

____

c) Which solvent has the highest MP, highest density but the lowest molar mass?

____

____

d) List a property of water that could be used to explain:

i) its ability to dissolve ionic solids

____

ii) its ability to narrow the range of temperature on Earth

____

iii) its safe use as a solvent

____

e) Comment on the specific heat capacity C of water compared with other solvents.

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f) Discuss the importance of the specific heat capacity C of water to aquatic life.

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Chemistry

Preliminary Course
Stage 6

NEW SOUTH WALES DEPARTMENT OF EDUCATION AND TRAINING

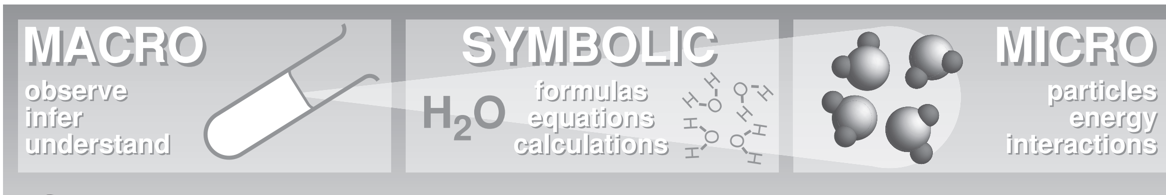
TAFE NSW

OPEN TRAINING AND EDUCATION NETWORK - DISTANCE EDUCATION

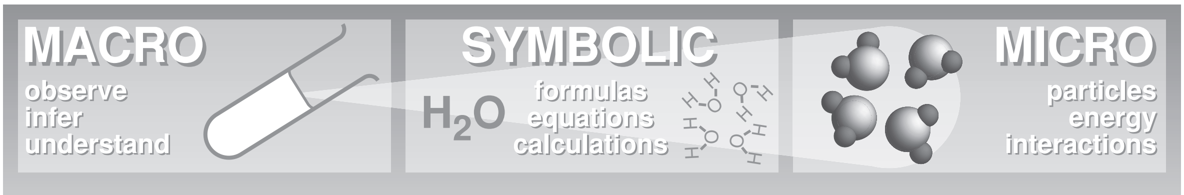
Water

Part 6: Review and report

Incorporating October 2002
AMENDMENTS


MACRO
observe
infer
understand
SYMBOLIC
H_2O
formulas
equations
calculations
MICRO
particles
energy
interactions

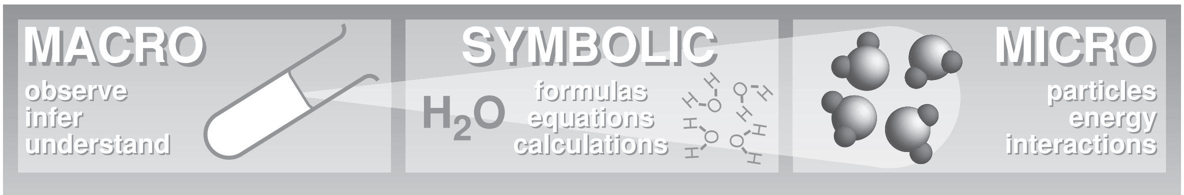
Contents



Introduction

In this part you will review the first five parts of this module, report on your open–ended investigation and assess your chemistry skills using information in your report.

Attach the report of your open–ended investigation (on the effect of salt on the boiling point of water) as Exercise 6.3.



Key ideas

Assess your understanding of this module by answering these true/false statements one part at a time. Write T or F at the end of each statement. Approximately half the statements are true. If you decide the statement is false, rewrite the statement so that it is true.

