Chemical substances

Introduction

This topic explores the key concepts of chemical substances as they relate to:

- materials and society
- matter
- atomic structure and the periodic table
- bonding
- ionic bonding
- metallic bonding
- covalent bonding
- plastics and fibres.

Key concepts of chemical substances

The activities in this topic are designed to explore the following key concepts:

- A chemical is a substance that is either naturally occurring or manufactured by humans.
- Objects that take up space and have mass are called ‘matter’.
- Matter can exist in one of four states: solid, liquid, gas or plasma.
- Substances can have a range of physical properties, including colour, density, electrical conductivity, hardness and flexibility.
- Substances have chemical properties that enable them to react chemically with other substances to form new substances.
- Matter consists of tiny particles called ‘atoms’.
- Atoms can bond with other atoms to form molecules or lattice structures.
- The temperature of a substance is related to the average kinetic energy of the particles that make up the substance.
- In general, atoms achieve a stable state by losing, gaining or sharing electrons to form chemical bonds with other atoms.
- All chemical bonds are electrostatic in nature.

Key concepts associated with atomic structure and the periodic table are listed in the sections ‘Atomic structure and the periodic table’ and ‘Bonding’ in this topic.

Students’ alternative conceptions of chemical substances

Research into students’ ideas about this topic has identified the following non-scientific conceptions:

- Heating a substance always means raising its temperature.
- A conductor is something that keeps things warm.
- Substances that insulate hot objects won’t insulate cold objects.
- Metals attract cold better than wooden objects do.
- Melting and dissolving are the same process.
- Bubbles in boiling water consist of air.
- Bubbles in boiling water consist of hydrogen and oxygen.
- Vapour is different from water.
- Heat is something physical.
- The space between atoms contains air.
- Copper atoms have the same properties as bulk copper. For example, copper atoms can melt.
- Gold atoms are gold in colour.
- Atoms are hard, like billiard balls.
- Atoms have electrons circling them like planets around a star.
- Atoms are like cells with a membrane and nucleus.
- Atoms can reproduce after the nuclei divide.
- There is only one correct model of the atom.
- The nucleus attracts all electrons around it equally.
- Molecules are glued together.
- Chemical bonds store energy.
- Atoms form bonds in order to satisfy the octet rule.
- Electron pairs are equally shared in all covalent bonds.
- Atoms can be seen with a microscope or an electron microscope.
- Molecules of solids are hard; molecules of gases are soft.
- Molecules of solids are bigger than molecules of gases.

Resources

Alphabetical listing of the elements

http://nobel.scas.bcit.ca/resource/ptable/elementsalpha.htm
This site gives the physical and chemical properties of the elements. It also gives the history of the discovery and common uses of the elements. A very good site.

Online periodic table

http://www.webelements.com/
This is a premium site that allows you to access any of the elements. A massive amount of information can be obtained from this site.

Periodic tables of elements

http://wwwndc.tokai.jaeri.go.jp/nucldata/periodic.html
This site has links to sites containing periodic tables of the elements. Very comprehensive; some sites are more advanced than others.

Science education at Jefferson Lab

http://education.jlab.org/index.html

Materials and society

Society uses many materials to enhance people's lives. In ancient times, the materials came directly from the environment ready to use (such as stone and wool) or materials were processed before use (such as bronze and iron). In recent times, there has been a revolution in the manufacture of materials, which began with the production of plastics (more generally 'polymers'). Scientists who work in the production of new materials are called 'chemists' and the science of materials is called 'chemistry'.

There are three aspects that make up the heart of chemistry:

- Chemistry is about making forms of matter that never existed before. From plastics and detergents to contraceptives and anticancer drugs, new materials have had an extraordinary impact on our lives.
- Chemistry is about analysing substances and working out how chemical reactions happen. This has allowed us to control industrial processes, monitor our environment and assess our health needs.
- Chemistry is about trying to make sense of our experiences.

Chemists are very interested in exploring the properties of materials so as to enhance their use for society. Once the properties of a material are known they can be considered as replacement materials or even used to make new products for use by society. For example, plastic bags are now commonly used in supermarkets to replace paper bags; my uncle used to distribute paper bags to shops during the 1960s and 1970s (I wonder if he would have had a job today?). An understanding of the electrical properties of silicon (found in common sand) has led to the revolution in computer technology. New materials are being produced by chemistry on a regular basis and will eventually find their way into our daily lives (if they have not already done so). Some common new materials include Teflon, Kevlar and graphite fibre. How often do we hear on news bulletins about the release of a new medicine that claims to assist in the treatment of such diseases as cancer or AIDS?

Manufactured or artificial substances are quite often referred to in everyday speech as 'chemicals'. There are even notions that chemicals can be dangerous for us; there are debates about food products that are grown with pesticides or growth enhancers in contrast to so-called 'organic' products that are grown without the addition of these synthetic materials. The use of the term 'chemical' for a synthetic material is incorrect from a scientific perspective. Everything is made up of chemicals whether humans had some involvement in its production or not. An organically grown apple contains chemicals as does one grown with the use of pesticides (however, the mix of chemicals may be different).

In considering a new material for a product, we need to identify the properties that make a particular material suitable for a particular application. As an example, consider plastic bags used as receptacles for grocery items. Plastic is a good material to use because it:

- is lightweight
- is strong
- is cheap to produce
- will stretch before breaking (it is elastic)
- is reusable and durable (it will last)
- is chemically inert for most grocery items (it will not react with grocery items)
- is recyclable
- is biodegradable (some plastics only—these plastics react with sunlight and over a period of time will degrade)
- is flexible (it can be made into different shapes; it is malleable)
- can be transparent.

Properties of new materials

Try out one or more of the following activities that explore the properties of new materials.

ACTIVITY: STAYING DRY

This activity is taken from Carrado 1993.

Have you ever wondered how the new disposable nappies that are so thin work quite well? There are tiny beads in the filling that are able to absorb more than 300 times their own weight of water. This activity explores the absorbent qualities of these beads.

You will need:

- **an 'ultra-absorbent' disposable nappy**
- **water**
- **table salt.**

Cut open the nappy and carefully peel away the cotton-like filling. You will notice that it feels gritty. Separate the small gritty beads from the cotton fibres (tweezers and a small kitchen strainer may help, but are not necessary). You should be able to collect approximately half a teaspoon of beads. Pour them into a clear glass, then add approximately half a cup of water and gently swirl, or pour the mixture back and forth between two glasses until it is too thick to pour (if you are able to get distilled, deionised water, it works better than 'hard' tap water). To 'unlock' your gel, sprinkle a little salt on top and stir it into the gel. When the water is released the now syrupy liquid can be washed down the drain.

The superabsorbent beads are a copolymer of polyacrylamide and sodium polyacrylate that can undergo physical changes quickly and reversibly with water.

Other uses for these polymers are for hydro-mulching of plants and removal of water from jet fuels.

Try your experiment again with a drop of food colouring (yellow fits the nappy theme well!).

ACTIVITY: 'TEARIBLE' TISSUES

This activity is taken from Parratore 1994, p. 112.

Have you ever noticed that if you spill something on some types of furniture and carpet they are easy to clean up without leaving a stain? Some fabrics are coated with a very thin layer that repels liquids somewhat (this makes them 'hydrophobic' substances). In this activity you can experiment with the absorbency of tissues with and without water-repellent substances.

You will need:

- **a can of Scotchguard spray (or other water-repellent spray)**
- **facial tissues**
- **two clear cups**
- **two rubber bands**
- **water.**

Open one tissue and drape it over one of the cups. Push the centre of the tissue slightly into the cup, forming a pocket. Secure it to the cup with a rubber band. Slowly pour some water into this pocket and observe what happens. Now spray a new tissue with Scotchguard, let it dry, and repeat the experiment. What happens? Does this tissue behave differently?

The spray forms a coating over the surface of the tissue. The coating is so smooth that any holes present are smaller even than tiny water particles. Therefore, water is not allowed to penetrate through the tissue. You can try this same test on different materials, such as pieces of scrap fabric.

ACTIVITY: ECOFOAM VERSUS STYROFOAM

This activity is taken from Gammon 1994, p. 1077.

Have you ever seen packaging pellets made of foam? Have you ever noticed two different kinds of these pellets? One kind is bright white and S-shaped, and the other is not so white and not so curved. In this experiment we will find out the differences between the two types.

You will need:

- **styrofoam pellets**
- **ecofoam pellets**
- **a cup of water**
- **nail-polish remover.**

The old-fashioned, white pellets are made of styrofoam, which is an expanded version of a polymer called 'polystyrene'. The newer kind is called 'ecofoam' and it is made from corn. The most obvious difference between them in terms of chemical properties is that ecofoam will dissolve in water, whereas styrofoam will not. Put a foam pellet in a cup of water and wait a few minutes. The warmer the water, the faster the ecofoam will dissolve. Can you think of a way to take advantage of this difference for a useful purpose? Of course, the answer is fairly obvious.

Ecofoam was deliberately made as an alternative to styrofoam because it quickly degrades in the environment. It is therefore a more ecologically and environmentally friendly form of packaging. Can you see that ecofoam would not make a very good beverage container, however?

If you have access to acetone (the active ingredient in many nail polish removers), you will observe that styrofoam will 'melt' in acetone, whereas ecofoam will not. This could therefore be one way to reduce the space that styrofoam products take up as waste.

ACTIVITY: TESTING THE TEXTURE OF TOOTHPASTE

This activity is taken from Renee B Adams' experiment at <www.ael.org/eisen/toothps.htm>.

Chemistry is so common that it can even be found in toothpaste. Chemists have worked hard to come up with the perfect stuff. Read the labels—you'll find out all kinds of interesting things. Here you'll find some information and learn some tests you can do to compare different brands.

What are the active ingredients in toothpaste? There is fluoride, of course, either as sodium fluoride or sodium monofluorophosphate. Fluoride reverses the process of tooth decay where acids (especially from sugar) dissolve minerals out of the teeth. There are antibacterial agents such as triclosan to control plaque, and antitartar agents to control mineralised plaque. Other, inactive or inert, ingredients are water, detergents (to loosen plaque), binders (to keep solid and liquid ingredients together), humectants (to keep the toothpaste moist in the tube), flavouring, preservatives (to stop bacteria from growing on the other stuff) and abrasives (for cleaning and polishing).

Using the tests that follow, you will observe, collect data and make informed decisions related to consumer choices.

You will need:

- **toothpicks**
- **four or five brands of toothpaste**
- **a toothbrush**
- **a microscope.**

Prepare a chart listing the brands of toothpaste, with sections for texture by 'touch', 'taste' and 'microscope'. Rub a bit of each brand between your fingers and note whether it feels smooth, gritty, etcetera. Then brush your teeth with each brand and record the texture by taste. Next, using a toothpick, smear some toothpaste on a microscope slide, add a drop of water and put on a coverslip. View the slide with the microscope and draw a picture on your chart of what it looks like. Now compare all the brands for texture, grit and appearance.

Which would you choose and why? Why is toothpaste better than just water to clean your teeth?

What do you suppose the abrasives are? This grit is often silica, alumina, calcium carbonate or sodium bicarbonate. Chemists are able to make toothpaste clean, polish and protect your teeth, plus make it taste good and sit up on your toothbrush!

Consumer science

There is no doubt that we currently live in a consumer society. We are constantly exposed to advertising that exhorts the purchase of a particular product. Quite often the advertisement gives the impression that one product is better than a rival product or that one product has special properties that no other product possesses. Are we to believe advertisements or can we be discerning consumers?

When choosing between products what criteria do you use? For example, the following is a list of some criteria used to compare potato chips (you may think of some that I have missed):

- flavour
- saltiness
- freshness (use-by date)
- size/shape of chip (the bigger the better—the corrugated chips always taste better)
- fat content (watch those hips!)
- carbohydrate content (I need an energy fix fast!)
- additives used (in case of allergies or hyperactivity)
- a fantastic competition you can enter (you can win a car!)
- packaging (my favourite colour is purple)
- cost (do I get a small or large pack?)
- value (cost per weight of chip—where's my calculator?).

The list of criteria that I use to select a product may not be the same list as another person uses. In selecting my product I might give extra weighting to some criteria over others. Other people might weigh the criteria differently.

The test that I apply for each criterion is important. This is an area where we need to be scientific. Is the test a fair test? For example, if I measure the total fat content of a large chip from one brand against a small chip from another brand the test is not considered a fair test. However, I can make it a fair test by using the same sized chip from each brand. I could also make the test fair by considering the weight of each chip (then I can compare the amount of fat per unit weight for each brand).

Other considerations for a scientific test include:

- Does the test give you a measurement that measures against the criteria?
- Does the test produce the same results repeatedly?
- Are the measurements carried out accurately?

Consider the following activities that relate to product testing. Consumer organisations, such as *Choice* magazine, specialise in product testing. We can research the criteria, the criteria weighting and the tests used to determine the merits of one product over another.

Product testing

Teaching note: For each product there is only limited information on which test to use or criteria to test. Do not be restricted by what is written here.

In judging different products you need to consider:

- What criteria are relevant for the evaluation?
- Is the test fair?
- Are the results reproducible?
- Is the method of comparison (scale, addition of scores etc.) appropriate?

ACTIVITY: ERASERS

You will need:

- **a grey lead pencil**
- **five types of eraser.**

A test for the effectiveness of an eraser is to draw a pencil line for each eraser to be tested (the lines must be the same length). When ruling the pencil line to be rubbed out you need to control (i) the strength of the pressure of the pencil to the paper, (ii) the length, width and density of the line, and (iii) the type of paper. Will you compare the lines after a given number of strokes of each eraser, or count the strokes needed to erase the lines completely? What other criteria apart from erasing efficiency might be appropriate?

ACTIVITY: POTATO CHIPS

You will need:

- **a variety of brands of potato chips**
- **some brown paper squares**
- **some brown paper bags**
- **a rolling pin**
- **a breadboard**
- **a jar of water.**

To test for salt content, you could get a person or a number of persons to taste the chips. It would be best to blindfold the people so they do not see the brand. These people decide from least salty to most salty (this test can also be used to test for flavour). As an alternative to the above test you could dissolve the chips into water (what needs to be controlled?). Crush a chip from each brand (samples need to be the same size) and put the crumbs of each into separate containers with 40 mL of water. Add a pinch of salt to another 40 mL of water. Have a clean glass of water on hand. Alternatively, taste the salted water and each sample of chip-water, taking a sip of fresh water in between tastes. Which is saltiest?

To test for oil content, place a chip between two sheets of brown paper on the breadboard, and then crush it with the rolling pin. How much oil appears on the brown paper? Measure the size of the spot using a ruler. Alternatively, place a chip on top of a pile of brown paper squares. Roll it with the rolling pin. How many thicknesses of paper does the oil penetrate? Hold the oil patch over some print or up to the light. How translucent is the patch? Repeat the experiment for other brands of chips.

To test the packaging, examine the package each brand of chip comes in. How is the manufacturer trying to sell the chips to you? What colours are used in the packaging? What is the salt or fat content according to the nutrition label? Is there a toy included? Is there a competition? How easy are the bags to open? How will

you score each of these criteria for packaging? Which brand of chip is considered to be the best according to the packaging? Why?

Rank the criteria in order of importance (you may like to add other criteria, such as value). Which chips would you recommend?

ACTIVITY: BALLS

You will need:

- **a variety of balls.**

Test for the bounciest ball. Does the surface make a difference? Investigate.

ACTIVITY: STICKY TAPE

You will need:

- **a variety of sticky tapes**
- **a mirror or smooth surface.**

Design a test to determine which tape adheres best to the smooth surface. Which tape re-sticks the best? Are there other relevant criteria?

ACTIVITY: PAPER TOWELS

You will need:

- **a variety of paper towels or toilet paper**
- **a glass of water**
- **an eye-dropper**
- **a ruler.**

Which brand absorbs the most water? Use the eye-dropper to place a drop into the centre of the paper and see how far the drop spreads. Alternatively, dip one end of the towel into the water and see how much is absorbed. Which is the strongest towel? Is dry or wet strength the most relevant?

ACTIVITY: GLUES

You will need:

- **icy-pole sticks**
- **scales**
- **a variety of glues.**

What do you expect from good paper glue? To test general-purpose glues, glue two icy-pole sticks together and check the weight needed to break the bond. Glue paper strips together and time how long the glue takes to dry. Alternatively, rate the stickiness of the glue. Does it mark the paper?

ACTIVITY: DETERGENT

You will need:

- **a variety of detergents**
- **microscope slides**
- **Vegemite**
- **butter or oil**
- **a 5¢ piece**
- **an eye-dropper**
- **beakers.**

Make up solutions of detergent and water of comparable strength. Compare the foaming actions. Put a drop of oil on a microscope slide. Devise a method for comparing how easily each detergent removes it. Repeat the same experiment

using Vegemite. How many drops fit onto a five-cent piece (the better the detergent, the fewer drops).

ACTIVITY: WATER RESISTANCE OF CLOTH

You will need:

- **a variety of types of fabric**
- **a beaker of water**
- **an eye-dropper**
- **a ruler.**

Design a test to determine the water-resistance of each cloth.

ACTIVITY: SUGAR SOLUBILITY

You will need:

- **different types of sugar (e.g. brown, coffee, white, cube, castor, icing)**
- **an eye-dropper**
- **saucers**
- **beakers.**

Place a small amount of sugar into a saucer. Count the number of drops of water needed to dissolve the sugar. Alternatively, put ten drops of water onto the sugar and record how much sugar dissolves.

Is there any difference between using warm water and using cold water?

Investigative chemistry

A large part of the work undertaken by scientists is determining the types of chemical substances that are in materials. Drug testing in sport is always current news. In drug testing, samples of urine or blood are taken and tests are undertaken on the samples. ‘Chromatography’ and ‘mass spectrometry’ are used to determine the constituent chemicals in each sample. Both these techniques separate the components in the sample according to the mass (called ‘molecular weight’) of its constituent particles. The scientists match the molecular weights found with those molecular weights of known substances and if matches are found then the names of the chemicals in the sample are known.

Some substances are known to react chemically only with other specific substances. Scientists use this fact to determine unknown substances. For example, geologists know that if a rock bubbles when acid is applied to its surface the rock contains lime.

Another method of determining the composition of substances is by spectral analysis. Pure white light is composed of all the colours of the rainbow. A sample substance is converted to a gas and white light is shone through it. Depending on the type of particles within the sample, certain colours will be absorbed and others will not. The light that passes through the gas is spread out with a prism and the pattern of colours formed is called the ‘spectra’. If there is no gas the spectra will be one continuous strip of colours (like that of the rainbow). With the gas present the spectra will have dark lines. The pattern of dark lines is matched with known particles that give the same patterns.

