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# Gases, Liquids and Solids



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# Gases, Liquids and Solids

**Philip Matthews**

**Series editor: Brian Ratcliff**



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# Introduction

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### The presentation of units

You will find that the books in this series use a bracketed convention in the presentation of units within tables and on graph axes. For example, ionisation energies of  $1000 \text{ kJ mol}^{-1}$  and  $2000 \text{ kJ mol}^{-1}$  will be represented in this way:

Measurement	Ionisation energy ( $\text{kJ mol}^{-1}$ )
1	1000
2	2000

OCR examination papers use the solidus as a convention, thus:

Measurement	Ionisation energy / $\text{kJ mol}^{-1}$
1	1000
2	2000

Any numbers appearing in brackets with the units, for example ( $10^{-5} \text{ mol dm}^{-3} \text{ s}^{-1}$ ), should be treated in exactly the same way as when preceded by the solidus,  $/10^{-5} \text{ mol dm}^{-3} \text{ s}^{-1}$ .

## Gases, Liquids and Solids – an A2 option text

*Gases, Liquids and Solids* contains everything needed to cover the A2 option of the same name. It is a brand new text which has been written specifically with the new OCR specification in mind. A specialised glossary of terms is included, linked to the main text via the index.

The book is divided into four chapters corresponding to the module sections States of Matter, Phase Diagrams, Distribution Between Phases, and Raoult's Law and Distillation.

The module builds upon material in *Chemistry 1*, in particular aspects of Structure and Bonding from the Foundation Chemistry module. Please note that some prior knowledge required to cover this module is not covered until the synoptic unit of the A2 course. However, supporting material can be found in *Chemistry 2*.



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# States of matter

By the end of this chapter you should be able to:

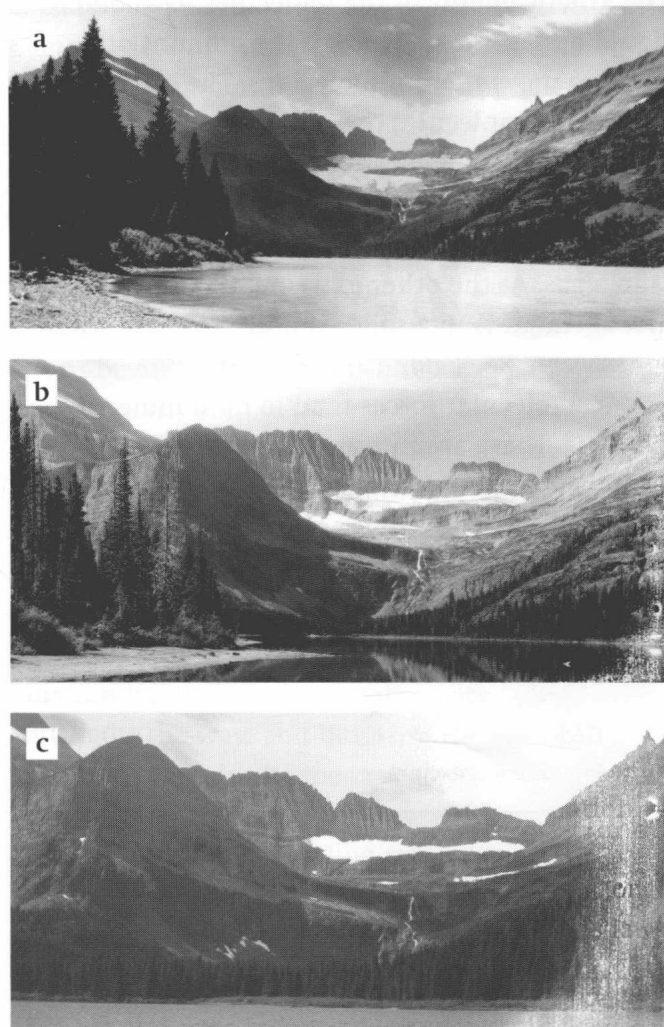
- 1 describe, using a *kinetic-molecular model*, the solid, liquid and gaseous states, melting, vaporisation and vapour pressure;
- 2 state the basic assumptions of the *kinetic theory* as applied to an ideal gas;
- 3 explain qualitatively, in terms of *intermolecular forces* and *molecular size*, (i) the conditions necessary for a gas to approach *ideal behaviour*, and (ii) the *limitations of ideality* at very high pressures and very low temperatures;
- 4 state and use the *ideal gas equation*  $PV = nRT$  in calculations, including the determination of the relative molecular mass of a volatile liquid.