Water on Earth

- Most of the hydrosphere's water is in the liquid state.
- Water has its maximum density in the solid state.
- Respiration removes carbon dioxide from the atmosphere.
- Fungi and bacteria help recycling of chemicals by decaying the bodies of dead plants and animals.
- Water is essential to the operation of the nitrogen cycle and the carbon cycle.
- Plate tectonic forces are involved in returning some carbon, nitrogen and oxygen atoms to the atmosphere.
- An iceblock would float higher in fresh water than salt water.
- Of the four spheres – biosphere, lithosphere, hydrosphere and atmosphere – the atmosphere contains the least amount of water.
- When liquid water solidifies the water molecules move further apart.
- Of the four spheres – biosphere, lithosphere, hydrosphere and atmosphere – the biosphere contains the greatest amount of water.

Structure and properties

- The intermolecular forces keeping water molecules together are dispersion forces, dipole-dipole forces and hydrogen bonding.
- The strongest intermolecular force between water molecules is hydrogen bonding.
- In the solid state each water molecule is surrounded hexagonally by six other water molecules.
- In the water molecule there are two electron pairs in the valence shell of the oxygen atom.
- Molecules containing polar bonds are always polar.
- A stream of non-polar liquid molecules is attracted towards an electrically charged object.
- Dispersion forces occur between all molecules.
- Hydrogen bonding occurs between molecules which have a hydrogen covalently bonded to a N, O or F atom.
- Surfactant added to water increases the surface tension of the water.
- The viscosity of a liquid usually increases as the temperature rises.

The important solvent

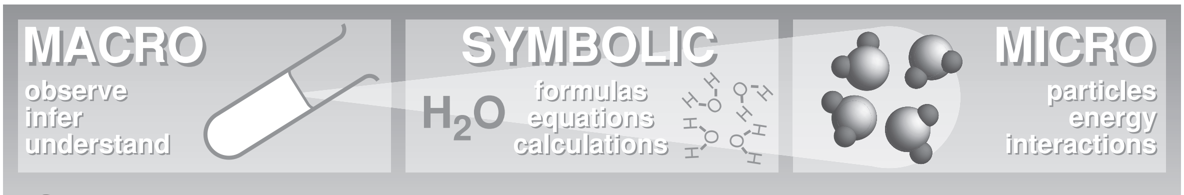
- Water is a solvent for many ionic solids and small molecules that can hydrogen bond.
- Water molecules between a cation and an anion increase the attractive forces between the ions.
- The shape of large biological molecules such as proteins is affected by the presence of water molecules.
- Non-polar substances usually dissolve in non-polar solvents.
- Solubility of gases in water increases with temperature rise.
- Gases which chemically react with water forming ions are the most soluble.
- Diffusion is the main way in which perfume molecules are carried into your nose.
- Diffusion is the main way in which small molecules pass in and out of cells.
- ‘Like dissolves like’ is an explanation, not a generalisation.
- Covalent network substances and large molecules are usually water insoluble.

Salts in water

- Equilibrium occurs when solid lithium chloride is in contact with a saturated solution of sodium chloride.
- If solid sodium chloride is placed in water in a beaker this is a system at equilibrium.
- A system at equilibrium has change occurring at the macro level but not the micro level.
- Most carbonates are soluble.
- All group 1 salts are soluble.
- 0.5 ppm = 500 ppb.
- Quantitative analysis using volumes of known concentration is called volumetric analysis
- A 2 mol L^{-1} Na_2CO_3 solution has $[Na^+]$ = 1 mol L^{-1}.
- A 2 mol L^{-1} Na_2CO_3 solution has $[CO_3^{2-}]$ = 1 mol L^{-1}
- 2 mol Na_2CO_3 dissolved in 1 litre of water is a 2 mol L^{-1} Na_2CO_3 solution.

Water and heat

- A specific heat of 4.18 J g^{-1} K^{-1} for water is equal to 4.18 kJ kg^{-1} K^{-1}
- 4.18 J g^{-1} K^{-1} = 4.18 J K^{-1} g^{-1}
- Calorimeters to which water is added are usually made from materials which are good conductors of heat.
- When an endothermic reaction is carried out in a calorimeter heat energy is released to the water of the calorimeter.
- When an endothermic reaction is carried out in a calorimeter the temperature rises in the calorimeter.
- Bond breaking absorbs energy.
- Bond formation releases energy.
- When acid or alkali dissolves in water the energy involved in bond breaking is greater than the energy involved in bond formation.
- When thermal pollution of a waterway occurs the dissolved oxygen levels decrease.
- The specific heat of water is generally lower than that of other solvents.