The activity *Mystery powders* allows students to undertake standard tests on substances that can then be used to determine the composition of unknown substances.

ACTIVITY: MYSTERY POWDERS

You will need:

- **six white powders:**
  - **sugar**
  - **flour**
  - **baking soda**
  - **salt**
  - **starch**
  - **plaster**
- **a container for each powder**
- **a magnifying glass**
- **jars or test tubes**
- **vinegar**
- **a frying pan or saucepan**
- **aluminium foil**
- **iodine.**

Put each powder into a container, remembering to clearly label each container. There are five tests to be done on each of these powders. You should record what happens to each powder for each test.

If you then complete the same tests on an unknown white powder (one of the six but unlabelled) or an unknown mixture (of two or more of the six white powders) and compare your test results with the previous results on the known powders, you will be able to identify the unknown powders. This is the same process that scientists use to determine the composition of unknown substances.

In completing these tests students should never taste a substance.

Test 1: Testing by sight

Examine each powder with a magnifying glass. Record your observations of the appearance of each substance. (You may like to make this exercise harder by using icing sugar and grinding up the salt crystals.)

How is the plaster different from the sugar? Which two powders have the largest particles? What is the shape of the salt crystals?

Test 2: Testing with water

Pour small amounts of water into six jars (or test tubes). Add a different powder to each jar. Record your observations for each powder (e.g. whether it dissolves, bubbles, floats or sinks).

List the powders that are soluble. List the powders that will not dissolve (i.e. are insoluble).

Test 3: Testing with an acid

Pour small amounts of vinegar (acid) into six small jars and add a different powder to each. Vinegar belongs to a group of substances called ‘acids’. Record your observations for each powder.

Are any of the reactions different from those with water? Which powders react most with acid?

Test 4: Testing with heat

Line a frying pan or shallow saucepan with aluminium foil. Place a small amount of each powder in the pan and heat it up. Record your observations of what happens when each powder is heated.

This test can be carried out as a class demonstration if necessary. You can also heat up more than one powder at once (make little aluminium patty pans for each powder and heat all of them at once).

Which powders show the greatest change? Which powders do not change? What happens to starch when it is heated?

Test 5: Testing with iodine solution

(Caution: iodine is poisonous; handle it carefully and keep it away from your eyes.)

Place each powder in a separate jar (or test tube) and add a few drops of iodine solution to each. Record any colour change.

What are the usual colours of iodine solutions? Does the iodine change colour when added to any of the powders? Which ones?

When all the tests have been completed, a summary of the results should be written up in table form. This will make it easier to determine unknown powders and mixtures.

Unknown powders and mixtures

Take an unknown powder and identify it by undertaking each of the five tests described above. After each test decide which of the six substances can be eliminated based on the results from each of your tests.

You may like to make this harder by using an unknown mixture that contains two or more of the white powders.

Make up a collection of unknown powders and mixtures to be tested.

Matter

What is matter?

Objects that take up space and have mass are called ‘matter’. Matter is all around us; in fact, the whole universe is made up of matter. All matter has certain characteristics called ‘properties’ that help us to distinguish one type of matter from another. The different properties of a substance make them suited to particular purposes and determine the way they are used by society. There are two basic types of properties: physical properties and chemical properties.

Physical and chemical properties of matter

Physical properties describe the nature of a substance. They represent those characteristics that can be observed or measured and that do not change how the substance looks. Some physical properties of a substance include colour, density, electrical conductivity, melting point, boiling point, hardness and flexibility. The following table gives an extended list of physical properties of matter.

TABLE 1: PHYSICAL PROPERTIES OF MATTER

Property	**Definition**
Melting point	The temperature at which a solid changes to a liquid
Boiling point	The temperature at which a liquid changes to a gas
Viscosity	A measure of how easily a substance flows
Durability	The length of time a substance lasts
Transparency	The ability of a substance to allow light to pass through it
Opacity	The impenetrability of a substance to light
Hardness	The ability of a substance to resist denting or scratching
Flexibility	The ease with which a substance can be bent
Brittleness	The readiness of a substance to snap when bent
Elasticity	The ability of a substance to return to its original shape when stretched
Malleability	The ability of a substance to be extended or shaped by hammering or pressure with rollers
Ductility	The ability of a substance to be drawn into a wire.
Conductivity of heat or electricity	The degree to which a substance allows heat or electric current to flow through it
Solubility	The degree to which a substance dissolves in water
Density	The mass per unit volume of a substance

(James et al. 1999, p. 6)

The chemical properties of a substance are those properties that involve the ability of the substance to react with other substances to form new substances.

For example, how does a substance react with acids, bases and water—does it burn, melt, fizz, corrode or explode? A piece of paper, when burnt, turns into a black substance. After the flame goes out you can no longer burn the new substance because its chemical properties have changed.

In other topics will explore the processes that underlie chemical change in materials. For example, what happens when paper burns? What is the new black substance that is formed?

Particle model of matter

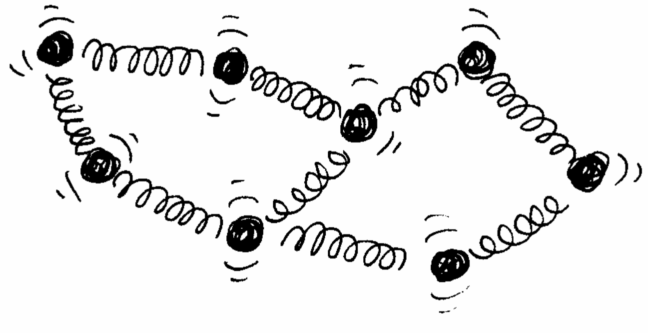
An initial exploration of matter tells us that it is composed of small particles. We know this because most objects can be broken up into smaller and smaller pieces. In some instances we can break an object into pieces that are too small to be seen with the naked eye. We also know that for solid objects the particles are connected to each other in some way. We know this because solid objects retain their shape.

A simple representation of matter in a solid state is shown in Figure 1. The particles are shown by small spheres, and connections between the particles are shown as springs. Later, in more complex models of matter, we will refer to the particles as either ‘atoms’ or ‘molecules’ and to the connections as ‘chemical bonds’. However, for the moment we will restrict our discussion to particles connected by bonds. The particles of matter are in constant vibratory motion. To understand this we need to bring the concept of temperature to our simplified model of matter.

The temperature scale we use today is called the ‘Celsius scale’; some of us may even remember the Fahrenheit scale. Both these temperature scales have a zero value which does not represent the lowest possible temperature as both scales go into the negative. However, a lowest temperature exists, which is approximately –273 ˚C on the Celsius scale. From a scientific perspective, at this temperature the particles of matter are stationary. In other words, they have no motion energy or kinetic energy. The amount of kinetic energy a particle has depends on its speed as well as its mass. For example, a car and truck have the same speed, the truck has more kinetic energy, but if the car and truck are stationary, neither has any kinetic energy. It is important to note that two objects with the same kinetic energy can have different speeds. In this case the object with less speed will have greater mass.

The temperature of an object is related to the average kinetic energy (motion energy) of all the particles contained within it. If the particles within an object are stationary they have no kinetic energy and therefore the temperature of the object is –273 ˚C. However, if an object has a temperature greater than –273 ˚C its constituent particles move. There is another temperature scale that begins at this lowest temperature: the ‘Kelvin scale’. The lowest temperature point is called ‘absolute zero’ and refers to 0 ˚K (Kelvin) or –273 ˚C.

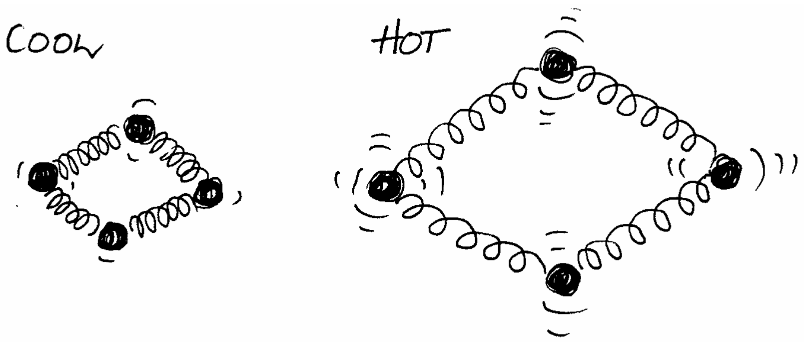
FIGURE 1: PARTICLE MODEL OF MATTER AS A SOLID



From our new understanding of temperature and our simplified model of matter, we can see that as an object increases in temperature its particles will have greater kinetic energy (on average) and so its particles will move faster. As kinetic energy is related to both the speed and the mass of the particles, two substances at the same temperature can have particles at different speeds although their particles have the same kinetic energy. The material with heavier particles will move slower.

The simplified model of matter (see Figure 1) can explain the expansion of solid objects when heated. As the temperature of an object rises, the kinetic energy of its particles increases and so the particles vibrate faster. This in turn tends to stretch the bonds (springs) that constrain the particles. As the bonds stretch, the particles move away from each other and so the object as a whole expands (see Figure 2). Have you ever wondered why railway lines, bridges and concrete slabs have expansion joints?

FIGURE 2: EXPANSION OF SOLIDS THROUGH HEATING



With continued heating of solid objects, the motion of the particles becomes so great that the bonds between the particles begin to break. When this occurs, the object goes through a change of state. This is described in more detail in the next section in this topic.

Changes of state of matter

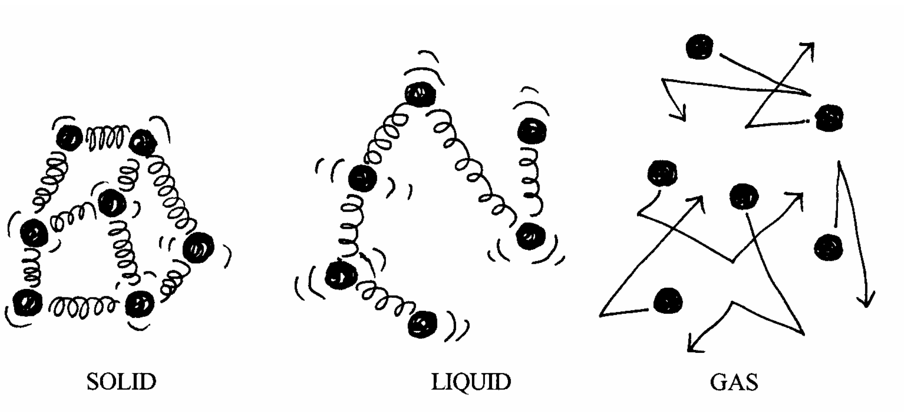
Most matter can exist in three separate states. On Earth there are most commonly three states: solid, liquid and gas. There is a fourth state, called ‘plasma’, which does not readily occur on Earth; this state is discussed later in this topic, in the section ‘Atomic structure and the periodic table’.

The three states of water are ice, liquid water and water vapour. Steam is sometimes considered to be water in a gaseous form, although if we can see the steam then it is in tiny droplets of liquid form. The same can be said for mist and clouds (that is, they exist as tiny droplets of water in a liquid state; they are held in the air by the collisions of air particles in a gas form). We know that the temperature of ice is lower than liquid water and much lower than steam. This means that, from a particle model perspective, the particles are slow in the solid, faster in the liquid and fastest in the gas. In addition, the particles in the solid are constrained by bonds, thus giving solids their shape.

In contrast, there are fewer bonds between particles in a liquid and no bonds between the particles in a gas. (Note: if the particles are molecules, which are atoms bonded together, then these bonds remain intact. In a change of state the bonds between the molecules break.)

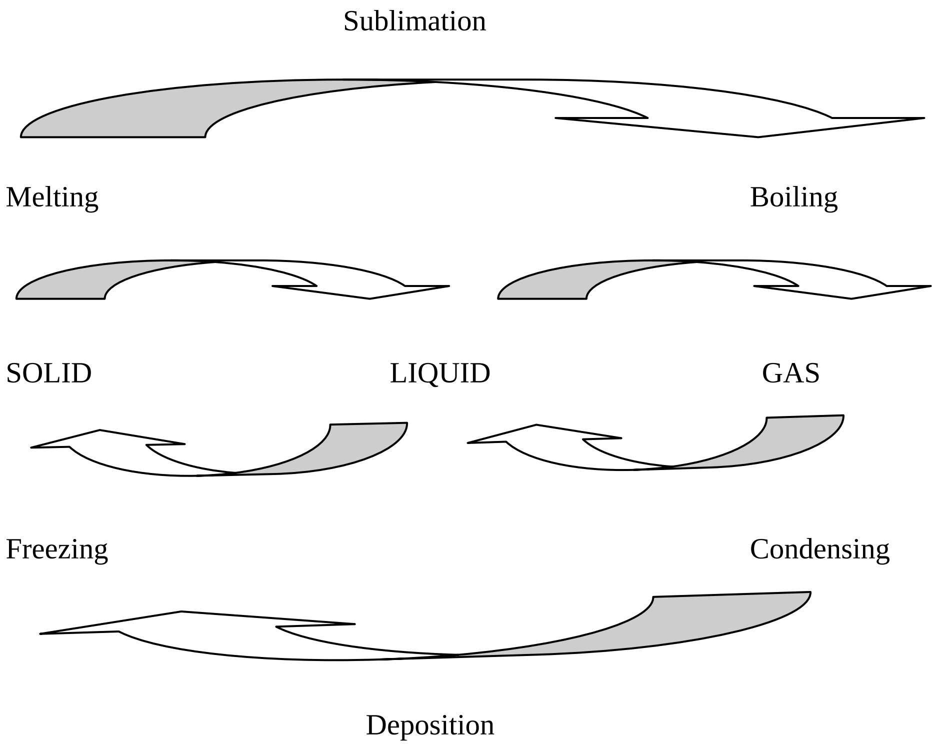
In gaseous form, the particles are free to move and thus take up the whole volume of the container that holds the gas. The three states of matter are shown in our simplified model of matter in Figure 3.

FIGURE 3: PARTICLE MODEL OF THE THREE STATES OF MATTER



The scientific terms that refer to specific changes of state are shown in Figure 4. Note that the change of state called ‘sublimation’ is where a solid changes directly into a gas, bypassing the liquid stage altogether. Conversely, a change of state from a gas to solid, bypassing the liquid state, is called ‘deposition’. An example of sublimation is where solid carbon dioxide (dry ice) is used as ‘fog’ or ‘smoke’ as a theatrical effect. The solid carbon dioxide sublimes to gaseous carbon dioxide. Incidentally, the ‘fog’ or ‘smoke’ that we see is not carbon dioxide gas (this is colourless) but water vapour (gaseous water). The cold carbon dioxide condenses the water vapour in the air to form tiny droplets of water that we see as fog. Other examples of materials that sublime include solid air-fresheners and mothballs. An example of deposition is the production of snow high in the atmosphere.

FIGURE 4: CHANGES OF STATES OF MATTER



ACTIVITY: STATES OF MATTER ROLE-PLAY

This activity will provide students with visual and physical representations for understanding the structure of matter and the processes involved in changes of state. The activity involves a role-play in which the students represent the particles of matter.

To model the solid state (e.g. a piece of chocolate) the students will need to be close together. Each student latches onto another two students (the links represent the bonds). There should be enough room for the students to vibrate (move backwards and forwards) to complete the role-play.

As the solid is heated (the chocolate is in a saucepan on a hot plate) the particles in the solid move faster and the solid changes to a liquid state. In the role-play the students moving faster can show this. In doing so they should let go of one student (breaking one bond).

As the liquid is heated it changes to a gas state. In the role-play the students breaking all connections and moving faster so that they fill the room show this.

Depending on the level of the class, the teacher can introduce the scientific terms used for changes of state (melting, boiling, condensing, etc.). A game can be played in which the students are initially in a particular state (e.g. liquid) and the teacher calls out a change of state (e.g. 'freeze'). The students are then expected to role-play the change of state.

Students need to be aware that for the purposes of the role-play nothing exists between them. Many will say that air does. But this cannot be as they are modelling the particles of matter.

ACTIVITY: STATES OF MATTER

The following set of activities relates to the various states of matter: solids, liquids and gases.

Although the three states of matter are solids, liquids and gases, not all substances seem to fit perfectly into one of these groups.

You will need:

- **shaving cream**
- **a paper towel**
- **a 5¢ piece**
- **a magnifying glass.**

Place a small mound of shaving cream on a paper towel. Look at the shaving cream. Would you call it a solid, a liquid or a gas? One characteristic is that it keeps its shape without being in a container. Does this make the shaving cream a solid? Why or why not?

Very gently place a five-cent piece on top of the shaving cream. What do you observe? Does the shaving cream act like a solid, liquid or gas?

Shaving cream is very light. Look at it very closely or use a magnifying glass if you have one. What do you think makes it so light? Does this make you change your opinion of whether it is a solid, liquid or gas?

Rub a little shaving cream between your thumb and index finger. Does it feel like a solid, liquid or gas?

Leave a blob of shaving cream out overnight. Look at it very closely the next day. How has it changed? Has its state changed? Leave it for a few more days and see if you think it has changed state.

Explanatory note: Shaving cream seems to have an unusual state because it is a liquid soap with a lot of gas bubbles mixed in with it. The gas makes it so thick and frothy that it keeps its shape and supports light objects like a solid. When you let the liquid from shaving cream evaporate, all that's left is the very light and thin solid soap and the spaces where the gas bubbles were. In other words, shaving cream is a mixture of three substances in different states: air in a gas state, water in a liquid state and soap in a solid state.

The line-up

You will need:

- **a number of labelled clear plastic containers that enclose a variety of materials, such as:**
  - **Vegemite**
  - **water**
  - **hundreds and thousands**
  - **shaving cream**
  - **aluminium foil**
  - **wood**
  - **wool.**

Line up the samples provided from 'most solid' to 'most liquid'. Attempt to decide what criteria to use to establish where the line between solid and liquid should go. Discuss how to define 'solid' and 'liquid'.

Explanatory note: The previous activity gives you some idea that it can be quite difficult to decide if a substance is a liquid or a solid, particularly if it is a mixture of both. The general definitions of a solid as having a fixed shape and a liquid as being a continuous flowing substance still apply. Some reference books suggest that glass can be considered a liquid as it is a very viscous fluid; they cite the example of stained glass windows that, over centuries, have slowly flowed so now the glass is thinner at the top and thicker at the bottom.

In completing this exercise you should note that if a solid such as wood is cut or shaved the wood shavings are still a solid. This is sometimes not understood by primary school students.