## The three states of matter

Figure 1.1 shows the Grinnell glacier, at various times in the past. You can see just how much of the glacier has disappeared between 1914 and 1997.

Many scientists believe that the increasing rate of melting of glacier ice, and ice at the polar ice caps, is a result of global warming. You may know that, if global warming continues, the sea level on Earth will gradually rise. Mainly the rise will be due to the expansion of sea water; but the melt water from glaciers and the polar ice caps will add to the increase as well. In fact, it has been estimated that, if all the ice over the Earth's land mass were to melt, sea level would rise by around 20 m. For many Europeans, the retreat of glaciers in mountainous regions and the end of reliable winter snow falls will be the end of skiing holidays. However, there are far more serious consequences world wide; for example, glaciers provide billions of gallons of water for drinking water and the irrigation of crops in lowland regions. In Africa, the end of glaciation will make vast tracts of land uninhabitable owing to the lack of melt water. It has been estimated that a rise of  $4^{\circ}\text{C}$  in average air temperature will be enough to melt all the glaciers in Europe, and even the Himalayas.

Studying the conditions that cause ice to form, and melt, and the change of water vapour into



• **Figure 1.1** Three photographs of the Grinnell glacier, Glacier National Park, USA, taken in a 1914, b 1938 and c 1997.

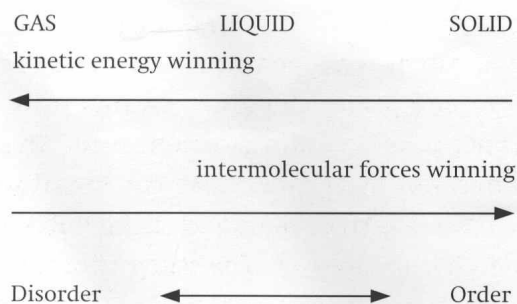
rain and snow, is part of chemistry. We know how and why water changes between its solid, liquid and gaseous forms; and this knowledge is important in predicting the results of global warming. However, the same processes are at work when any solid, liquid or gas changes from one form to another. This book will introduce you to the main factors that we believe are responsible for the different properties of gases, liquids and solids. However, as in all good stories, we should start at the beginning; and that means you need to know what we mean by states of matter.

The three **states of matter** are solid, liquid and gas. Whether a substance exists as a solid, liquid or gas mainly depends on two things:

- 1 *Kinetic energy* – which increases as a substance is heated.
- 2 *Intermolecular forces* – the forces between the molecules that make up the substance.

The kinetic energy of the molecules in a solid, liquid or gas is a measure of the amount of random movement of molecules. The more kinetic energy the molecules of a substance have, the greater is the tendency for its molecules to be jumbled up, i.e. to be more disordered. The most *disorderly* arrangement that molecules can achieve is in a gas. At the other extreme, the most *orderly* arrangement is in a solid. Liquids are somewhere in between. See *figure 1.2*.

Intermolecular forces tend to hold molecules together. There are intermolecular forces between all molecules; but between some they are very weak, and between others they are quite strong. When the forces are weak, the molecules are not likely to cling together to make a liquid or solid unless they have very little kinetic energy. The



● **Figure 1.2** The relation between kinetic energy, disorder and intermolecular forces.

noble gases are excellent examples of this. For instance, helium will not liquefy until the temperature is almost as low as  $-269^{\circ}\text{C}$ , or 4 K. On the other hand, the intermolecular forces between water molecules are very strong – strong enough to hold them together up to  $100^{\circ}\text{C}$ .

To summarise, we can say that:

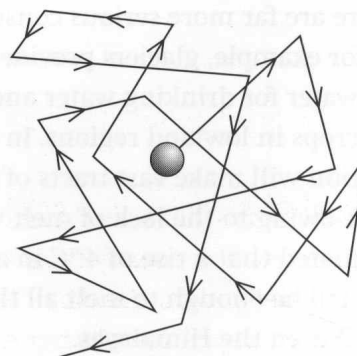
Intermolecular forces tend to bring order to the movements of molecules.

Kinetic energy brings disorder, and leads in the direction of randomness or chaos.