Suggested answers

Water on Earth

- Most of the hydrosphere's water is in the liquid state.   T
- Water has its maximum density in the solid state.   F

  Water has maximum density at 4°C in the liquid state.
- Respiration removes carbon dioxide from the atmosphere.   F

  Respiration releases carbon dioxide to the atmosphere.
- Fungi and bacteria help recycling of chemicals by decaying the bodies of dead plants and animals.   T
- Water is essential to the operation of the nitrogen cycle and the carbon cycle.   T
- Plate tectonic forces are involved in returning some carbon, nitrogen and oxygen atoms to the atmosphere.   T
- An iceblock would float higher in fresh water than salt water.   F

  Salt water is denser than fresh water so an ice block should float higher in salt water.
- Of the four spheres – biosphere, lithosphere, hydrosphere and atmosphere – the atmosphere contains the least amount of water.   F

  The biosphere living matter contains the least amount of water if the water in which living things live is excluded.
- When liquid water solidifies the water molecules move further apart. T
- Of the four spheres – biosphere, lithosphere, hydrosphere and atmosphere – the biosphere contains the greatest amount of water. F

  98.3% of the Earth's water is in the hydrosphere, 1.7% groundwater in the lithosphere, 0.0010% in the atmosphere and 0.0001% in the biosphere's living matter.

Structure and properties

- The intermolecular forces keeping water molecules together are dispersion forces, dipole-dipole forces and hydrogen bonding. T
- The strongest intermolecular force between water molecules is hydrogen bonding. T
- In the solid state each water molecule is surrounded hexagonally by six other water molecules. F

  In ice each water molecule is tetrahedrally surrounded by four water molecules. The structure made up of tetrahedrally joined water molecules has hexagonal shaped spaces.
- In the water molecule there are two electron pairs in the valence shell of the oxygen atom. F

  In the water molecule there are two bonding electron pairs connecting hydrogen atoms and two nonbonding electron pairs in the valence shell of the oxygen atom.
- Molecules containing polar bonds are always polar. F

  Molecules containing polar bonds can be non polar if the polar bonds cancel out eg CO_2, BF_3, CCl_4
- A stream of non-polar liquid molecules is attracted towards an electrically charged object. F

  Only polar molecules, those with dipoles, are attracted towards electrically charged objects.
- Dispersion forces occur between all molecules. T
- Hydrogen bonding occurs between molecules which have a hydrogen covalently bonded to a N, O or F atom. T
- Surfactant added to water increases the surface tension of the water. F *Surfactant molecules between the water molecules at a surface decrease hydrogen bonding between the water molecules at the surface. This decreases the surface tension of water.*
- The viscosity of a liquid usually increases as the temperature rises. F

  Viscosity decreases as temperature rises and the liquid flows more easily.

The important solvent

- Water is a solvent for many ionic solids and small molecules that can hydrogen bond. T
- Water molecules between a cation and an anion increase the attractive forces between the ions. F

  Water molecules decrease attractive force between ions to about 1/80 of their original strength.
- The shape of large biological molecules such as proteins is affected by the presence of water molecules. T
- Non-polar substances usually dissolve in non-polar solvents. T
- Solubility of gases in water increases with temperature rise. F

  Solubility of gases in water decreases with temperature rise.
- Gases which chemically react with water forming ions are the most soluble. T
- Diffusion is the main way in which perfume molecules are carried into your nose. F

  Convection currents and breathing in of air are much faster than diffusion.
- Diffusion is the main way in which small molecules pass in and out of cells. T
- 'Like dissolves like' is an explanation, not a generalisation. F

  'Like dissolves like' is a generalisation, not an explanation.
- Covalent network substances and large molecules are usually water insoluble. T

Salts in water

- Equilibrium occurs when solid lithium chloride is in contact with a saturated solution of sodium chloride. F

  Equilibrium occurs when a solid is in contact with a saturated solution of itself. LiCl is actually more soluble than NaCl so the solid LiCl would dissolve on contacting the saturated solution of NaCl.
- If solid sodium chloride is placed in water in a beaker this is a system at equilibrium. F

  This is not equilibrium as the solid will dissolve. The rate of dissolution will be much greater than the rate of any precipitation.
- A system at equilibrium has change occurring at the macro level but not the micro level. F

A system at equilibrium has change occurring unseen at the micro level but not the observable macro level.