ACTIVITY: CHANGING STATE

These activities relate to a specific change of state. In explaining the phenomena you should draw on aspects of our simplified particle model of matter.

Water boil

You will need:

- **a pan**
- **water**
- **a heat source.**

Heat up a pan of water and observe the boiling process carefully. What is happening? What are the bubbles? What is happening at each stage. Boil water for 10 min. Where has all the water gone?

Explanatory note: This activity is more interesting than you might think. The first bubbles appear quickly at the edges. These are dissolved air coming out of the solution (a uniform mixture of two or more substances, in this case the water and air). Water always tastes flat when it has boiled because of the lack of dissolved air. It would not be a good idea to use boiled water for a fish tank.

In the next phase bubbles appear where heat has been applied. These grow and detach then disappear before hitting the surface. These bubbles are just water in a gas state (steam!). Many people believe that the bubbles are air (or even oxygen and hydrogen) but this is not true. The bubbles form at the bottom because that is where the hot plate is. From the particle model the fast moving hot plate particles collide with the bottom pan particles, which collide with the water particles. Therefore, the water particles at the heat source increase in kinetic energy first and so change state first.

To convince you that steam is given off, hold a cold spoon above the boiling water. The steam condenses back to liquid water.

Where has the water come from? Fill a can with ice. What happens to the outside of the can?

Explanatory note: This can lead to a discussion of steam, water vapour, bathroom mirrors, visible breath on cold mornings, etc. Students may be sceptical that the water on the outside of the can comes from condensation of gaseous water (water vapour) in the air and so some testing may need to take place (e.g. making the can airtight). For example, to test if the liquid comes from the air, fill two metal cups with ice and a small amount of water. Now, carefully cover one of the cups with a plastic zip-lock storage bag. Squeeze as much air out of the bag as you can and seal the bag. Watch the two cups for a few minutes. Does condensation form on the outside of both cups (condensation is just liquid water, but many people think that it is not)?

Iceblock melt

You will need:

- **an iceblock**
- **a heat source.**

You are given an iceblock and your task is to make as much water in a minute as possible.

Explanatory note: A change of state occurs at the surface of the ice if it is exposed to a heat source. Therefore, if you increase the surface area by crushing the ice and then heating it in some way (placing the pieces in your mouth) it will melt quicker. Pressure also works well. Placing a warm cloth or aluminium foil around the iceblock has an insulating effect and will slow the melting process. Heating the iceblock on a hot plate only heats the iceblock from one side and may not melt it as quickly as you think.

Liquid melt

You will need:

- **an iceblock container**
- **milk**
- **black coffee**
- **lemonade**
- **brine.**

Place the liquids into the iceblock container and freeze. Predict which will melt first. Can you explain any differences using the simplified model of matter?

The great melt

You will need:

- **a frying pan**
- **small aluminium patty pans**
- **a variety of solids (sugar, butter, candle wax, copha, salt, margarine, cheese, aluminium foil)**
- **a timer.**

Predict which of the solids will melt first. How will you know the solid has melted?

Test your predictions and note how long each solid takes to melt. Observe how each one reacts to the temperature change. What happens as each solid cools?

Explanatory note: The solids should solidify (freeze) in the reverse order that they melt. The judgment on melting should be when the substance first begins to melt rather than later, as other factors such as size or heat conduction will affect the results. The butter and cheese will melt, but because they are mixtures of substances (water, fats, milk solids) the melting point is not clean (the different substances have different melting points) and they tend to separate out somewhat. The cheese burns easily and the sugar caramelises to form a toffee, which is a chemical burning reaction. This brings up the question of reversibility. A change of state is reversible (it is a physical change) but most chemical reactions are not.

Salt water

How cool is salty water?

You will need:

- **a bowl of crushed ice**
- **salt**
- **test tubes**
- **orange juice**
- **water**
- **a thermometer**
- **a timer.**

Add salt to a bowl of crushed ice and take temperature readings. Take a test tube of orange juice and a test tube of water and place them in the crushed ice. What happens? Take the test tubes out at intervals and observe the freezing process.

Melting and boiling points

Common physical properties of substances are their melting and boiling points. Incidentally, the melting point is the same as the freezing point, and the boiling point is the same as the condensation point. The following table gives data on some common substances. Chemical scientists use melting and boiling points to determine unknown substances.

TABLE 2: MELTING AND BOILING POINTS

Substance	Melting/Freezing point (˚C)	Boiling/Condensation point (˚C)
Water	0	100
Helium	–270	–269
Oxygen	–219	–183
Silver	961	2210
Gold	1063	2970
Iron	1540	3000
White sand (silicon dioxide)	1700	2230

Evaporation and sense of smell

Our particle model of matter can explain the processes involved in evaporation and our sense of smell. Evaporation is a change of state from a liquid to a gas. This can occur at temperatures lower than the boiling point of a substance (at the boiling point of a substance all the substance changes state). Scents, smells or odours from solids and liquids also represent a change of state from liquid or solid state to a gas state.

As we have already mentioned, the particles of an object are in constant vibratory motion and the object’s temperature is determined by the average kinetic energy of its particles. However, all the particles do not necessarily have the same kinetic energy (temperature is a measure of the average kinetic energy of all the particles in a substance), so some particles will move faster than others. This means that on the surface of most objects the faster particles sometime get free of their bonds and move into the air. This can occur when the object is in a solid or liquid state.

If you can smell an object, like a chocolate bar, then particles that have been dislodged from the chocolate bar have moved up your nose and activated smell receptors. The receptors then send a message to your brain giving you the perception of the sense of smell. You should also note that if you take a chocolate bar from a refrigerator on a warm day you can smell the chocolate better as it heats up, and even more so when it melts. Once again, our model of matter can explain this situation. As the chocolate heats up, the particles of chocolate move faster and so more of them will be freed from the surface of the chocolate. Can you now explain why you get fewer fumes from cutting up an onion if you cool it under the water from a tap or in the refrigerator?

The following activities explore the concept of evaporation and use of sense of smell.

ACTIVITY: EXPLORING EVAPORATION

You will need:

- **a paper towel**
- **a bottle of perfume, eucalyptus oil or spray freshener**
- **methylated spirits**
- **water**
- **a hair dryer**
- **a refrigerator**
- **a cover.**

Place a wet handprint on a paper towel and watch the handprint disappear. Where has the water gone?

How quickly does smell travel? Open a bottle of perfume, eucalyptus oil or spray freshener and investigate the development of the smell throughout the room. How quickly does it spread? Investigate how the smell can be spread more quickly. The particle model suggests that if you can smell the substance then it should be slowly disappearing. Does the amount of substance decrease with time?

Place a drop of liquid (e.g. methylated spirits, water) onto a ceramic tile and investigate how each evaporates under different conditions—for example, in the open air, under a hair dryer, in a refrigerator or under a cover. Which evaporates the liquid faster?

Explanatory note: You can explain the different evaporation rates through our particle model of matter. The particles in different materials may have the same temperature but bonds of different strength. A material whose particles are weakly held will evaporate more quickly than a material whose particles are strongly bonded.

Water from leaves

Place a plastic bag over a group of leaves on a tree and tie the bag in place. Return after about an hour. What can you see? Explain your observation. When would be the best time to collect the most amount of water?

The nose knows!

You will need:

- **perfume**
- **peppermint lollies**
- **two small plastic cups**
- **a paper napkin**
- **plastic wrap**
- **water.**

Place a few drops of perfume in a cup and quickly cover with the paper napkin. Do you smell anything? Wait 1 or 2 min and smell again. Is the smell any stronger? Can you explain your observation (particularly from a particle model perspective)?

Now try placing a piece of plastic wrap over the cup. Smell the outside of the wrap. Do you smell anything? Wait 1 or 2 min and smell again. Is the smell any stronger? Can you explain your observation (particularly from a particle model perspective)?

Place a peppermint lolly in each of the two plastic cups. Smell the air over the lollies in each cup. Can you smell the peppermint? Now add quarter of a cup of water to one of the cups. Swirl the water and smell the air over the cup. Can you smell the peppermint now?

Explanatory note: From these activities you should realise that some particles produce a stronger smell than others. The particles of perfume are small enough to pass through the gaps in the paper towel. The particles from the peppermint lolly need to be dissolved before they evaporate easily.

Physical properties of solids

Most of the objects we see in our daily lives are in a solid state. These solids vary considerably in a number of ways and so just a few properties will be explored in this section.

The density of a solid is related to its mass and how much volume it has (liquids and gasses also have density). The density is the amount of mass per unit volume. For those of you who like formulas, density can be expressed as: density = mass/volume.

For example, a rock weighs 100 g and has a volume of 50 cm^3. The density of the rock is then 100/50 = 2 g/cm^3. Table 3 gives the densities of some common substances (including liquids and gases for comparison).

TABLE 3: DENSITY OF SUBSTANCES

Substance	Density (g/cm^3)
Air (20 ˚C)	0.0013
Carbon dioxide gas (20 ˚C)	0.002
Water (liquid)	1.00
Mercury	13.6
Ice	0.92
Aluminium	2.7
Granite	2.60
Wood (pine)	0.43
Lead	11.3
Sugar	1.54
Salt	2.18
Gold	19.3

(Heffernan & Learmonth 1997, p. 112)

ACTIVITY: FINDING THE DENSITY OF SOLIDS

To determine the density of an object you require two measurements, the mass (in grams) and the volume (in cubic centimetres). Once you know these two measurements, use a calculator to divide the volume into the mass. Your answer will be the density in units of grams per cubic centimetre.

The mass of your object can be easily determined by use of a weighing device such as kitchen scales. To determine the volume of your object can be a little more difficult. If it is a regular shape like a sphere, cube or box (prism) then you can use a mathematical formula. But who wants to deal with formulas? And what if the object is an unusual shape?

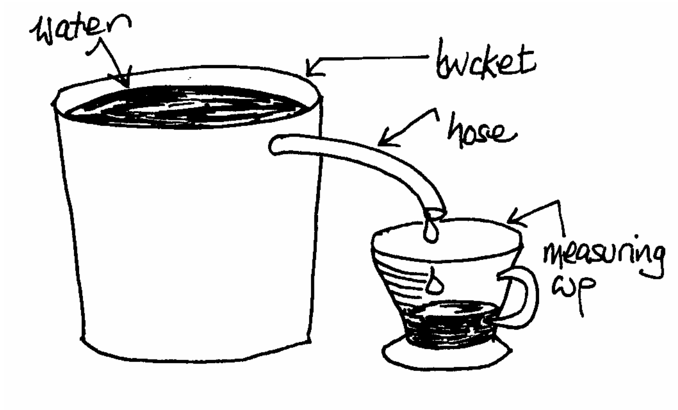
You will need:

- **a bucket**
- **water**
- **a piece of hose**
- **a measuring jug or cylinder (it needs to measure in millilitres).**

Cut a hole near the top of the bucket and insert the piece of hose. Fill the bucket with water until it just begins to leak out of the hose (the hose needs to go to another container, which can be the measuring jug). You have now made an overflow bucket.

To determine the volume of your object, place it in the bucket of water until it is completely submerged (you may need to push it down if it floats). The same volume of water will then flow out of the hose into the measuring cup. The volume of water in the measuring cup (in millilitres) is equal to the volume of the object (in cubic centimetres).

FIGURE: DEVICE TO FIND THE VOLUME OF OBJECTS



Now find the density of various objects and compare the densities found with those listed in the table above.

ACTIVITY: WILL IT FLOAT OR SINK?

An object floats or sinks in a liquid depending on its density relative to the density of the liquid. An object will float if its density is equal to or less than the density of the liquid and an object will sink if its density is greater than that of the liquid. From the table of densities above (Table 3) we see that lead has a greater density than liquid water so a lead bar will sink. However, if the lead bar is shaped so that it contains air (as in a boat) then its overall density is less than that of water and it will float. Can you see from the table why ice floats? Note that sugar and salt solutions are denser than pure water. This fact should explain the observations in the two following activities.

Floating peanuts

You will need:

- **some shelled peanuts**
- **two clear plastic containers**
- **water**
- **salt.**

Fill both containers with water and add salt to one of them. Add a peanut to each container and observe what happens. The peanut floats in salt water but not in pure water.

How sweet it is!

You will need:

- **a can of ordinary cola**
- **a can of diet cola**
- **a tub of water**
- **a straw**
- **an eye-dropper.**

Place both cans in the tub and observe what happens. The ordinary cola should drop to the bottom while the diet cola floats on or near the surface of the water. The densities of the two cans obviously differ. This is because ordinary cola contains approximately 18 g of added sugar.

Fill a clear plastic cup two-thirds full of water and place the straw on an angle all the way in. Use the eye-dropper to carefully drip some ordinary cola through the straw and onto the bottom of the cup. Do not stir the liquids, and observe where the cola stays. Now repeat with a new straw and a rinsed eye-dropper using diet cola. Does the diet cola behave the same way as the ordinary cola?

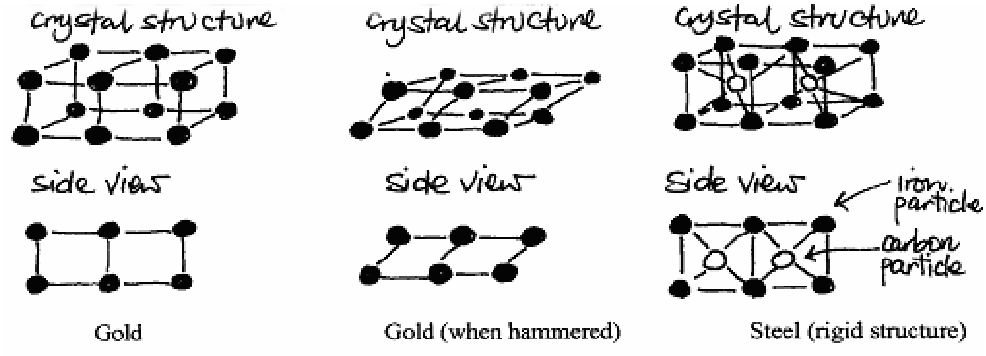
Physical types of solids

All solid materials are described as being either amorphous or crystalline. Crystalline solids contain particles that are arranged in a regular way to form definite shapes. Most solids are crystalline solids. Amorphous solids on the other hand have particles that are arranged in an irregular manner that have no shape or form. Glass is an example of an amorphous solid.

The arrangement of particles and the bonds between the particles explain a number of characteristics of materials. For example, pure metals such as gold or silver have particles that have a bonding structure that can be displaced. The bonds remain intact. This explains why such metals are ductile (able to be drawn into a wire) and malleable (able to be reshaped by hammering or rolling). See Figure 5.

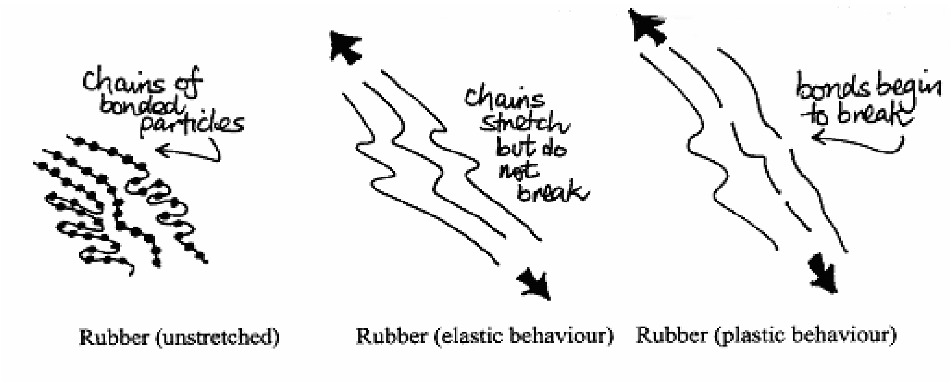
Conversely, metals such as steel are not as ductile and malleable. Steel contains iron particles that have had carbon particles added in a heating process. The smaller carbon particles fit between the particles of iron and form bonds with them. This makes the bonding structure more rigid. This results in a material that is harder (able to resist denting or scratching) and less ductile and malleable. Because the bonding structure is so rigid, if enough force is applied to steel it will not stretch very far before breaking. That is, it is brittle (see Figure 5). Other materials that have a rigid bonding structure and are also brittle include diamonds and concrete.

FIGURE 5: BONDING IN GOLD AND STEEL



Some plastics and rubber substances have particles that are connected by bonds in long chains. When these materials are stretched the chains remains together but begin to unravel. This explains their elastic behaviour, which is the ability to stretch and return to their original shape (see Figure 6). These substances can be over-stretched. When this occurs, the rubber or plastic is permanently stretched; the material has undergone plastic behaviour. When this happens the chains of particles get stretched to the point that bonds begin to break. However, as the chains do not all have the same length, the material does not break all at once (like brittle materials) but gradually when further stretched.

FIGURE 6: BONDING IN RUBBER



Physical properties of liquids

One of the main properties of liquids is that they can flow. Liquid can be poured from one container to another in a continuous stream. This is different from, say, sand, which can be poured from one container to another, but the material is not continuous. We know that some liquids, such as water, flow quite easily, and other liquids, such as tomato sauce, flow quite slowly. Honey is a liquid that changes the manner in which it flows markedly with changes in temperature. Honey spreads quite easily on a hot day but is very stubborn on a cold day or if the honey has been left in the refrigerator. 'Viscosity' relates to how well a liquid flows.

Viscosity is a measure of the 'thickness' of a liquid. Liquids with a low viscosity flow freely, whereas those with a high viscosity are more difficult to pour or stir. The different viscosities are attributed to the size of the particles and/or the strength of the bonds between the particles. For example, petrol (low viscosity) has smaller particles with weaker bonds than oil (high viscosity), which has larger particles and stronger bonds.

The viscosity of a liquid decreases when the temperature of the liquid increases. Honey has already been given as an example. From the particle model of matter, this can be explained by the view that as the temperature increases the particles move faster, which results in the breaking of some of the bonds. Table 4 gives viscosity values for some common materials. Note the decrease in viscosity with an increase in temperature.

TABLE 4: VISCOSITY VALUES

Liquid	**Viscosity coefficients**	
	10 ˚C	**30 ˚C**
Glycerine	2.10	0.35
Water	0.0013	0.00080
Castor oil	2.42	0.451
Light oil	0.15	0.05
Heavy oil	0.60	0.20

(Martin & Connor 1970, p. 158)

Some liquids change their viscosity when a force is applied, such as stirring or shaking. These liquids are called ‘non-Newtonian fluids’ because they do not fit Newton’s laws of how true liquids behave. There are two types of non-Newtonian fluids: ‘rheopectic’ liquids and ‘thixotropic’ liquids.