Thus, at a given temperature, a substance will exist as a solid, liquid or gas depending on where the balance between these two opposing influences lies. We saw in *Chemistry 1*, chapter 15, and shall see on page 8 in this book, that an *equilibrium* can be set up when a substance changes state.

## How do we know that gases are disorderly?

One piece of evidence for this comes indirectly from the experiments first performed by Robert Brown in 1827. He observed the movement of pollen on the surface of water, which he found to be completely unpredictable. The random movements of the pollen, known as **Brownian motion**, were finally given a mathematical explanation by Albert Einstein (of relativity fame) in 1905. He showed that a grain of pollen went on a random walk. A **random walk** (*figure 1.3*) is the sort of walk that a very drunk person would go on if put out in an open space. If we assume that the drunk found it impossible to make a conscious choice, he (or she) would be as likely to walk in one direction as any other. The reason why the grains



● **Figure 1.3** A random walk of a pollen grain.



behave in this way is that they are being bombarded by molecules in the liquid, which are themselves moving in a random way.

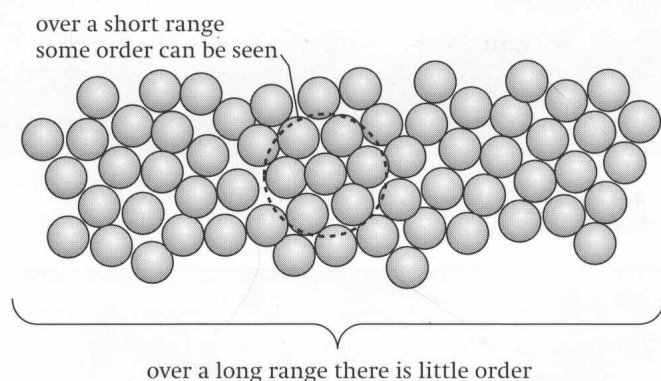
Around 1908 Jean Perrin made observations of Brownian motion in gases. He showed that small particles, much larger than individual molecules but still very small (less than  $10^{-6}$  m in diameter), also went on random walks. This could only be explained along the same lines as Brownian motion in liquids. The particles were being struck by the randomly moving gas molecules.

### How much order is there in a liquid?

The particles in a liquid group together, and it is just this tendency that produces some order in their arrangement (*figure 1.4*). However, the order is over a relatively short range, perhaps over a distance of  $10^{-9}$  m (about 10 molecular diameters). Over greater distances, the degree of order diminishes, i.e. the groups themselves are randomly arranged. We can summarise the situation in this way:

In a liquid there is short-range order, and long-range disorder.

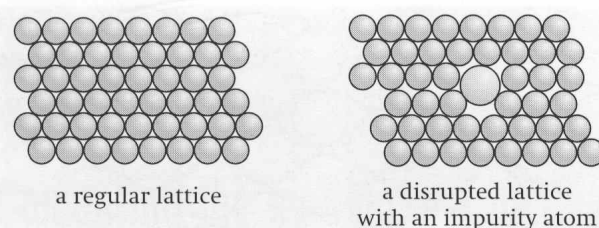
However, as in a gas, the positions of the particles in a liquid are constantly changing; so membership of the groups is always changing.



● **Figure 1.4** Order and disorder in a liquid.

### The arrangement of particles in a solid

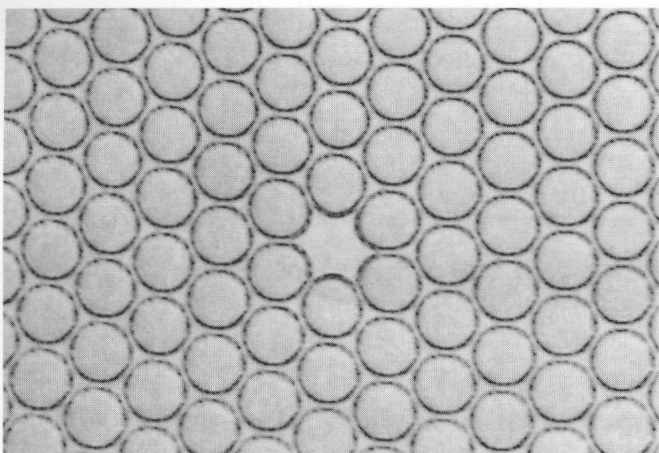
First, a reminder: there are many types of solid, whose properties depend on the particles that they contain. For example, metal crystals consist of lattices of atoms, which are best viewed as positive ions existing in a 'sea' of electrons; ionic