- Most carbonates are soluble. F

  Most carbonates are insoluble.

- All group 1 salts are soluble. T
- 0.5 ppm = 500 ppb. T
- Quantitative analysis using volumes of known concentration is called volumetric analysis T
- A 2 mol L^{-1} Na_2CO_3 solution has $[Na^+]$ = 1 mol L^{-1}. F

  Because each unit of Na_2CO_3 contains two sodium ions the concentration of Na^+ will be twice the concentration of sodium carbonate, that is 2x2 = 4 mol L^{-1}.

- A 2 mol L^{-1} Na_2CO_3 solution has $[CO_3^{2-}]$ = 1 mol L^{-1} F

  Because each unit of Na_2CO_3 contains one carbonate ion the concentration of CO_3^{2-} will be the same as the concentration of sodium carbonate = 2 mol L^{-1}.

- 2 mol Na_2CO_3 dissolved in 1 litre of water is a 2 mol L^{-1} Na_2CO_3 solution. F

  2 mol Na_2CO_3 dissolved in 1 litre of solution is a 2 mol L^{-1} Na_2CO_3 solution. To prepare the solution 2 mol of solute is added to water and the total volume of solution made up to 1 litre. The volume of water used will be close to but probably not exactly 1 litre.

Water and heat

- A specific heat of 4.18 J g^{-1} K^{-1} for water is equal to 4.18 kJ kg^{-1} K^{-1}

  T

- 4.18 J g^{-1} K^{-1} = 4.18 J K^{-1} g^{-1} T
- Calorimeters to which water is added are usually made from materials which are good conductors of heat. F

  Calorimeters are usually made from good heat insulators such as glass or expanded polystyrene.

- When an endothermic reaction is carried out in a calorimeter heat energy is released to the water of the calorimeter. F

  Endothermic reactions absorb energy from their surroundings so heat energy is removed from the water.

- When an endothermic reaction is carried out in a calorimeter the temperature rises in the calorimeter. F

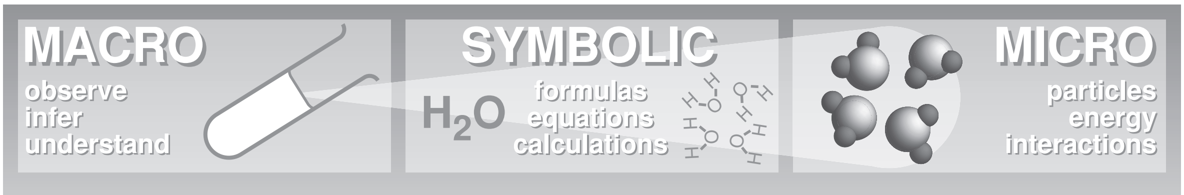
An endothermic reaction absorbs heat from the calorimeter and the temperature decreases.

- Bond breaking absorbs energy. T
- Bond formation releases energy. T
- When acid or alkali dissolves in water the energy involved in bond breaking is greater than the energy involved in bond formation. F

  Dissolution of acid or alkali is exothermic so the energy released by bond formation must be greater than that absorbed in bond breaking.