Rheopectic liquids increase in viscosity under a constantly applied force and are sometimes called ‘stir-thickening liquids’. Examples of these liquids include cornflour and water, starch solutions and quicksand (water-saturated sand). You can stir these liquids quite easily if you use a slow stirring motion. However, if you try to apply pressure by stirring fast you will find it impossible. This is because the viscosity is increased.

Thixotropic liquids decrease in viscosity under a constantly applied force and are sometimes called ‘stir-thinning liquids’. Examples of these liquids are toothpaste, butter, paint, margarine, shaving cream, blood and tomato sauce. When painting you must stir the paint to make it thin enough to spread. The action of the roller or the paintbrush thins the paint further. When you stop, it thickens and sticks. By shaking your tomato sauce bottle you are thinning the sauce (lowering its viscosity) so that it can easily pour from the bottle.

ACTIVITY: VISCOSITY INVESTIGATIONS

The following investigations have been adapted from activities described in Bennett & Gardner 1994.

Measuring viscosity

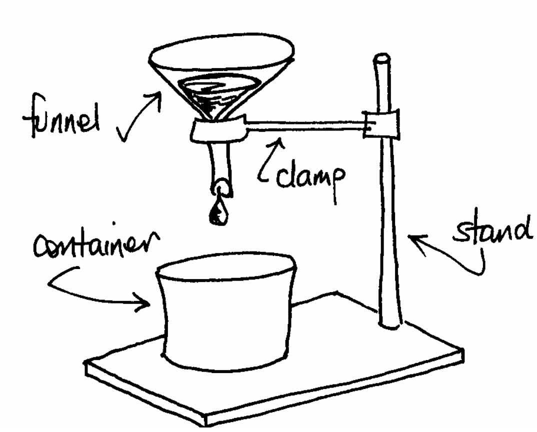
You will need:

- **a funnel with a stand**
- **a container**
- **a watch (preferably a stopwatch)**
- **various liquids.**

Set up the stand as shown in the figure *Testing viscosity*.

Measure the time it takes for a funnel of water to empty. Now record the time for various other liquids, such as cooking oil, honey, sauce, detergent and glycerine.

FIGURE: TESTING VISCOSITY



Temperature change and viscosity

You will need:

- **a jar of honey**
- **a knife**
- **a bowl.**

Dip the knife into the jar of honey, which is at room temperature. Lift the knife above the jar and observe the flow of honey. Predict what would happen to the flow of honey if you raise the knife higher. Test it out.

Raise the temperature of the honey by placing the jar into a bowl of hot water. How high can you raise the knife now? Why?

Repeat both parts again, but this time, find out in which direction the honey coils: clockwise or anticlockwise?

FIGURE: INVESTIGATING THE VISCOSITY OF HONEY

Unbottling tomato sauce

You will need:

- **a bottle of tomato sauce.**

Try pouring the tomato sauce out of the bottle. What happens? Shake the bottle well and try again. What happens?

Explanatory note: Tomato sauce is a stir-thinning liquid.

Cornflour sludge

You will need:

- **cornflour**
- **water**
- **an ice-cream container.**

Mix some cornflour and water into a sludge in an ice-cream container (you may need to experiment). Let it run through your fingers. Now slap the mixture with your fist many times. Is it still runny? Take a handful and add force to it, forming it into a ball. Now stop the force. What happens?

Explanatory note: This mixture is a stir-thickening liquid.

Making a bubble tower

You will need:

- **a glass cup or container**
- **an eye-dropper**
- **various liquids, such as glycerine, honey, clear detergent, water.**

Fill the container with one of the liquids. Put the eye-dropper on the bottom of the container and create different-sized bubbles of air. Which bubbles rise faster, the small or large ones? How does the rising speed of the bubbles relate to the viscosity of the liquid?

Explanatory note: The greater the viscosity the slower the bubbles will rise to the surface of the liquid.

Testing for viscosity

You will need:

- **five test tubes**
- **milk**
- **tomato sauce**
- **water**
- **glycerine**
- **cooking oil**
- **a marble or a paperclip.**

Fill each of the five test tubes with one of these liquids: milk, tomato sauce, water, glycerine, cooking oil. Now find the time it takes for a marble or paperclip to drop to the bottom of each of the test tubes. Which liquid is most viscous?

It's slime time

Slime is made by a reaction between just two compounds. One is a long chain molecule, a polymer called 'polyvinyl alcohol' (PVA). This is cross-linked with a simple solution of borax (sodium borate). Cross-linking means that the long chains are joined to each other at a few points along the chain. Such a process makes the molecules so heavy that they are no longer soluble in water, and a gel begins to form.

You will need:

- **food colouring**
- **a styrofoam cup of water**
- **borax**
- **an icy-pole stick.**

Add a drop of food colouring to a 4% PVA solution in water in a styrofoam cup. Add an equal volume of 4% borax and stir the gel with an icy-pole stick.

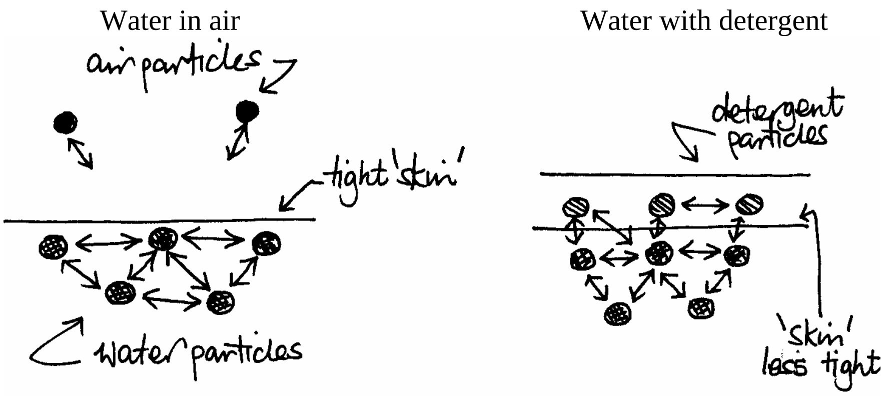
Your slime can be kneaded into an elastic, semirigid glob that has unusual physical properties. If the gel is simply suspended from your hand, it will flow and stretch. It can also be stretched by slowly pulling, but it will break if pulled quickly. When placed in a container the gel assumes the shape of the container. Similarly, it will flow into a film on a flat surface. Because of these physical properties, slime is another example of a stir-thickening liquid.

Bubbles and surface tension

Liquids form droplets. This is because the particles of the liquid are attracted to each other and so form bonds. Inside the droplet or under the surface of a liquid the particles pull towards each other in all directions. However, at the surface the particles are being pulled towards only those particles from the sides and below. (There is some pulling towards the particles in the air but it is not as much as in the liquid. See Figure 7). The consequence of this is that the surface of the liquid is pulled tight, much like a stretchy skin. This effect is called 'surface tension'.

Because of surface tension, water droplets form a sphere, which becomes teardrop-shaped when gravity pulls on it. Surface tension also allows insects and spiders to walk on water.

FIGURE 7: SURFACE TENSION IN A LIQUID



Surface tension in water can be lessened by the addition of soaps, oils and detergents. These substances form a thin layer on the surface of the water. In doing so, the particles on the surface of the water now have particles attracting them from all sides and so the surface 'skin' effect is lessened. The ability of water to form large droplets diminishes with the addition of detergents. Bubbles form when soaps and detergents are added to water because the surface tension is lessened allowing the water to form thin films. Try out some of the following activities that relate to surface tension.

ACTIVITY: INVESTIGATING SURFACE TENSION

Elastic skin

You will need:

- **a glass tumbler**
- **water**
- **pins or paperclips**
- **an eye-dropper**
- **a 20¢ piece**
- **detergent.**

Fill the tumbler to the brim with water. Determine how many pins or paperclips you can carefully add to the tumbler before water spills. Does the water bulge above the glass just before it spills? How can it do this?

Determine how many drops of water from an eye-dropper you can place on a twenty-cent piece before the water spills. Note the bulge in the water as this happens.

Repeat both experiments above to the point where there is an obvious bulge in the water. Now add a drop of detergent and observe what happens. Explain your observations. Alternatively, repeat the experiments but first add some detergent to the water. Are you able to add the same number of pins or droplets as before? Explain why.

Motor boats

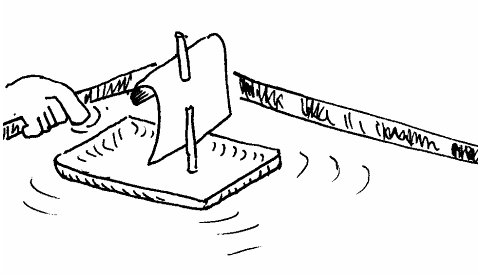
You will need:

- **a polystyrene food dish**
- **a large tray**
- **water**
- **detergent.**

Half-fill a large tray with water and let it settle. Make a boat out of the polystyrene food dish and float it at one end of the tray. Now touch the water between the tray and the boat with a detergent-wet finger. What happens? Why?

Before adding the detergent the surface is drawn tight under tension but when the detergent is added this tension is released, thus spreading the surface. The boat will move with this spreading effect.

FIGURE:
BOAT POWERED BY SURFACE TENSION



Spreading powder and twister

You will need:

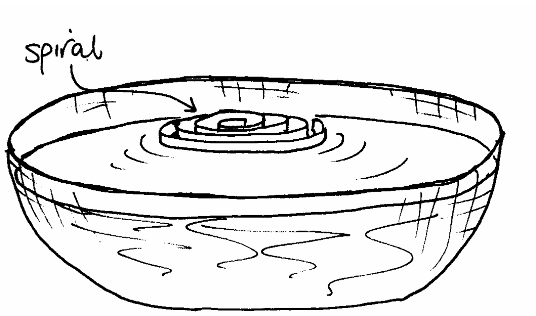
- **a bowl of water**
- **talcum powder**
- **a strip of paper**
- **detergent.**

Fill the bowl with water and sprinkle talcum powder over the surface. Dip a finger in detergent and touch the powder. Observe what happens. Why?

A similar effect can be obtained by floating a thin paper spiral on top of the water (make sure the bowl has been emptied and refilled from the last experiment). Put a drop of detergent in the centre of the spiral and observe. Explain what happens.

These effects can be explained in the same way as the activity *Motor boats*.

FIGURE:
TWISTER



Bubbles, bubbles, bubbles

The following websites about investigating bubbles provide lots of information, including:

- **the best bubble mix recipes**
- **suggestions for various soap bubble 'tools'**
- **the theory of soap bubbles which explains bubble shapes and colours**
- **activities for students.**

Floating Soap Bubbles
http://scifun.chem.wisc.edu/HOMEEXPTS/SOAPBUBL.html

ChemShorts for Kids: giant bubbles
http://membership.acs.org/C/Chicago/ChmShort/cs93.html#1.93

ChemShorts for Kids: science of soap bubbles
http://membership.acs.org/C/Chicago/ChmShort/cs95.html#1.95

ACTIVITY: INVESTIGATING PROPERTIES OF LIQUIDS

Liquid sandwich

You will need:

- **glass containers**
- **glycerine**
- **brine (salty water) or sugar syrup**
- **water**
- **alcohol (e.g. methylated spirits)**
- **vegetable oil.**

Pour each liquid into a tall glass container. These liquids do not mix (they are immiscible) because they have different densities and so will form layers. The effect is more striking if each liquid is coloured with food colouring first.

Liquid line-up

You will need:

- **a series of clear containers with a variety of liquids inside.**

Put the liquids provided into groups according to their properties. How many criteria can you use?

Where did it go?

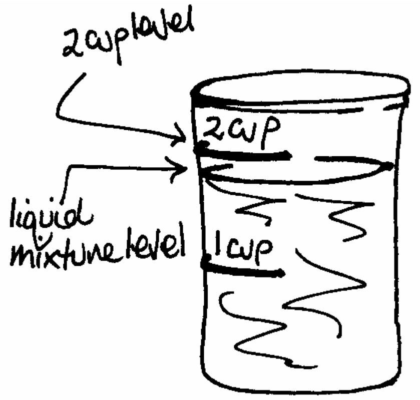
This activity is taken from van Cleave 1989, p. 26. Its purpose is to illustrate that there are pockets of space between water particles (this can be explained by the particle model of matter).

You will need:

- **a measuring jar or cylinder (a tall jar would do just as well)**
- **1 cup rubbing alcohol**
- **1 cup water**
- **a measuring cup (or cylinder)**
- **blue food colouring.**

Pour two accurately measured cups of water into the jar. Mark the water level for the 2 cups. Empty the jar. Add five drops of food colouring to 1 cup of water (needs to be accurately measured) to make the water level easier to see. Pour the coloured water into the measuring jar. Add 1 cup of rubbing alcohol to the coloured water. Observe the height of the liquid. The liquid level should be below the 2-cup mark.

FIGURE: LIQUID LAYERS



Physical properties of gases

There are many substances surrounding us that are in a gaseous state. In general, this is because the temperature of the environment exceeds the boiling point of the substances. For example, the boiling points for oxygen, nitrogen and carbon dioxide are –183 ˚C, –196 ˚C and –78 ˚C, respectively. These three substances form the main components of our atmosphere.

There is a perception that gases are weightless, but this is not true. All gases have mass. If we draw on our particle model of matter, then the only difference between gases and liquids or solids is the bonds that bind the particles together. The particles are still there and it is the particles that have mass. If a gas fills a container, its mass is spread over the whole volume of the container.

Density is the quantity that refers to how much mass there is per unit volume. Therefore, gases can have very small densities compared to their other states. For example, a teaspoon of liquid water has a density of 1 g/cm^3. If you were to boil the water in a room that measures 4 m × 3 m × 2 m the density of the gas would be reduced to 0.00000004 g/cm^3, as all the water particles would now be distributed to all parts of the room. The particles of water would be so spread out that it is no wonder we can't see water vapour, or most gases for that matter.

A physical property of all gases is their ability to be easily compressed. If you hold the end of a bicycle pump you are still able to push the plunger in a measurable distance. However, you can't compress the air in a bicycle pump very much as the pressure inside the pump gets too great. Pressure is the amount of force the gas applies to the sides of its container. From the perspective of our particle model of matter the particles in a gas are in constant motion. This causes the particles to collide with each other and the container that holds them. The pressure of the gas can be imagined to be the number of collisions per unit area that occur. If you reduce the size of the container then the particles have less room to move and so more collisions will occur; the pressure will rise.

Activities relating to the gas state

The following activities confirm that matter in a gas state:

- does exist, even though most of the time it is invisible
- expands when heated
- has mass
- takes up space
- exerts a pressure.

ACTIVITY:
AIR EXPANSION/
CONTRACTIONS

This activity shows how a gas (air) expands when heated and contracts when cooled.

You will need:

- **a half-litre plastic drink bottle**
- **liquid dishwashing detergent**
- **water**
- **a paper or plastic cup**
- **ice**
- **a bowl.**

Make a bubble solution by mixing 2 teaspoons of water with 1 teaspoon of detergent. Lower the end of the bottle into the bubble mixture to create a bubble over the opening of the bottle. Observe what happens to the bubble when: (i) you gently hold your hand around the bottle, making sure you do not squeeze it, (ii) you place the bottom of the bottle in some hot water, and (iii) you place the bottom of the bottle in some iced water. You may need to re-dip the top of the bottle into the bubble mixture after each step.

Explanatory note: The expansion and contraction of the air inside the bottle can be explained by the particle model of matter.

ACTIVITY:
AIR TAKES UP
SPACE

The following activities show that air takes up space and will resist being compressed very much.

Pushing a plastic bag into a jar

You will need:

- **a plastic bag**
- **a jar or plastic container**
- **a rubber band**
- **sticky tape.**

Open out the plastic bag by blowing into it then fix the bag over the top of the container. Attach it firmly with rubber bands and tape, so that it is airtight. Try to push the plastic bag into the jar (without causing any air leaks). What do you think will happen? What did you discover? Was it easy? Could you open the plastic bag more by gently pulling? Try it.

FIGURE:
PUSHING A
PLASTIC BAG
INTO A JAR

Plastic bag cushion

You will need:

- **a polythene bag (a freezer bag bursts too easily) or a bladder from a wine cask.**

Fill the plastic with air and ensure the opening is tightly secured. You have made an air cushion. How much will the cushion hold up? Will it hold a book? Will it hold a person? Try sitting on it. As a variation to this activity, upend a table and sit it on six to eight balloons. These are best semi-inflated and the floor and table must be free of dust. Predict how many students the floating table can support. You will need to be careful to keep the table steady, so one person (the teacher) should be steadying the table at all times. Another variation is to place a balloon under a pile of books. By blowing into the balloon it can act like a hydraulic lift. Investigate how many books can be lifted in this way.

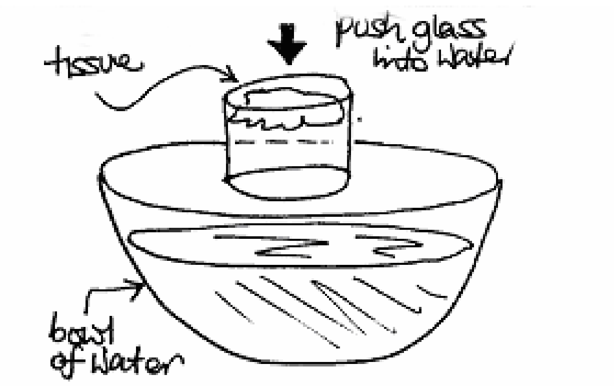
Dunking a tissue

You will need:

- **a tub of water at least 15 cm deep**
- **a glass tumbler**
- **a box of tissues.**

Push some dry tissue paper into the bottom of a glass tumbler, so that it won't fall out when the glass is inverted. Predict what will happen to the tissue if you push the upside-down glass underneath the water in the tub. Do you think the tissue will get very wet? Take the glass tumbler out and feel the paper. Can you explain your observations?

FIGURE:
DUNKING A TISSUE



Explanatory note: The key idea in this activity is that air takes up space; it has pressure.

ACTIVITY:
HOT AIR
EXPANDS

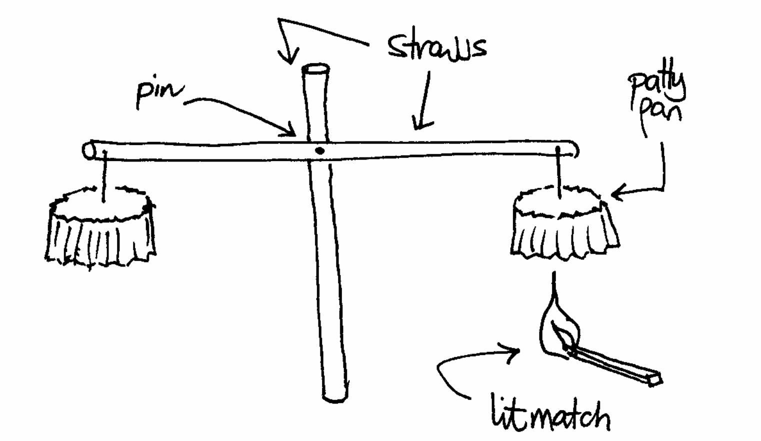
Our particle model of matter suggests that if you heat up matter it expands.

You will need:

- **two small identical patty pans (paper cupcake containers)**
- **two drinking straws**
- **a pin**
- **two short lengths of thread**
- **tape**
- **matches.**

Attach a thread to the centre of the bottom of each of the patty pans. Hang the patty pans upside down by attaching them to either end of a straw. Attach the straw holding the patty pans to another straw through its centre with the pin (see the figure *Hot air expands*). Make sure the straws move freely around the pin. Let one person hold the vertical straw, strike a match and hold the flame under one of the patty pans. (Be careful not to set the patty pan alight.) Watch the pan move. Take away the flame and the equilibrium will be restored. Hold the flame under the other patty pan and observe a similar movement.

FIGURE: HOT AIR EXPANDS



Explanatory note: When the air is heated the particles within it move faster and tend to move further away from each other. This results in an expansion of the air when heated. As the heated air is less dense than cooler air, it will rise up. This pushes one of the patty pans upwards.

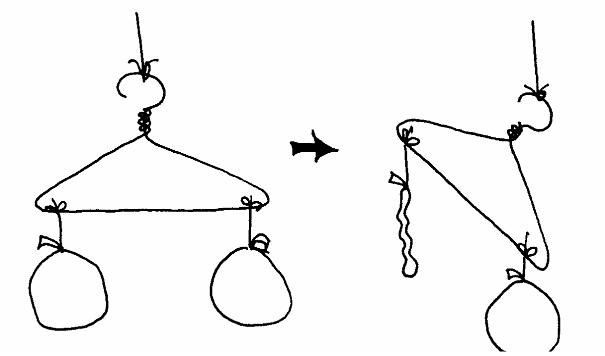
ACTIVITY: AIR HAS MASS

You will need:

- **two balloons**
- **a wire coathanger.**

Blow up two balloons and tie them to either end of a wire coathanger. Hang the coathanger so that the balloons are free of any impediments. Gently put a small hole in the top of one of the balloons, near where it is attached to the coathanger. As air escapes from the punctured balloon, the full balloon will pull the coathanger down on that side.

FIGURE: AIR HAS MASS



Atomic structure and the periodic table

To come to some understanding of the chemical processes that occur in nature and artificially in laboratories and industries, first you need to understand the microscopic structure of matter to a greater level than by thinking of matter as being composed of small spherical particles. However, we have seen in the last section that many phenomena can be explained by use of the simplified particle model of matter—for example, changes of state in matter, evaporation, your sense of smell, the expansion of solids, liquids and gases with heat, and the surface tension in liquids.

Model of the atom

As has already been established matter is composed of small particles. The early Greeks called these small particles 'atomos', which means 'indivisible'. In modern times, we have adopted the name 'atoms' to describe these tiny bits of matter.

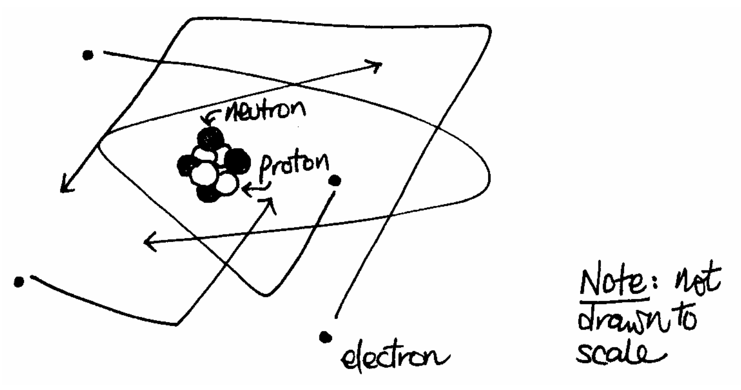
Atoms were first thought (by early Greeks) to be little spheres that simply could not be divided. However, we have a different view today. While we are unable to 'see' the individual atoms that make up matter, scientists have a different representation, or model, of atoms, which varies significantly from solid spheres. By undertaking the task below you will get some insight into a better model of the atom than we have had previously (the particle model). However, bear in mind that this is just another model: a representation of what scientists believe the smallest pieces of matter are like. This model will be able to explain quite a lot, particularly chemical reactions and bonding, but it does not explain all that we know about matter. For more complex phenomena about matter scientists have other models.

Key concepts of atomic structure

The following concepts relate to the structure of atoms:

- All matter consists of extremely small particles called 'atoms'.
- Atoms are mostly space (vacuum).
- Atoms consist of a central nucleus composed of protons and neutrons surrounded by electrons (see Figure 8).

FIGURE 8: MODEL OF THE ATOM



- Protons are subatomic particles that are contained within the nucleus of the atom and have a positive electrical charge.
- Neutrons are subatomic particles that are contained within the nucleus of the atom and have an electrically neutral charge.
- Electrons are subatomic particles that surround the nucleus of the atom and have an electrically negative charge. (We shall see later that electrons are important in understanding chemical bonds and chemical reactions.)
- An atom gets its name according to the number of protons it has in its nucleus. For example, a hydrogen (H) atom has one proton in its nucleus whereas a carbon (C) atom has six protons in its nucleus. The number of protons an atom of an element has is called its 'atomic number'.
- An element is a substance that has the same type of atom contained within it. That is, each atom of an element contains the same number of protons in its nucleus. There are very few materials that we come across in our daily lives that are pure elements but most materials have more than one type of atom. Examples of pure elements include gold, copper, iron, carbon (diamond), oxygen and nitrogen.
- Isotopes of a particular element have the same number of protons in their nuclei but differ in the number of neutrons that are present. For example, two isotopes of carbon each have six protons in their nucleus (this makes them carbon atoms) but one isotope has twelve neutrons while the other has thirteen neutrons.
- Atoms are electrically neutral, which means that they have the same number of positively charged protons as they have negatively charged electrons.

Elements are often written as symbols—for example, aluminium (Al), gold (Au) and carbon (C). (I remember these by reciting the mnemonic 'Happy Healthy Little Beggar Boys Catching Naughty Oversized Fish'. The first nine elements in order of weight are hydrogen, helium, lithium, beryllium, boron, carbon, nitrogen, oxygen and fluorine).

Plasma as the fourth state of matter

It is commonly known that there are three states of matter: solid, liquid and gas. We are now in a position to describe a fourth state, called 'plasma'. Plasma is a very hot state of matter in which electrons have been stripped from atoms to leave positively charged ions (nuclei), which mingle freely with the electrons. All the matter inside a star is in the form of plasma. Plasma occurs on Earth around lightning and within nuclear explosions and so is quite rare.

Electron arrangement within the atom

How electrons are arranged within the atom is important as it explains a lot about the chemistry of elements. The main concepts to understand here are:

- There are regions around the atom that scientists refer to as shells or energy levels in which the electrons move.

- The electrons that move in the shells or energy levels closest to the nucleus require the greatest energy to be removed from the atom. Electrons in the shells or energy levels further out from the nucleus require less energy to be expelled.
- The arrangement of electrons in shells or energy levels around the nucleus is called the ‘electronic configuration’.

Rules for filling shells or energy levels

When considering the electronic configuration of an atom we use the following rules:

- Electrons fill the innermost shell first (only two electrons can occupy this shell or energy level).
- The second shell is filled next; it can contain only eight electrons.
- The third shell is filled next; it can contain only eighteen electrons. However, there is an overriding rule that the outermost shell cannot have more than eight electrons in it. Therefore, if there are, say, ten electrons to fill the third shell, then eight will fill this shell and two will occupy the fourth shell.

The electrons in the outermost shell of the atom are called ‘valence electrons’. These valence electrons are responsible for many of the chemical properties of matter. (This is why I have explained the concepts and rules above.)

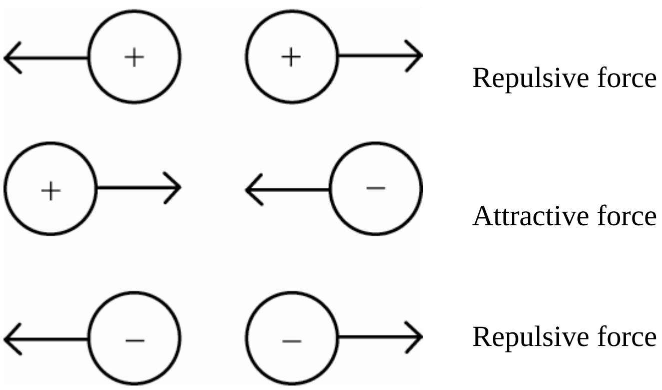
A very important concept in bonding and chemical reactions is that any atom is considered stable (i.e. it will not try to react with another atom) if its outer shell is full or has eight electrons in it. This is why some atoms will attach themselves to (bond with) other atoms. In doing so they take or give electrons to other atoms or share electrons so that they have a full outer shell or eight outer-shell electrons. This will be explained in more detail in the next sections.

Ions and electrostatic forces

To understand bonding between atoms you need to understand how electrostatic forces arise and how ions are formed. It has already been mentioned that if an atom has equal numbers of electrons and protons it is considered to be electrically neutral. However, an atom can gain and lose electrons from its outer shell and in doing so become overall electrically charged. Ions are electrically charged atoms; they have an unequal number of electrons and protons. Positive ions have fewer electrons than they have protons, whereas negative ions have more electrons than they have protons.

Forces exist between charged particles. Particles that have the same charge will be forced away from each other (they have a repulsive electrostatic force). Particles that have opposite charge will be forced together (they have an attractive electrostatic force). These forces are shown in Figure 9.

FIGURE 9: ELECTROSTATIC FORCES ON CHARGED PARTICLES



The charged particles can be protons, electrons or ions. Protons stay together within the nucleus (even though protons repel each other) because there is a much stronger force holding them together (nuclear force). As the protons are constrained within the nucleus they do not take part in chemical reactions (except for the hydrogen atom nucleus as it contains only one proton). Reactions do occur within the nucleus (e.g. in nuclear bombs and radioactivity) and are considered to be nuclear reactions, but these are considered as part of the study of nuclear physics and not chemistry.

ACTIVITY: HOW BIG IS AN ATOM?

The following activity will give you some idea of how small an atom is.

You will need:

- **a calculator**
- **a large cooking tray**
- **dishwashing detergent**
- **an eye-dropper**
- **a small measuring cylinder (with measurements in 1 mL divisions)**
- **a toothpick**
- **a ceramic tile**
- **a cloth**
- **talcum powder.**

Fill the tray with a 1 to 2 cm of water and lightly sprinkle the top of the water with talcum powder. Fill the eye-dropper with dishwashing liquid and place one drop onto the tile. Now put the tip of the toothpick into the droplet of detergent and then put the tip into the centre of the tray of water. You should note that the tiny amount of detergent spreads the talcum powder into a large area (this is because is the detergent lessens the surface tension of the water; refer to the discussion of bubbles and surface tension in the section 'Matter' in this topic for more details).

To get an estimate of the size of an atom we can assume that the small toothpick drop of detergent spreads out on the water to one layer of atoms thick. Therefore, if we can work out the thickness of the layer of detergent then we know the thickness of an atom. To complete this calculation, follow the procedure below.

Step 1

Find the volume (in mL) of a drop of detergent from the eye-dropper. To do this, work out how many drops make up 10 mL using the measuring cylinder. Now

divide the number of drops into 10 mL to find the volume of one drop. For example, if there are 25 drops in 10 mL then the volume of one drop is 10 mL/25 = 0.4 mL. As 1 mL = 1 cm^3, the volume of the drop is 0.4 cm^3.

Step 2

To find the volume (in cm^3) of a toothpick droplet of detergent, you need to find the number of toothpick droplets in one drop of detergent from an eye-dropper. Dip the toothpick into the drop of detergent on the tile. Now wipe the toothpick with the cloth. Dip the toothpick again into the drop of detergent and again wipe it clean. Repeat until there is no more of the drop left on the tile. Count how many times you were able to do this. To calculate the volume of one droplet divide the number of droplets into the volume of a drop from step one. For example, if there are 32 droplets, the volume of one droplet is 0.4 cm^3/32 = 0.0125 cm^3.

Step 3

To find the area of the detergent spread on the tray, you need to estimate how many square centimetres the area of detergent spread out on the tray. Let's assume it is 200 cm^2. Given that the volume of the droplet of detergent on the water equals the thickness of the layer times the area, then the thickness of the layer equals the volume of the droplet divided by the area. In our example, the thickness of the detergent layer = 0.0125/200 = 0.0000625 cm. Therefore, based on our assumption that the layer of the detergent is one atom thick, the diameter of an atom is 0.0000625 cm.

Explanatory note: This method assumes that the layer of detergent on the water is one atom thick. Which is not true. However, the exercise does show you that the atom must be extremely small in size. The diameter of atoms, determined by scientists, varies from atom to atom. For example, a carbon atom has a diameter of 0.0000000077 cm, whereas a rubidium atom has a diameter of 0.0000000216 cm.

Sections of the periodic table

An enormous amount of information can be obtained about the physical and chemical properties of elements if the elements are organised in a specific way. Scientists have found it very useful to organise the elements into a specific tabular format, which has become known as the periodic table. In most periodic tables you will find the following information given for each element:

- Symbol: Generally the symbol can be obtained from the first one or two letters of the name of the element, but not always.
- Atomic number: This gives the number of protons in the nucleus of the atom. The number of protons in the nucleus gives the atom its name. The number of neutrons can vary but these different types of atoms (isotopes) still have the same atomic name. In addition, the number of electrons can vary but these different atoms (ions) still have the same atomic number and therefore the same atomic name.
- Relative atomic mass: This gives the average weight of the isotopes of the element compared to the (approximate) weight of a hydrogen atom. For example, an average atom (isotope) of lead (Pb) is 207 times heavier that an atom (isotope) of hydrogen (H).

- Boiling point and melting point: Note that for some of the very heavy elements the melting and boiling points are not known. This is because they have been made in the laboratory and are very unstable and so don't stay around long enough to determine these points.
- Electronegativity: This concept will be explained in more detail in the section 'Bonding'. It is a measure of the relative attraction the element has for electrons in a bond.
- Electron structure: The coding refers to a more complex model of electronic configuration. For our purposes we only need to consider the 'group number' of the element. For example, carbon is in Group IV, so will have four valence electrons. If the element resides in the d Block or f Block, refer to the small superscripted number at the end of the code. For example, iron (Fe) has the code $[Ar]3d^6 4s^2$. The small '2' at the end of the code means this element has two valence electrons. You will find that most of the d and f Block elements have two valence electrons.

Key concepts of the periodic table

- The periodic table is organised into a table of elements of increasing atomic size where the vertical columns are called 'Groups' and the horizontal rows are called 'Periods'.
- The Groups, labelled I, II, III, IV, V, VI and VII, indicate how many valence electrons (outer-shell electrons) there are. All the elements in these groups have full inner shells (or an unfilled shell of eight electrons). There are blocks of elements called the 'transition elements' (d Block elements on your periodic table), lanthanides and actinides (f Block elements on your periodic table), most of which have two valence electrons but have unfilled inner shells.
- The Periods, labelled 1, 2, 3, 4, 5, 6 and 7 indicate the shell the valence electrons occupy.
- The position of an element on the periodic table depends on the number of valence electrons it has and the shell it occupies (another consideration is if the inner shells are full or not). For example, the electronic configuration of argon (Ar) is 2, 8, 8. All inner shells are full, so this element is in Group VIII (eight valence electrons) and the 3rd Period (valence electrons occupy the third shell). Silver (Ag) has the electronic configuration 2, 8, 18, 17, 2. Silver has an unfilled shell that does not contain 8 electrons and so it is a transition element with a period of five because the two valence electrons occupy the fifth shell.
- Of the elements in the periodic table 80% are metals; the rest are considered non-metals.

The elements in the periodic table have similar physical and chemical properties within groups. More will be said about this in the next section.

Elements, compounds, molecules and mixtures

In your readings you will have come across the three terms ‘elements’, ‘compounds’ and ‘molecules’. Elements are substances that contain the same type of atom. Compounds in contrast are substances that contain atoms from different elements chemically combined in some way (atoms have bonds). While there are only ninety naturally occurring elements (and twenty or more artificially made elements), there are millions of compounds.

The scientific name given to a ‘compound’ gives a clue to its constituent elements. For example, ‘dihydrogen oxide’ indicates that the compound contains hydrogen, and ‘oxide’ refers to the element oxygen. Compounds have symbolic formulas that not only give the elements present but also the proportions. The compound dihydrogen oxide has the common name water and has the formula H_2O. This means that water consists of hydrogen atoms and oxygen atoms. The subscripted number ‘2’ means that for every atom of oxygen there are two atoms of hydrogen.

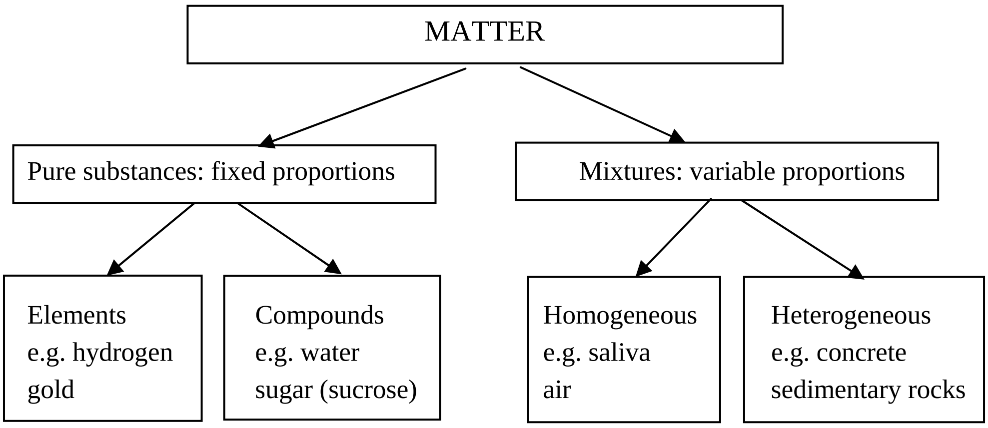
The formula for compounds does not give an indication of how the atoms are bonded. (More details will be given in the section ‘Bonding’.) Bonds can connect atoms in a three-dimensional lattice arrangement, as occurs in crystals like salt. The chemical name for salt is ‘sodium chloride’ (NaCl), which shows that salt has equal numbers of sodium and chlorine atoms. Alternatively, there can be strong bonds between a small number of atoms that form what are called ‘molecules’. In a solid, molecules are weakly bonded to each other. These weak bonds break during a change of state but the strong bonds within the molecules remain intact. Some examples of molecular compounds include, silicon dioxide (sand, SiO_2), ammonia (NH_3), aspirin ($C_9H_8O_4$) and hydrogen gas (H_2).