- When thermal pollution of a waterway occurs the dissolved oxygen levels decrease. T
- The specific heat of water is generally lower than that of other solvents. F

  Water has a higher specific heat than practically all other solvents. (liquid ammonia and aqueous ammonia solution are the only common solvents with a higher specific heat)



Appendix

Risk assessment

1 Record (written, simple diagrams could be useful) the procedures you will be carrying out.

- chemicals to be used
- chemicals to be made; give details of how you intend to make them
- quantities of chemicals to be used or made
- equipment
- techniques
- non-chemical hazards

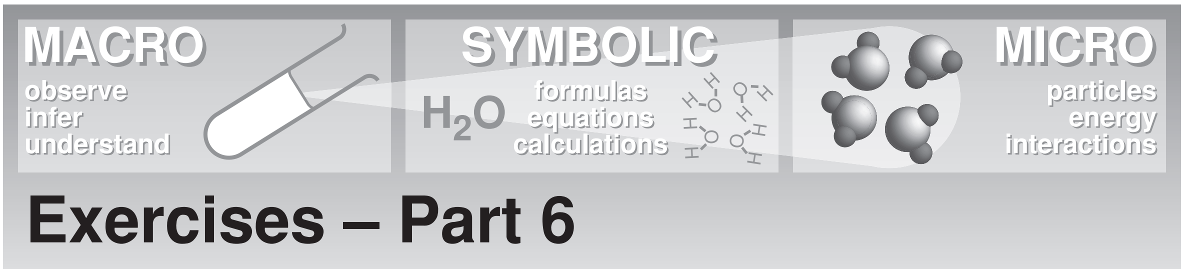
2 Identify any hazardous chemicals you are planning to use. Look at labels and information from suppliers such as MSDSs (Material Safety Data Sheets).

Identify products and possible by-products that could be hazardous.

3 Record potential hazards and ways that you might be exposed to these hazards.

4 Decide how you will minimise exposure to hazards.

5 Plan safe and environmentally considerate disposal of chemicals.



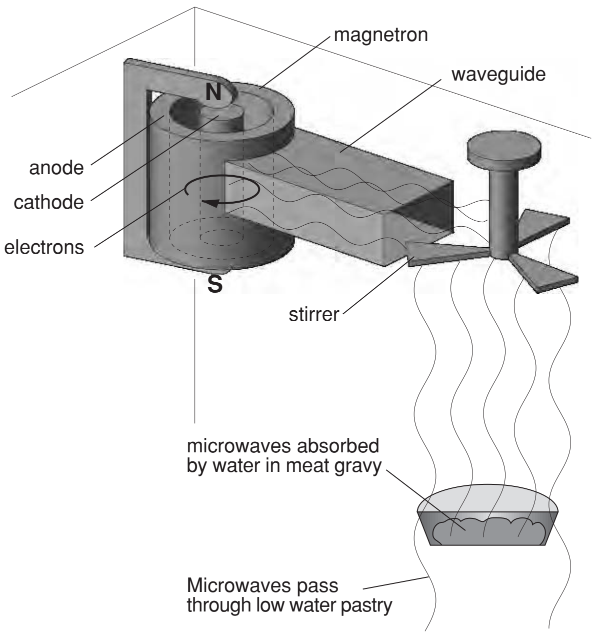
Exercises – Part 6

Exercises 6.1 to 6.3 Name: ________________________________

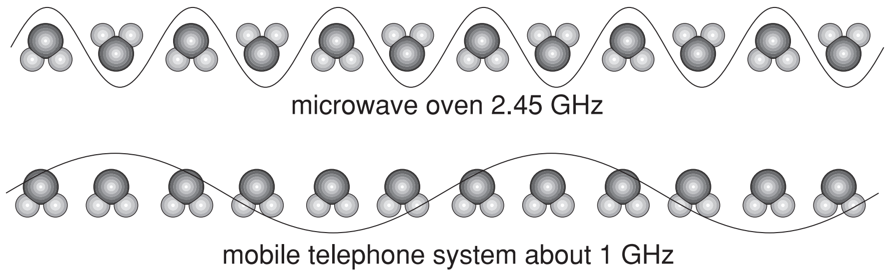
Exercise 6.1: Interpreting diagrams

The two diagrams below show

- a meat pie being heated in a microwave oven



- a simplification of what happens when 2.45 GHz frequency microwaves pass through water. (the water molecules do not move as much as shown if hydrogen bonded to other water molecules or carbohydrates or proteins).