Broadly, there are two basic kinds of matter. These are pure substances and mixtures. Pure substances are elements or compounds. These are chemically bonded in fixed proportions. For example carbon dioxide (CO_2) is a pure substance (compound), as carbon atoms are bonded to oxygen atoms in a fixed proportion of one carbon atom to two oxygen atoms.

Mixtures such as air, blood, wood, sea water and petrol are each made up of a number of elements and/or compounds combined physically. The proportions of the elements or compounds that make up the mixture can vary considerably. For example, sea water is a mixture of water (H_2O) and sodium chloride (NaCl). The proportion of salt to water can vary. Mixtures can be heterogenous or homogeneous. A homogeneous mixture is one where the particles present are well mixed even on the microscopic level (for example, sea water). Alternatively, a heterogenous mixture is one where the constituent particles are easily identified (for example, concrete).

The classes of matter are shown in Figure 10.

FIGURE 10: CLASSES OF MATTER



Bonding

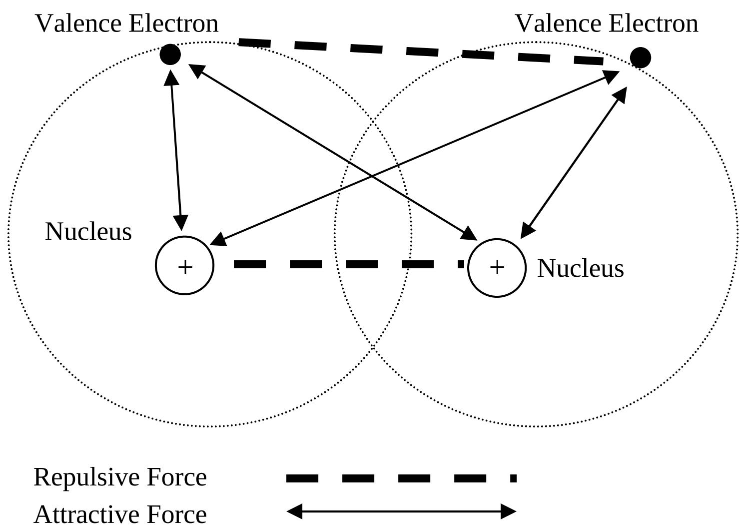
The particles of matter we now call ‘atoms’ are very often connected to each other. In other words, atoms have bonds between them to form structures. From the previous section we have learned that the nucleus of an atom is positively charged (contains protons) and the electron(s) that surround the nucleus are negatively charged. We have also learned that electrical forces exist between charged particles. It is these electrical forces that form the basis of chemical bonding between atoms.

Chemical bond

Electrons and the nucleus are the parts of the atom that are held together by the attractive forces between the negatively charged electrons and the positively charged nucleus. However, the model of the atom discussed previously does not explain why the electrons do not get sucked into the nucleus or why electrons are in different shells or energy levels. In addition, the model does not explain why the nucleus doesn’t fly apart because of the repulsive forces that exist between the positively charged protons. For this course of study we need to be content with thinking that there are other forces at play here. (There are more complex models of the atom that explain these situations.)

If two atoms are in the same vicinity there are electrostatic forces between neighbouring electrons (the valence electrons have the most effect as they occupy the outer shell) and nuclei. There are forces of attraction between the positive nucleus of one atom and the negative electrons of the other. However, the forces of repulsion between the electrons and the nuclei of the two atoms counterbalance these (see Figure 11). Because of the arrangement of the electrons (one can imagine them as a dispersed ‘cloud’), the net attractions outweigh the net repulsions between the atoms and they bond together.

FIGURE 11: FORCES BETWEEN ATOMS



The strength of the bonding is determined by the properties of the two participating atoms. It ranges from extremely weak, such as between two atoms of helium, to extremely strong, such as the bonding between carbon atoms in diamond or between nitrogen atoms in a molecule of nitrogen (N_2). Bonding between molecules of nitrogen (N_2), in contrast, is extremely weak.

An understanding of the relative strengths of the electrical interactions (or bonding) of atoms is very important because it explains the physical and chemical properties of matter. For example, substances with strong bonding are solids with high melting points; bonding is weak in liquids and very weak in gases.

Electronegativity and bonding

The attraction by one atom to electrons in another atom is called the ‘electronegativity’ of the atom. If you refer back to a periodic table of the elements, you will notice that one of the items of information on each element is the electronegativity. Atoms with high electronegativities are able to attract electrons easily, whereas atoms with low electronegativities do not attract electrons as readily. Most elements have electronegativities and hence form bonds with other elements and so do not exist as single atoms. Conversely, the Group VIII elements do not have electronegativities and so exist as single atoms (this is mostly true, although some compounds of Group VIII elements are known to exist).

Previously, we looked at the way chemists arrange elements into the periodic table according to the number of electrons in their outer shell. We have also mentioned the stability associated with atoms having eight electrons in their outer shell. Those elements naturally occurring with eight outer-shell electrons, or valence electrons, (neon, argon, xenon, etc.) are most reluctant to be involved in any chemical activity at all. (They are called ‘noble’ gases or inert gases because of this and are considered to have no electrovalency). All other elements seek to react with other elements in such a way as to attain atoms with the chemically stable state of eight outer-shell electrons. Chemists call this the ‘stable octet’. They talk about the atoms having ‘a filled outer shell’. Many books refer to it as the ‘octet rule’. The ‘Chem4Kids’ website <http://www.chem4kids.com> calls them ‘happy atoms’. That is, atoms are ‘happy’ when outer shells are full or have eight electrons in them.

There are essentially three ways which elements can reach the stable state. They can do so by:

- losing electrons (for metallic elements)
- gaining electrons (for non-metallic elements)
- sharing electrons (between metallic and non-metallic elements).

Gaining and losing electrons leads to ionic bonding. Sharing electrons leads to either covalent bonding or metallic bonding. There is a fourth type of bonding that occurs between molecules, called ‘intermolecular bonding’.

Each type of bonding is discussed in detail in separate sections below. An understanding of the concepts contained in this section is fundamental to an understanding of the nature and properties of all matter and the chemical and physical changes matter undergoes. Keep in mind the important concept that all chemical bonds are electrostatic in nature. That is, positively charged objects (ions, atomic nuclei) are attracted to negatively charged objects (electrons, ions).

Ionic bonding

From atoms to ions

The concept of electronegativity is a crucial one. Do you understand how the electron shell diagrams are produced? These give a pictorial representation of the arrangements of electrons in shells within the atom. Note the link between the electrovalency of the ion and the position of the element in the periodic table.

The key concepts are:

- Atoms of elements become stable if they have full outer shells (or eight electrons in their outer shell—the ‘octet rule’).
- Atoms become stable by losing, gaining or sharing electrons with other atoms. They form chemical bonds.
- If atoms lose or gain electrons they become ions (cations are positive and anions are negative).
- The electrovalency of an ion is the number of electrons lost (negative number) or gained (positive number). The electrovalency of an element can be predicted from its Group in the periodic table (i.e. the number of valence electrons).

Bonding between ions

The examples, first of all, explain in what circumstances some atoms will lose electrons and others gain electrons. A transfer of electrons will not occur unless all atoms in the interaction become stable. The examples also explain why a chemical bond forms. Positive charges are attracted to negative charges. In this situation positive ions (cations) are attracted to negative ions (anions).

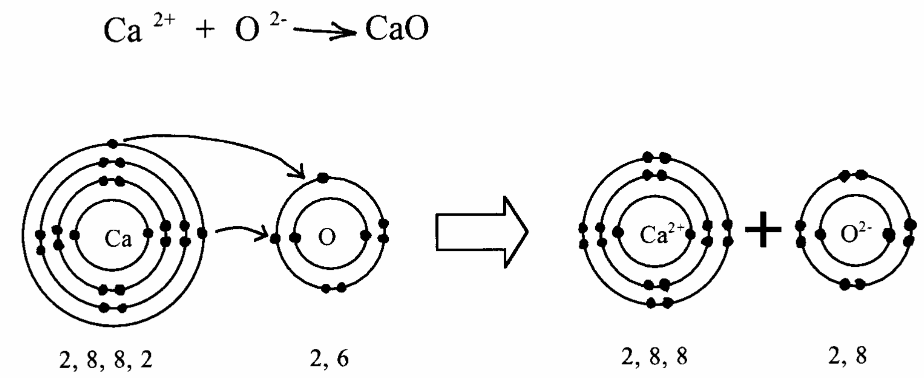
The key concepts are:

- Atoms become ions to become stable.
- Positive ions attract negative ions. The attractive force is called an ‘ionic bond’.
- The attractive force is also called an ‘electrostatic force’.
- Ionic bonding occurs between ions.

Substances that have ionic bonding are called ‘ionic compounds’.

Note the use of electron shell structures to explain the bonding. For example, Figure 12 shows the shell structures of calcium and oxygen in the formation of calcium oxide.

FIGURE 12: ELECTRON SHELL STRUCTURES FOR CA AND O IN THE FORMATION OF CAO



Structure and properties of ionic substances

The key concepts are:

- Ionic bonds are strong. This explains why ionic compounds are solids with high melting points that are hard and not easily scratched.
- Ionic bonds are directional. This means that they have a fixed lattice structure and are crystalline in nature. Ionic compounds are brittle and shatter easily when impacted. This explains why crystal structures shatter. As force is applied to the crystal, the lattice is distorted so that ‘like’ charged ions line up, bringing repulsive forces into play. In this arrangement many, many bonds will break at once giving a shatter effect.
- Ions have different sizes. Because of the range of sizes of the ions there can be different packing patterns. Thus, there is a range of crystalline forms of ionic compounds. This explains the different crystal shapes found in nature. Sodium chloride is said to have a ‘cubic close packed’ crystalline structure.
- Ions in an ionic compound are free to move around in the liquid state (the crystal structure breaks down because of the vibrating motion of the ions).

But this occurs only at high temperatures. In the liquid state the ions are free to move around and so will easily conduct electricity.

Many ionic compounds will dissolve in water. In this situation the crystal structure breaks down and ions are free to move around in the water. The water molecules are responsible for breaking the ionic bonds. A fuller explanation of this is given in the topic ‘Water’. The solution produced will conduct electricity (ions are free to move) and the solution is called an ‘electrolyte’.

Exceptions to the rule

There are some exceptions to the rules governing the type of ion an atom becomes. So far, the requirement is that an atom gains or loses electrons to have a full outer shell (or have eight electrons in it). This is true for Group I, II, III, V, VI and VII elements. These elements can gain or lose up to three electrons. However, Group IV elements do not lose or gain four electrons to form ions. In addition, there are some transition metals, such as chromium and lead, which can form different types of ions. For example, the different ions of iron (Fe) are Fe^{2+} and Fe^{3+}.

Another exception, which is not explained by the discussions given thus far, is the existence of polyatomic ions. Polyatomic ions are groups of atoms that are strongly bonded and behave as single units with an overall charge. Some examples of polyatomic ions are the dichromate ion ($Cr_2O_7^{2-}$) and the ammonium ion (NH_4^+). These exceptions cannot be predicted from looking at the periodic table.

Naming ionic compounds

The electrovalency of an atom or a group of atoms gives the numerical charge on the ion produced. It is a positive number if the ion has a positive charge and a negative number if the ion has a negative charge. The name of the ion remains the same as the element if it is positively charged (cation) but changes if it is negatively charged (anion). For example, oxygen changes to oxide, and nitrogen changes to nitride.

Table 5 lists some common cations and anions and their electrovalencies.

TABLE 5: ELECTRO-VALENCIES OF ANIONS AND CATIONS

Cations +1	**+2**	**+3**	**+4**
Sodium Na^+	Magnesium Mg^{2+}	Aluminium Al^{3+}	Tin (IV) Sn^{4+}
Silver Ag^+	Copper (II) Cu^{2+}	Chromium Cr^{3+}	Lead (IV) Pb^{4+}
Ammonium NH_4^+	Zinc Zn^{2+}	Iron (III) Fe^{3+}	
Anions –1	**–2**	**–3**	
Fluoride F^-	Oxide O^{2-}	Nitride N^{3-}	
Chloride Cl^-	Sulfate SO_4^{2-}	Phosphate PO_4^{3-}	
Chlorate ClO_3^-	Carbonate CO_3^{2-}	Phosphide P^{3-}	
Cyanide CN^-	Phosphate HPO_4^{2-}		

Knowing the names of the anions and cations and their valencies, you can name ionic compounds and give their empirical formulas.

The ionic compound that contains the sodium cation and the chloride anion is called ‘sodium chloride’. The convention in naming the ionic compound is to name the cation first. The formula for the ionic compound is based on the rule that there must be equal numbers of positive and negative charges. In this way the whole ionic compound is neutrally charged. As the sodium ion Na^{+} has a +1 charge and the chloride ion Cl^{-} has a –1 charge, the formula is NaCl. This means that the compound has equal numbers of each ion.

It becomes a little more complicated if the cation does not have the same numbered charge as the anion. See if you can follow the combinations in Table 6.

TABLE 6: FORMULAS FOR COMPOUNDS

Example	Ionic compound	Valency of cation	Valency of anion	Empirical formula
1	Magnesium chloride	+2 (Mg^{2+})	+1 (Cl^{-})	$Mg(Cl)_2$ or $MgCl_2$
2	Silver oxide	+1 (Ag^{+})	–2 (O^{2-})	$(Ag)_2O$ or Ag_2O
3	Barium phosphate	+2 (Ba^{2+})	–3 (PO_4^{3-})	$(Ba)_3(HPO_4)_2$ or $Ba_3(HPO_4)_2$

In example 1 in Table 6, two Cl^{-} ions balance out the +2 charge on the Mg^{2+} ion. In the formula $MgCl_2$, the ‘2’ means there are two Cl^{-} ions for every Mg^{2+} ion.

In example 2, two Ag^{+} ions balance out the –2 charge on the O^{2-} ion.

In example 3, balance of charges only occurs when we have three Ba^{2+} ions for every two HPO_4^{2-} ions.

Table 7 gives the names of some common ionic compounds and their uses.

TABLE 7: COMMON IONIC COMPOUNDS

Common name	Scientific name	Formula	Used to manufacture
Baking soda	Sodium hydrogen carbonate	$NaHCO_3$	Bread products
Lime	Calcium oxide	CaO	Mortar
Potash	Potassium nitrate	KNO_3	Gunpowder
Gypsum	Calcium sulfate	$CaSO_4$	Plaster
Lye	Sodium hydroxide	NaOH	Soap

Metallic bonding

In your study of elements in the periodic table you will have noted that the elements are divided broadly into two groups: metals and non-metals. Metallic elements, of which there are more than eighty, have a huge range of atomic numbers and sizes. Not surprisingly, some of their properties (such as melting point, density etc.) vary widely. However, you will note that they are all grouped together on the left-hand side of the periodic table and that they all have low electronegativities.

You will be very familiar with the materials that we call metals. Sometimes they are pure metals, like copper, or 24-carat gold, which contain only one type of atom. More often they are mixtures of two or more metallic elements, such as steel or brass or solder; we call these mixtures alloys. Pure metals and alloys have many quite similar properties (such as good heat and electrical conductivity, and the ability to be formed or worked without losing strength) and we can predict that a similar type of bonding exists between their atoms.

Table 8 gives the constituents of some common alloys.

TABLE 8: CONSTITUENTS OF ALLOYS

Alloy	**Constituent elements**
Mild steel	Iron (Fe) with a few traces of carbon (C)
Stainless steel	73% iron (Fe), 18% chromium (Cr), 9% nickel (Ni)
Bronze	90% copper (Cu), 6% tin (Sn), 4% zinc (Zn)
Brass	65%–70% copper (Cu), 25%–35%, zinc (Zn)
Amalgam (dental filling)	50% mercury (Hg), 35% silver (Ag), 13% tin (Sn)
18-carat gold	75% gold (Au), 12.5% silver (Ag), 12.5% copper (Cu)
Sterling silver	92.5% silver (Ag), 7.5% copper (Cu)
Solder	70% lead (Pb), 30% tin (Sn)
Five-, ten-, twenty- and fifty-cent pieces	25% nickel (Ni), 75% copper (Cu)
$1 and $2 coins	2% nickel (Ni), 6% aluminium (Al), 92 % copper (Cu)

In this section we look at the properties of metals and from these infer the type of bonding that exists between the atoms.

Properties and structure of metals

This section describes how the properties of metals—lustre, conduction of heat, conduction of electricity, malleability, ductility, tensile strength, melting point, hardness and density—are related to the regular crystalline lattice structure of metals.

The key to this behaviour is their low electronegativity (refer to your periodic table and look at the electronegativities of the metallic elements compared to the others). This means that these atoms have a low electron-attracting power and so the small number of valence electrons that metal atoms have (usually one or two, occasionally three or four) is only weakly held to particular nuclei. In fact the outer shell electrons of metallic atoms move freely around to the outer shell of neighbouring atoms.

As the valence electrons do not attach themselves to any one particular atom, the atoms remain positive ions. The metal ions then arrange themselves in a three-dimensional lattice arrangement where the positive ions are neatly stacked and held together by the electrostatic forces between the ions and the free-roaming electrons.

The structure of a metal is a fixed three-dimensional lattice of positive ions embedded in a 'sea' of electrons. You can imagine them to be bricks and the free electrons as mortar that holds the bricks or ions in place. Metallic bonding is electrostatic, which means that there are attractions between the positively charged metallic ions and the negatively charged electrons. This type of bonding is called 'metallic bonding'.

Before proceeding you should be able to explain the properties of metal listed above in terms of the model of metals where there are positive ions in a 'sea' of mobile electrons. For example, the property of electrical conductivity means that metals easily carry electric currents. Currents are charges moving along a conductor. The fact that valence electrons, as negatively charged particles, are free to move in metals explains the property of electrical conductivity.

Properties and uses of metals

Metals are very widely used to make artefacts and as construction materials because of the useful properties they possess. Hardness, high strength and the ability to be cast, milled or beaten into a variety of shapes are essential requirements for building and construction materials. Because they are electrical conductors and are ductile, metals are the ideal material for making electrical wiring. Their colour and lustre make them popular materials for decorative purposes. Other properties make them useful for a great variety of purposes. But because different metals possess the so-called 'metallic properties' to differing extents, certain metals are preferred for certain tasks. The economics of production also plays an important role in deciding what metal is used for what purpose.

The bubble raft model, described in a later section, is a very useful model to try to understand the conflicting properties of strength, hardness and flexibility. It helps explain how either the existence of discontinuities in crystal structure or the introduction of foreign atoms affects these properties. Metallic bonding explains how the properties of a metal may be modified by the processes of work hardening, by the heat processes of annealing, quenching and tempering, and by the process of alloying with other (usually metal) atoms.

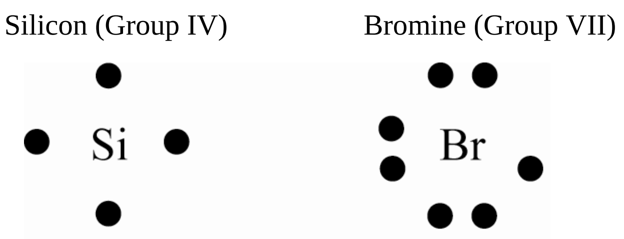
Covalent bonding

Covalent bonding characteristically occurs between atoms of non-metals. We have seen that non-metal elements have nearly completed valence shells, they have moderate to high electronegativity and are always found on the right-hand side of the periodic table. Atoms of these elements (unlike metal atoms) do not give up their electrons easily in their attempt to form a filled outer-electron shell, rather they achieve the stable octet state (full outer shell) by sharing outer-shell electrons with other non-metal atoms and forming molecules. A single covalent bond consists of a pair of electrons, usually one from each atom, situated in a region between the two bonding atoms. These negative electrons act as a 'glue' to stick the positively charged nuclei together. What results is a strong 'directional' covalent bond between the two atoms.

Covalent bonds

Electron-dot diagrams (sometimes called 'Lewis diagrams') are used a great deal by chemists to represent and try to explain the chemical behaviour of atoms. Only the outer-shell electrons are shown. See Figure 13.

FIGURE 13: LEWIS DIAGRAMS OF SOME ATOMS



In the Lewis diagram (see Figure 13), the number of valence electrons is determined by the Group number. Notice how the electrons are grouped around the atom. In the bromine atom the electrons form pairs, whereas in the silicon atom they do not. We must refine our model of the atom. It has been stated previously that electrons exist in shells or energy levels according to various rules. Now we must add to this idea so that electrons within shells occupy certain positions.

Atoms can contain no more than eight outer-shell electrons. These occupy four regions of space called 'orbitals' around the central atom. This means that as electrons fill shells they also fill regions within each shell called 'orbitals'. Orbitals can contain zero, one or two electrons; they cannot contain more than two electrons. The Lewis diagrams in Figure 13 give you some insight into how electrons fill orbitals in the outer shell. Before an orbital can contain two electrons (then it will be full) all other orbitals must have one electron in them (refer to the Lewis diagram for silicon). Therefore, if there are only four orbitals, they start to fill with five or more electrons. Notice in the bromine atom that three of the orbitals are full (three pairs) and one orbital has one electron. Electrons in full orbitals are non-bonding electrons or lone pairs and single electrons are bonding electrons. Only the bonding electrons are available to share with other atoms, thus forming a bond (called a 'covalent bond'). In Figure 13, silicon can form four covalent bonds, whereas bromide can only form one covalent bond.

Atoms with four or fewer outer-shell electrons have only one electron in each orbital and each electron can form a single covalent bond; for example, carbon (C), with four valence electrons can form four single covalent bonds.

Atoms with five electrons—for example, nitrogen (N)—have three orbitals, each containing one electron, and one filled orbital with two electrons. A nitrogen atom has three bonding electrons, so can form three single covalent bonds and it has one non-bonding pair.

Formation of molecules

When considering the formation of molecules, the number of single bonds that an atom can form is equal to the number of bonding electrons it contains. The formula and structure of the molecule that can be formed between different elements can be inferred from this.

The three different ways of representing molecules are:

- electron-dot (or Lewis) diagrams
- structural diagrams where the dots of an electron pair, of both bonding and non-bonding pairs, are replaced by a dash
- semi-structural diagrams where the dashes representing lone pairs are omitted and only the bonds shown.

The use of dots and crosses to represent valence electrons does not mean that there are different types of electron. All electrons are identical. The use of dots and crosses helps identify the source of electrons.

A double covalent bond may be formed when two elements each have two available bonding electrons. An example is oxygen; two oxygen atoms form an O_2 molecule. A double bond may be represented by two dashes.

A triple covalent bond may be formed when two elements each have three available bonding electrons. An example is nitrogen; two nitrogen atoms form an N_2 molecule. A triple bond may be represented by three dashes.

The structure and bonding of the important molecular compounds carbon dioxide (CO_2), water (H_2O), methane (CH_4) and ammonia (NH_3) are discussed in later sections.

The three-dimensional shape of molecules

The three-dimensional structure of molecules can be predicted using the valence-shell electron-pair repulsion (VSEPR) theory. The structure can be described by ball-and-stick or three-dimensional structural diagrams.

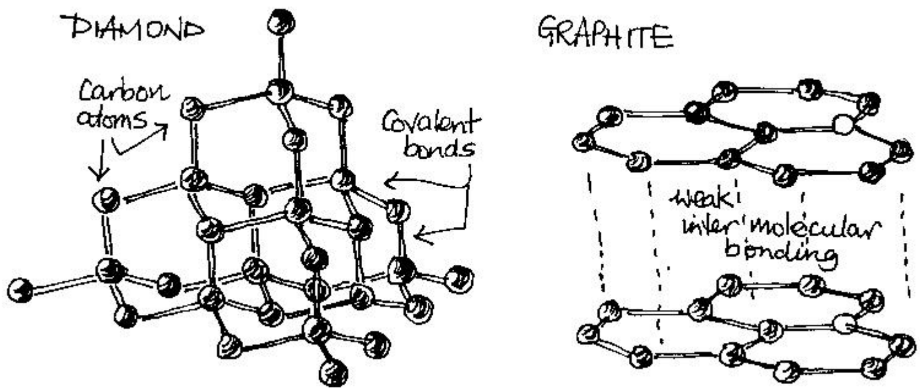
Shapes of some common molecules are described in Figure 14 and Table 9. You should particularly note the tetrahedral structure of methane. This is a common bonding structure of the organic molecules containing carbon atoms.

Covalent lattices

Sometimes covalent bonding extends in three dimensions with multiple atoms forming lattice arrangements instead of single molecules. These extended three-dimensional structures are sometimes called ‘giant molecules’. Perhaps the most famous is the diamond structure, which consists entirely of the element carbon. Each carbon atom, with four valence electrons, is bonded via a strong covalent bond to four other carbon atoms to form a giant tetrahedral lattice. Diamond, with its strong three-dimensional crystalline structure, is the hardest substance known. Another very hard covalent crystalline substance is quartz, consisting chemically of silicon dioxide (SiO_2) units. Many precious gems also have this type of structure.

A different allotrope (physical form) of carbon that also has an extended three-dimensional structure is graphite, the ‘lead’ of lead pencils (which is not lead at all). This soft, greasy, black substance has a layered structure. The carbon atoms within the layers are held together by strong covalent bonds with each carbon atom in the layer bonded to three other atoms in an extended hexagonal lattice. The layers are loosely held together by electrons which are free to move. Because of its structure and properties, graphite is used in pencils; it is used as a lubricant between metal surfaces like locks, and is fabricated into electrodes as it is a relatively good conductor of electricity and has a high melting temperature. The structural arrangements of diamond and graphite are shown in Figure 14.

FIGURE 14: LATTICE STRUCTURE OF DIAMOND AND GRAPHITE



Intermolecular bonding

In this section, the nature of the forces that hold molecules together (intermolecular forces) is investigated.

Molecular substances are either gases or liquids or low-melting-point solids and can be converted from one form to another quite easily by the application or removal of heat. Water, H_2O, turns to ice when it is refrigerated to below 0 °C; it changes to the gaseous state when it is heated above 100 °C. Nitrogen, N_2, the major component of air, is a gas at normal temperatures but will turn to a liquid if it is cooled below –196 °C and to a solid at even lower temperatures. Oxygen, O_2, behaves similarly. Many organic substances are molecular and can exist in the three states, and the changes between them, which occur quite sharply at a particular temperature, are familiar. One example is the evaporation of alcohol,

C_2H_6, at 78 °C when methylated spirits is heated. Another is candle wax (a mixture of higher hydrocarbons) which first melts when heated, then vaporises before burning to form the candle flame.

When a molecular compound goes through a change the molecules stay intact. From this you should see that the bonds between the molecules are weaker than those between the atoms in the molecules. Chemists talk about strong intramolecular bonding but weak intermolecular bonding.

Strength of intermolecular bonding

Melting points and boiling points are measures of the strength of the intermolecular bonding. As we know, different molecules have different melting and boiling points; we must look for reasons for this.

This is partly because of the size of the atoms: how many protons and electrons they have. As a rule, the larger the atoms or molecules the stronger the bonding and the higher the melting and boiling points. This is illustrated in the table of melting and boiling points and molecular masses of some common diatomic molecules (Table 9).

TABLE 9: MOLECULAR MASS AND MELTING AND BOILING POINTS

Diatomic molecule	**H_2**	**N_2**	**O_2**	**F_2**	**Cl_2**	**Br_2**	**I_2**
Molecular mass	2	28	32	38	71	160	254
Melting point (°C)	–259.2	–210	–218.8	–219.6	–101	–7.2	113.7
Boiling point (°C)	–252.7	–195.8	–183	–186.2	–34.7	58	183

The other significant factor that relates to the strength of intermolecular bonding occurs when the atoms making up the molecule are different, especially if they have a large difference in electronegativity. Table 9 gives the electronegativities (bonding electron attracting power) of some elements.

TABLE 10: ELECTRO-NEGATIVITIES OF ATOMS

Element	**Electronegativity**
H	2.1
C	2.5
N	3.0
O	3.5
F	4.0
P	2.1
S	2.5
Cl	3.0
Br	2.8

Take the example of the molecule hydrogen chloride (HCl). This molecule consists of two different atoms: hydrogen and chlorine. From Table 10, note that the electronegativity of Cl is greater than that of H. The electronegativity is a measure of the relative pull by the atom on the electrons in the bonding pair. This means that, although H and Cl share electrons, Cl will take more of the share and the bonding electrons will be closer to Cl than to H. As electrons are negatively charged, being closer to Cl means that the Cl end of the molecule will have a slightly negative charge, whereas the H end of the molecule will have a slightly positive charge.

Dipole–dipole bonding

The asymmetric distribution of charge on the molecule is called a 'dipole'. Molecules that form diploes are called 'polar molecules'. Can you see now how molecules can be attracted to each other? Once again, the bonds that are formed are electrostatic. The positive end of one polar molecule attracts the negative end of another. The greater the difference in electronegativities between the atoms in the molecule, the greater the polarity and the greater the strength of the intermolecular bond. This type of intermolecular bonding is called 'dipole–dipole bonding'.

In bulk amounts of molecules like HCl the dipoles tend to align themselves positive to negative and the dipole–dipole bonds formed result in stronger intermolecular bonding and higher melting and boiling points than expected. All polar molecules form dipole–dipole bonds and tend to align themselves into a regular crystalline order in the solid phase.

Hydrogen bonding

The strongest intermolecular bonding occurs between molecules containing hydrogen atoms and one of the three high-electronegativity small atoms, nitrogen, oxygen and fluorine. The melting points of ammonia (NH_3), water (H_2O) and hydrogen fluoride (HF) are much higher than expected.

The so-called 'hydrogen bonds' require closer study because they are crucial to the understanding of much of the behaviour of substances in contact with water, including a lot of biochemistry like protein behaviour and DNA replication.

Hydrogen has a lower electronegativity than most other atoms. This means that in covalent molecules the bonding electrons stay closer to the other atom. As hydrogen only has one electron and one proton, then in the covalent compound; the protons in the hydrogen atoms are exposed on one side. Therefore, the molecules that contain hydrogen, such as water, are very polar. The hydrogen atoms in the molecule are positively charged and the non-hydrogen atoms in the molecule are negatively charged. The intermolecular bonding occurs between the positive hydrogen atoms in the molecule and the negative non-hydrogen atoms of another molecule. This type of bonding is called 'hydrogen bonding'.

In ice, this intermolecular hydrogen bonding is strong enough to hold the molecules in a rigid tetrahedral three-dimensional structure. This relatively open structure holds the water molecules further apart than they are in liquid water.

Ice is the only chemical substance whose solid form is less dense than its liquid form; hence, ice floats on water. Life in polar regions depends on this unique property.

Of the many other manifestations of hydrogen bonding that are crucial to life, perhaps the most famous is the replication of the genetic material DNA, which occurs in the nucleus of cells when they divide and the translation of the DNA code in the formation of proteins.

Plastics and fibres

The word ‘plastic’ has been in our vocabulary for a lot longer than the substances we know as plastic have been in existence. The word plastic simply means ‘able to be moulded’. The substances that we call ‘plastic’ have been in existence for less than 100 years. They have developed in parallel with the petroleum industry, for they are largely made from petrochemicals, refined and fabricated from crude oil or other fossil fuels. Major components of most plastics are the elements carbon and hydrogen.

Chemists and manufacturers use the technical name ‘polymer’ (from the Greek, *poly*, many, and *mer*, unit) when referring to plastics. As we will see later, this is in recognition of the fact that plastics are very large molecules (macromolecules) made up of many small units. In other words, a polymer is made up of many monomer units. These units may be linked or bonded in the form of a long chain (thermoplastic), or they may be cross-linked to each other to form a rigid lattice (thermosetting).

Fibres are long-chain polymers that have been drawn out to very thin threads. All thermoplastics could be made into fibres but certain ones (often containing polar atoms or groups) have much higher tensile strength because the bonds between their individual molecules are stronger and they ‘stick together’ better in the narrow cross-section profile of a fibre.

Properties of plastics

A very wide range of substances goes under the name of plastic. Some of them aren’t strictly plastic at all (in the sense that once formed they can’t be melted or reformed). These are called ‘thermosetting polymers’ (or ‘cross-linked polymers’). Most, however, are solid at room temperature but melt as the temperature is increased. These are called ‘thermoplastics’. These ‘true’ plastics can be heat-moulded to form objects that, when they cease being of use, can be melted down and recycled. Thermoplastics usually don’t have a sharp melting point but soften gradually before becoming liquid. This is because they are made up of polymer molecules of different chain lengths with different numbers of monomer units.

Plastics usually have a low density and a high strength-to-weight ratio and are hard wearing. They are chemically inert. They are, however, invariably flammable. Because of their organic nature they are often soluble in organic solvents, although there are water-soluble polymers. Plastics are mostly easy to process. They are often clear materials but take up fillers and colours well. Polymer technology has advanced to the state where particular polymer molecules can now be chemically tailored to produce plastics that have specific desired properties.

So, not surprisingly, plastics are now ubiquitous. They are increasingly replacing more traditional materials such as wood, stone, metal, glass, ceramics, leather and natural fibres.

Uses of plastics

As a case study we have taken just one area—modern sports—to look at the impact that the development of modern plastic technology has had. Much of the improvement in athletic excellence has been made possible by plastics. Table 11 gives a summary of some of the names of plastics used in modern sports together with their properties.

Athletic excellence made possible by plastics <http://www.americanplasticscouncil.org/athletic/home.html> is a site with many links that provide information about the use of plastics in sport. Table 11 gives a summary of some of the names of plastics used in modern sports, together with their properties.

TABLE 11: PLASTICS IN SPORT

Plastic	Properties	Sport
ABS foam (acrylonitrile-butadiene-styrene)	Impact resistance, impact absorption	Baseball, softball, athletics, shooting
Carbon fibres/graphite (carbon-reinforced plastic)	Strength, durability, abrasion resistance, lightweight structure, impact resistance, shatter resistance, stiffness	Archery, canoeing, kayaking, rowing, cycling, tennis
EPP foam (expanded polypropylene)	Impact absorption	Canoeing, kayaking, rowing
EPS/EPU/EPE foams (expanded polystyrene) (expanded polyurethane) (expanded polyethylene)	Impact absorption	Cycling, field hockey, equestrian
EVA foam (ethylene-vinyl-acetate)	Impact absorption	Athletics, baseball, softball, beach or indoor volleyball, boxing, field hockey, weight-lifting

Fibreglass (glass reinforced with plastic)	Lightweight structure, durability, flexibility, impact resistance, abrasion resistance, shatter resistance	Archery, athletics, canoeing, kayaking, field hockey, gymnastics, sailing
Kevlar	Strength, impact resistance, abrasion resistance, stiffness	Archery, beach or indoor volleyball, canoeing, kayaking, rowing, sailing, fencing, tennis
Spandex/Dorlastan/ Lycra	Flexibility, moisture absorption, aerodynamics, comfort	Wrestling, cycling, swimming, synchronized swimming, diving, water polo
Neoprene	Slip resistance, impact absorption	Gymnastics, weight-lifting
Nylon	Durability, lightweight structure, elasticity, strength, water resistance	Beach/indoor volleyball, cycling, sailing, fencing, equestrian, tennis, water polo, weight-lifting
Polycarbonate	Impact resistance, shatter resistance, abrasion resistance, optical clarity, durability	Cycling, field hockey, gymnastics, shooting, swimming, synchronized swimming, diving, tennis
Polyester	Durability, elasticity, strength, water resistance	Beach or indoor volleyball, sailing, tennis
Polyethylene	Impact absorption, durability, rebound ability	Baseball, softball, gymnastics
Polypropylene	Impact absorption, impact resistance, strength, elasticity, water resistance	baseball, softball, field hockey, sailing
Polyurethane	Durability, strength, rebound ability, impact absorption, shape retention, weather resistance	Archery, athletics, beach or indoor volleyball, boxing, field hockey, football, gymnastics, tennis, weight-lifting
Polyurethane/polymer Gel coats	Durability, weather resistance	Canoeing, kayaking, rowing, equestrian, diving, sailing

PVC (poly-vinyl-chloride)	Durability, shape retention, weather resistance, comfort, watertightness	Athletics, beach or indoor volleyball, boxing, wrestling, field hockey, football, gymnastics, swimming, synchronized swimming, diving
Spectra, Vectran	Elasticity, strength, flexibility, water resistance, durability	Sailing
Vinyl nitrate foam	Impact absorption with little rebound ability	Wrestling

Polymer structure and the chemistry of polymerisation

It is beyond the scope of this topic to go into detail about the processes of polymerisation and the chemical nature of the different polymers and plastics.