Use the information provided in the diagrams to describe in words the importance of water in microwave heating of food.

Exercise 6.2: Chemistry skills check list

Now that you have completed three of the four Preliminary course modules it is time for you to check that you have had opportunities to develop skills in:

1 planning investigations
2 conducting investigations
3 communicating information and understanding
4 developing scientific thinking and problem solving techniques
5 working individually and in teams.

Place a tick √ to the right hand side of any skill you have done in any of the Preliminary modules. Place a dash – to the right hand side if you cannot tick the skill.

1 Planning investigations. Have you:

a) carried out an open-ended investigation?

b) collected

i) qualitative data such as colour, change in appearance, and so on?

ii) quantitative data, that is, made measurements?

c) identified the order of magnitude and uncertainty in a measurement? eg a measurement of 984 g to the nearest g has an order of magnitude of 10^3 g and uncertainty of ±0.5g

d) used and recorded the correct units for measurements?

e) organised gathered information so that it can be efficiently analysed? eg into a table or on a graph

f) understood the difference between a dependent variable and an independent variable?

The dependent variable is what you are interested in measuring. The independent variables are quantities that could affect the value of the dependent variable measurement.
For example, consider measuring the effect of temperature on colour of a gas mixture. Colour is the dependent variable and temperature, concentration, pressure, for example, would be independent variables.

g) identified variables eg. temperature, amounts of chemical, that need to be kept constant, developed strategies eg carry out experiment at a constant temperature, to ensure variables are kept constant,

demonstrated the use of a control eg compared result between experiments that differ by one factor only?

h) aimed for validity (leading to effective results and worthwhile conclusions) and reliable (trustworthy) data?

i) designed and reported on appropriate procedures?

j) predicted issues eg safety, and acted accordingly eg used safety procedures?

k) identified and set up appropriate equipment for investigations?

l) carried out a risk assessment of intended experimental procedures (if not, see the Appendix)

identified and addressed potential hazards eg flammable liquid use near flame or electrical sparks?

m) identified technology that could be used eg computer or thermometer, and determined its suitability?

n) recognise the difference between destructive and non-destructive testing of a material? eg. actual destruction of the material in a lab compared with computer simulation of destruction of the material

o) analysed potentially different results of these two procedures? eg consider if the results of an actual test to destruction of a material will be different from computer simulation of that destruction

2 Conducting investigations aims to produce valid and reliable data.

Have you:

a) modified procedures and analysed the effect of these modifications?

b) minimised hazards and waste?

c) disposed of waste materials carefully and safely?

d) identified and used safe work practices?

e) used appropriate data collection techniques eg enter measurements in a table, and use appropriate sensors eg. a thermometer?

f) repeated trials of experiments as appropriate?

g) accessed information from a range of resources including the Internet?

h) practised efficient data collection techniques eg used search engines on the Internet?

i) extracted information from graphs, tables, written and spoken material?

j) summarised and collated (gathered together) information from a range of resources?

k) identified practising male and female Australian scientists, areas in which they are working and information about their research eg from TV programs, newspapers, magazines or internet sites such as www.abc.net.au/science

l) assessed (judged the value of) the accuracy of measurements and calculations and their relative importance?

m) identified and applied mathematical formulas

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

$$n = \frac{m}{M}$$

$$c = \frac{n}{V}$$

$$cV = \text{constant}$$

n) illustrated trends and patterns by selecting and using appropriate methods, including computer assisted analysis eg. using a computer spreadsheet program?

o) evaluated the relevance of information?

p) assessed the reliability of information by considering information from various sources?

q) compared accuracy of scientific information from mass media by comparison with similar information from scientific journals
eg Scientific American www.sciam.com
New Scientist www.newscientist.com
Nature www.nature.com

3 Communicating information and understanding.

Have you:

a) used different text types. For example:

- set out procedure steps in your report for the open-ended investigation?
- offered explanations in answering questions?
- argued a point of view (exposition)?
- described an experiment using aim, method, results, conclusion?

b) selected and used appropriate media. For example:

– writing
– drawing
– photograph
– video?

c) acknowledged sources of information?

d) used symbols, formulas and appropriate units?

e) used a variety of pictorial representations to show relationships and present information clearly and succinctly eg

– made molecular models
– drawn chemical formulas
– written chemical equations in symbols
– used flow charting
– used a key
– used a table
– used a Venn diagram?

f) selected and drawn appropriate graphs?

g) identified situations where use of curve of best fit is appropriate on a graph?

4 Developing scientific thinking and problem-solving techniques to draw valid conclusions from gathered data and information

Have you analysed information to:

a) identify trends, patterns and relationships as well as contradictions? eg MP and BP of CH_4 NH_3 H_2O HF Ne

b) justify inferences (what you first think happens) and conclusions (what you are more certain happened)?

c) identify and explain how data (observed information) supports or refutes an hypothesis (satisfactory explanation for a variety of observations), a prediction (what you think will happen) or a proposed solution to a problem?

d) predict outcomes and generate plausible explanations related to observations?

e) make and justify generalisations? eg 'like dissolves like'

f) use models to explain and/or make predictions? eg particle theory

g) use cause and effect relationships to explain? eg thermal pollution and death of fish

h) identify examples where scientific ideas interconnect?
eg. biogeochemical cycles

Have you solved problems by:

a) identifying and explaining the nature of the problem?

b) choosing a strategy to solve a problem?

c) developing a range of possible solutions?

d) evaluating the appropriateness of different strategies?

Have you used available evidence to:

a) design and produce creative solutions to problems?

b) propose coherent and logical ideas using scientific principles?

c) apply critical thinking in considering predictions, hypotheses and results?

d) formulate cause and effect relationships?

5 Have you worked on some activities in chemistry:

– by yourself

– with another person or persons?

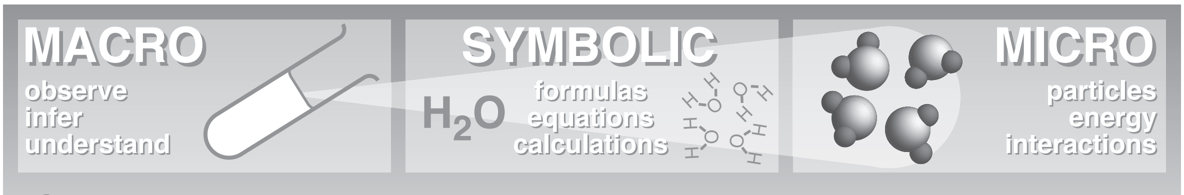
If you ticked most of these skills – congratulations!!!

You are well on the way towards completing the Preliminary course chemistry skills (8.1 pp19-20 *Chemistry Stage 6 Syllabus* © Board of Studies NSW, 2002).

When you have completed the last Preliminary module *Energy* check through this list again concentrating on the skills you have just put a – next to.

Exercise 6.3: Report on the open-ended investigation

‘identify and describe the effect of salt on the boiling point of water’



Student evaluation of module

Name: ______________________ Location: ______________________

We need your input! Can you please complete this short evaluation to provide us with information about this module. This information will help us to improve the design of these materials for future publications.

1 Did you find the information in the module clear and easy to understand?

____

2 What did you most like learning about? Why?

____

____

3 Which sort of learning activity did you enjoy the most? Why?

____

____

4 Did you complete the module within 30 hours? (Please indicate the approximate length of time spent on the module.)

____

____

5 Do you have access to the appropriate resources? eg a computer, the internet, scientific equipment, chemicals, people that can provide information and help with understanding science

____

____

Please return this information to your teacher, who will pass it along to the materials developers at OTEN – DE.

Learning Materials Production
Open Training and Education Network – Distance Education
NSW Department of Education and Training