ACTIVITY: TESTING FIBRE AND FABRICS

You will need:

- **a heat source (e.g. candles)**
- **tongs**
- **a microscope**
- **a magnifying glass**
- **a heat-proof mat or large tray**
- **water**
- **a measuring cylinder**
- **an electric iron**
- **a wooden block and sandpaper**
- **a large collection of different fibres and fabrics.**

Collect natural and synthetic fibres and fabrics.

Classification of features: Make observations of your set of fabrics using all of your senses by looking, pulling, stretching, feeling, listening etc. Use a magnifying glass, and if possible a microscope, to extend your observations. How are the materials similar and how are they different?

Flammability: You need to use the simplest and safest method to undertake flammability tests. You should use a heat source that releases just enough energy to complete the task. Household candles are usually sufficient. Use tongs to hold the material and the heat-proof mat to hold the candle and burning materials. Alternatively, use large trays.

Two types of investigation include:

- **moving the flame close to the fabric or thread but not applying it directly to it**
- **direct application of the flame to the material to be tested.**

Does the flame change colour? Does the material burn? How fast? Is there any melting? Is there any ash left after burning? What colour is it?) Do the fumes and smoke smell? How quickly does the material ignite?

Speed of soaking: Cut strips of fabric about 10 cm × 1 cm and place one end of each strip to an equal depth (approximately 1 cm) in water. Measure the height the water has travelled along the strips over the remainder of the practical session. The water travels vertically because of capillary action.

Absorption: Predict which fabric or fibre will absorb most water and grade them. Cut out fabric squares and carefully insert them in a measuring cylinder full of water. Let them soak and lift them clear of the water. Record volume changes. What procedure will you adopt to ensure a fair test? Compare your results with your predictions and attempt an explanation of the comparison.

Drying times of fabrics and fibres: In recent times the sought after qualities of 'ease of drying' and 'crease resistance' have greatly affected fabric composition and construction.

What factors would increase efficiency of drying? Conduct an investigation into either or both of these characteristics. Where significant changes occur—for example, creasing of the material—tease out a fibre and examine it using a magnifying glass. Use an electric iron to press the creased piece of fabric and then once again examine a fibre comparing it with: (i) the original fibre; (ii) fibre from the creased piece; and (iii) fibre from the pressed piece of fabric.

Wearing a fabric: Use the fabric samples, the wooden block and some sandpaper, or the rim of a twenty-cent piece; devise a fair test to determine which fabrics wear best.

Design a fabric: Use your knowledge gained from the tests to describe the structure and material that would be used in a fabric that is light, water-resistant and stretchable.

ACTIVITY: NEEDLE THROUGH A BALLOON

This activity explores polymers.

You will need:

- **balloons**
- **an upholstery needle (30 to 35 cm), a sharpened knitting needle, a bamboo skewer or a wire coathanger sharpened to a point**
- **a small amount of cooking oil**
- **a paper towel or cloth.**

Inflate the balloon and tie it off. Let a little air out of the balloon before tying it off, so it will be easier to puncture the balloon without breaking it. Make sure the balloon is not longer than the needle.

Dip the tip of the needle (or bamboo skewer etc.) into the cooking oil. Alternatively, use a paper towel or cloth to spread the oil along the length of the needle.

Using a gentle twisting motion, insert the needle into the nipple end of the balloon (the end opposite the knot) where the balloon is thicker.

Pull the needle out slowly through the tied end. The balloon will slowly deflate.

After the needle is out, jab the balloon sharply with the needle; it will pop.

Explanatory note: The balloon is made of long molecular chains called 'polymer chains', which stretch and seal around the needle that is slowly inserted. When the needle jabs the balloon, the molecules do not have time to stretch around the needle and so the air rushing out of the balloon breaks the bonds of the long chains.

A similar phenomenon occurs if you use a zip-lock plastic bag filled with water and slowly push a sharp pencil in one side through to the other side. Rubber tyres on cars also work this way. A gummy layer on the inside of the tyre seals around any nails or sharp objects that poke directly into the tyre.

ACTIVITY: INVESTIGATING PLASTICS

You will need:

- **a collection of as many different types of plastic as you can find (e.g. toys, plastic food wrap, transparencies, shopping bags, milk bottles, drink bottles, polystyrene cups, freezer bags, plastic fabrics such as nylon)**
- **plastic household items**
- **an onion**
- **chocolate**
- **sherry or vinegar**
- **wearable plastic items**
- **polystyrene, china, thin plastic, metal and paper cups**
- **a thermometer**
- **a variety of plastic bags**
- **plastic bottles**
- **a strip of plastic**
- **nylon line**
- **plastic and glass containers.**

Comparing plastics: Compare different types of plastic in terms of their density, flexibility, ability to float or sink, reaction to hot water, scratchibility and reaction to tearing and cutting (ease, type of edge). Use tools where appropriate to extend your investigations.

Household items: Collect a number of household items that have plastic or have components that are plastic. Compare the types of plastic used in the household objects on display. Identify the plastics used. What properties of the plastic make it an appropriate material for the item in each case? What would the main competitor be in each case (wood, metal etc.)?

Plastic to wear: Make a list of the plastic objects you commonly wear. Why do we wear plastics? Choose a room in your home. What plastic items can you find there?

Keep it warm!: Which do you think will keep a cup of coffee warm the longest: a cup made of polystyrene, china, thin plastic, metal or paper? Predict the order of effectiveness. Use hot water from a tap and a thermometer to compare the cooling rate over time. What precautions do you have to take to ensure the test is fair?

Strength of plastics: Cut 1-cm-wide strips of different plastics from bags and compare their strength by hanging weights from them. Which sort of shopping bags hold the most weight? Check how many bricks they can hold. Check for any special design features they have—for example, reinforcement, handle design. Devise a method to test value for money in garbage bags. Test the strength of plastic bottles by filling them with water and dropping them from various heights. Does the area of first contact between the bottle and the ground have any effect?

Plastic stretches over time: Hang weights from a strip of plastic and a piece of nylon line, and record the length over a time interval. Does the plastic go back to the original length if the weight is taken off?

Do plastic containers contaminate the taste of food?: Taste chocolates left in plastic and glass containers that have previously contained onion, sherry or vinegar.

Explanatory note: In explaining your observations try to think about the type of bonding contained within the plastic. To do this you will need to refer to the notes in relation to the different types of plastics and the various bonding arrangements that account for their properties.

ACTIVITY: RUBBER BANDS AND HEAT

This activity was taken from the website 'Science is fun' <http://scifun.chem.wisc.edu/HOMEEXPTS/rubberband.html> (viewed 9 March 2005).

You will need:

- **rubber bands**
- **a hair dryer.**

Almost everyone has used rubber bands, but few people have taken the time to observe the less obvious properties of these everyday objects. In this activity, you will examine the thermal properties of rubber—that is, the behaviour of rubber as it relates to heat energy.

In the first experiment you will attempt to detect heat flow into or out of a rubber band. To do this, you need a rather sensitive heat detector.

Fortunately, you have such a detector with you at all times. Surely you've felt the heat of a flame or the cold of an ice cube. Therefore, you know that your skin is sensitive to heat flow. In this experiment, you will detect heat flow using some of your most sensitive skin, that on your forehead or on your lips.

1) **Place your thumbs through a heavy rubber band, one on each end. Without stretching the band, hold it to your forehead or lip. Does the band feel cool or warm or about the same as your skin? Repeat the test several times until you are sure of the result.**
2) **Move the rubber band slightly away from your face, so it is not touching your skin. Quickly stretch the band about as far as you can and, holding it in the stretched position, touch it again to your forehead or lip. Does it feel warmer or cooler or about the same as it did when it was relaxed?**
3) **Move the stretched rubber band away from you face. Quickly let it relax to its original size and again hold it to your skin. Does it feel warm or cool?**
4) **Repeat the stretching and testing, and relaxing and testing several times until you are sure of the results. An object feels cool or cold to you when heat flows from your skin to the object. Conversely, an object feels warm or hot when heat flows from the object into your skin. If the stretched rubber band feels**

cool, then it absorbs heat energy from your skin. If it feels warm, then it gives off heat energy to your skin. If the band feels neither warm nor cool, there is no detectable heat flow. These three cases can be represented as follows:

- **case 1: relaxed band + heat-stretched band**
- **case 2: relaxed and stretched band + heat**
- **case 3: relaxed band stretched band (no heat)**

Which of these three cases best describes what you observed?

There is another way to test which of the three statements is correct. We can see what happens to the length of a rubber band if we heat or cool it.

1) **Hang one end of the rubber band from the wall or ceiling and suspend a weight from the other end of the rubber band. (What you use for a weight will depend on what is available. The weight should be heavy enough to stretch the rubber band, but not so heavy that it is likely to break it—e.g. hang the band over a door knob and suspend a hammer from the band.)**
2) **Heat the rubber band with a hair dryer. Start the dryer and, when it has warmed up, turn its heat on the stretched rubber band. Does the stretched rubber band become longer or shorter when it is heated?**

Does this observation agree with what you found in the first part of the experiment? Doing an experiment several ways and checking for agreement in the results is an important strategy in science.

When rubber is heated it behaves differently from most familiar materials. Most materials expand when they are heated. Consider the liquid in a thermometer. The thermometer works because the liquid expands when its temperature increases. Similarly, a wire made of metal, such as copper, becomes longer as it gets hotter. The expansion of metals with increasing temperature is the principle behind the functioning of home thermostats and jumping discs.

In materials made up of small, compact molecules—for example, the liquid in a thermometer—as the molecules move about more, they push their neighbouring molecules away. Rubber, on the other hand, contains very large, threadlike molecules. When rubber is heated, the sections of the molecules move about more vigorously. In order for one part of the molecule to move more vigorously as it is heated, it must pull its neighbouring parts closer. To visualise this, think of a molecule of the stretched rubber band as a piece of string laid out straight on a table. Heating the stretched rubber band causes segments of the molecules to move more vigorously, which can be represented by wiggling the middle of the string back and forth. As the middle of the string moves, the ends of the string get closer together. In a similar fashion, the molecules of rubber become shorter as the rubber is heated, causing the stretched rubber band to contract.

ACTIVITY:
A STAMPEDE!

This activity is taken from taken from Chemshorts for Kids <http://membership.acs.org/C/Chicago/ChmShort/cs94.html> (viewed 9 March 2005).

Printing presses use rubber rollers to pick up ink and apply it to the surface of paper. Because of the chemicals that make up rubber, it has a way of picking up ink and then releasing the ink to paper. Rubber works because of the unique interactions between molecules in rubber and the molecules in ink and paper. So

why can an eraser rub out pencil marks from paper? Because the attraction between rubber and graphite (pencil lead) is stronger than that between paper and graphite.

You will need:

- a 5 cm square of cardboard
- a rectangular pink school eraser
- a ballpoint pen
- ink
- white paper.

Now let’s make our own rubber stamps. Cut out a 5 cm square of cardboard. Draw a line down the middle of a rectangular pink school eraser. Now draw a simple shape (star or heart) on both the cardboard and on one-half of the eraser using a ballpoint pen. Fill in the shapes with a lot of ink to make them dark. Press the shapes down hard on a piece of white paper. Which one prints better? On the other half of the eraser draw a simple picture such as a flower or tree using red, blue and green ballpoint pens. See how many good clear prints you can make on your paper without re-inking the eraser.

Use another eraser to make a stamp of your name. To have your name come out correctly when you print it, you have to write it backwards on the eraser. Add some designs around your name using different colours; be creative and have fun!

ACTIVITY: BUBBLEGUM CHEMISTRY

This activity is taken from Marsella (1994).

All you really have to do in this ‘experiment’ is chew your favourite kind of bubblegum for a while. Think about what you learn here while you are chewing.

You will need:

- bubblegum.

Bubblegum is a mixture of several chemicals, but rubber is the most important. A good bubblegum must be strong enough to stretch to a thin film without breaking, but still be soft enough to chew easily. That’s a tall order. The other chemicals in bubblegum—resins, waxes, fillers, flavours, sugar, humectants and emulsifiers—are all there either to provide flavour or to modify when and by how much the rubber stretches. Rubber molecules are polymers, which are long chain-like molecules formed when many smaller molecules bond together end to end. A natural polymer called ‘latex’, which is from trees, is used to provide the stretchy part of bubblegum. Many bubblegum companies use a synthetic, food-grade version of the same rubber that goes into truck tires! This polymer is a mixture of styrene and butadiene and is abbreviated to ‘SBR’.

Of the twenty or so chemicals in bubblegum, some dissolve in water and some do not. Most of the water-insoluble portion of bubblegum is called ‘gum base’. That’s where the rubber is. Some of the additives in the bubblegum purposely actually restrict the size to which the bubbles can be blown, so as not to completely alienate parents! The most intense fragrances and flavourings in fruits are often essential oils like limonene (which is from orange and lemon rinds). They are well suited to bubblegum because they are not water soluble and do not dissolve out of the bubblegum in your mouth. Bubblegum does seem to lose flavour after a while, but that is usually because the sugar, which intensifies the fruit flavour, has dissolved.

Chemists must think not only of how the bubblegum tastes and how big the bubbles get, but also how it feels in the mouth. It must soften without getting gooey, take up water without dissolving, and keep its flavour for as long as possible. On top of all that, it must not dry out on store shelves, it should not stick to the wrapper, and it should be easy to work with in the factory. Chemists know how to tweak all the ingredients to make a formulation that is just right; who knew a simple thing like bubblegum could be so complicated!

Recycling plastics

Most plastics are produced from crude oil, coal and gas and so their production contributes to the depletion of limited resources. Disused plastics produce problems in terms of waste disposal as most are not biodegradable. Plastics have been known to be hazardous to marine life. For these reasons recycling of plastics is important.

Not all plastics can be recycled. In addition, not all recycled plastics are made of the same type of plastic and so need to be sorted according to their type. The plastics industry has introduced a voluntary coding system that identifies the resin composition of plastic containers. The system makes it easier for recyclers to identify and separate plastics for recycling. Once sorted, the plastic groups are shredded into flakes so that they can be washed, melted and remoulded.

Table 12 gives details of this code.

TABLE 12: COOLING SYSTEM TO IDENTIFY PLASTICS

Code Symbol	Type of Plastic	Characteristics	Applications
1 PET	Polyethylene Terephthalate PET	Clear, tough, solvent resistant, often used as a fibre.	Soft drink bottles, pillow and sleeping bag filling, textile fibres.
2 HDPE	High Density Polyethylene HDPE	Hard to semi-flexible, waxy surface, opaque, melts at 135°C	Crinkly shopping bags, freezer bags, milk bottles, buckets, rigid agricultural pipes.
3 V	Unplasticised Ployvinyl Chloride UPVC	Hard, rigid, can be clear, can be solvent welded	Electrical conduit, plumbing pipes and fittings, blister packs, clear cordial and fruit juice containers.
	Plasticised Ployvinyl Chloride PPVC	Flexible, clear, elastic, can be solvent welded	Garden hose, show soles, cable sheathing, blood bags and tubing, watch straps.
4 LDPE	Low Density Polyethylene LDPE	Soft, flexible, waxy surface translucent, withstands solvents	Garbage bags, squeeze bottles, black irrigation tube, garbage bins.
5 PP	Polypropylene PP	Hard but still flexible, waxy surface, melts at 145°C, translucent, withstands solvents. Very versatile with many applications.	Potato crisp bags, drinking straws, microwave ware, plastic kettles, plastic garden settings, baby baths, plastic hinged boxes.
6 PS	Polystyrene PS	Clear, glassy, rigid, opaque semi-tough, melts at 95°C. Affected by fats and solvents.	Plastic cutlery, imitation 'crystal glassware', low cost brittle toys.
	Expanded Polystyrene EPS	Foamed, light weight, energy absorbing, heat insulating.	Panel insulation, produces boxes, protective packaging for fragile items.
7 OTHER	OTHER: Includes all other resins and multi materials (eg laminates). Examples are polyamide, acrylonitrile butadiene styrene (ABS), acrylic, nylon, polyurethane (PU) and phenolics.		

(http://www.plastic bottles.com.au/recycle.htm, viewed 10 March 2005)

References

Bennet, D & Gardner, E 1994, *Doing science*, Horwitz Publications, Sydney.

Carrado, KA 1993, *ChemShorts for Kids*, Argonne National Labs, viewed 9 March 2005, <http://membership.acs.org/C/Chicago/ChmShort/cs93.html#1.93>.

Gammon, SD 1994, *Journal of Chemical Education*, vol. 71, p. 1077.

Heffernan, DA & Learmonth, MS 1997, *The world of science: book one*, 3rd edn, Addison Wesley, Melbourne.

James, M, Derbogosian, M, Raphael, S & Moloney, J 1999, *Chemical connections: 1*, John Wiley & Sons, Milton, Qld.

Marsella, G 1994, ‘Bubblegum’, *ChemMatters*, vol. 12, no. 3.

Martin, S & Connor, A 1970, *Basic physics 1*, Whitcombe & Tombs, Sydney.

Parratore, P 1994, *Wacky science: a cookbook for elementary teachers*, Kendall/Hubt, Dubuque, P. 112.

Van Cleave, J 1989, *Chemistry for every kid*, John Wiley & Sons, N.Y.

Further resources

Agler, L 1987, *Liquid explorations*, University of California, Lawrence Hall of Science, Berkeley.

Ahouse, JJ 1987, *Fingerprinting*, University of California, Lawrence Hall of Science, Berkeley.

Barber, J 1989, *Cabbages and chemistry*, University of California, Lawrence Hall of Science, Berkeley.

Barber, J 1989, *Crime lab chemistry*, University of California, Lawrence Hall of Science, Berkeley.

Barber, J 1988, *Vitamin C testing*, University of California, Lawrence Hall of Science, Berkeley.

Bird, J & Catherall, E 1976, *Fibres and fabrics*, Macdonald Educational, London.

Elvins, C, Jones, D, Lukins, N, Ross, B & Sanders, R 1999, *Chemistry one*, 3rd edn, Heinemann, Melbourne, Victoria.

Kinnear, J & Martin, M 1992, *Nature of biology: Book one*, The Jacaranda Press, Milton, Qld.

Smith, A & Dwyer, C 1991, *Key chemistry*, Melbourne University Press.