● **Figure 1.5** Disorder can appear even in very orderly solids.

substances like sodium chloride have lattices built from positively and negatively charged ions ( $\text{Na}^+$  and  $\text{Cl}^-$ ); iodine crystals have a lattice of iodine molecules,  $\text{I}_2$ ; graphite crystals contain layers of hexagonal rings of carbon atoms; and diamond is a giant lattice of carbon atoms all bonded in a tetrahedral arrangement. In general, metals and ionic substances have high melting points (although there are exceptions); molecular crystals (like iodine) have low melting points; and giant lattices of interlocked atoms (like diamond) have very high melting points. For the sake of keeping our description of solids fairly simple, we shall use metal crystals as our examples of solids. Much of what we shall say about the structures of metals applies to other types of solids (and when it doesn't, we shall say why).

When a liquid metal starts to crystallise, the atoms begin to fit together in regular patterns. A particularly simple pattern is shown in *figure 1.5*. It is clear that this is a very orderly arrangement. However, even at 0 K, the atoms are not completely still; they vibrate about the same average position. The very regular packing of particles in a solid extends over far greater distances than in a liquid. However, eventually the regularity breaks down. This can happen because of impurity atoms getting in the way, and blocking the normal pattern. It can also happen when crystals start to grow in several places and grow towards each other. Where the crystalline regions meet, the two lattices may not meet exactly. We can visualise such situations in the laboratory using bubble rafts, like that in *figure 1.6*. The study of order and disorder in metal crystals is of huge importance in industry. Variations in a metal lattice can lead to greater strength or, more worryingly, to metal fatigue (*figure 1.7*).



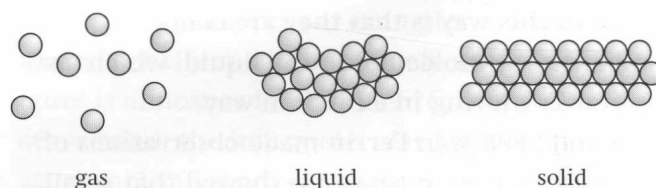
● **Figure 1.6** Bubble rafts can be used to show perfect order and imperfections.



● **Figure 1.7** Metal fatigue in the rails is likely to have contributed to the October 2000 Hatfield train crash.

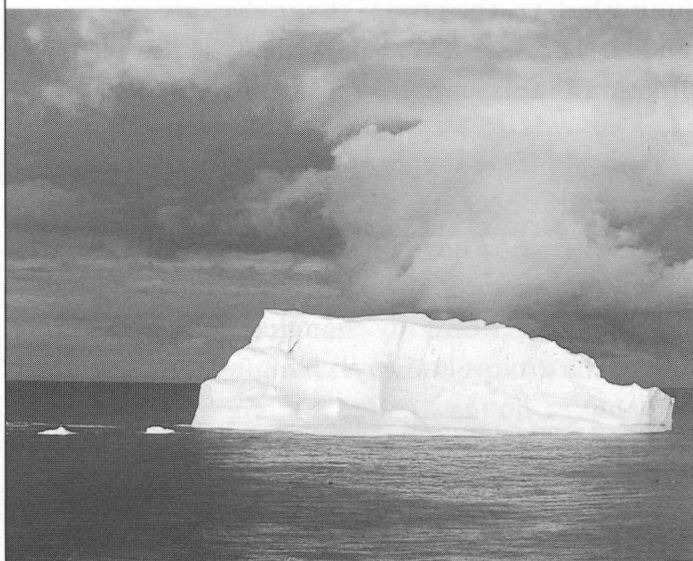
## Differences in properties of solids, liquids and gases

As we have seen, the particles in a gas are, on average, much further apart than they are in a liquid or solid (figure 1.8). There is very little difference



● **Figure 1.8** The arrangements of particles in a gas, a liquid and a solid.

between the spacing of atoms in liquids and solids; that is why both liquids and solids are hard to compress. Also, the particles in a gas travel very much faster than those in a liquid. The differences in spacing, and in speed, are the main reason for the different properties of the three states of matter (figure 1.9 and table 1.1). For example, notice that gases are not very good conductors of heat. For heat to be conducted by atoms or molecules, the movement energy of the molecules must be



● **Figure 1.9** Cloud, sea and iceberg – the three states of matter.

	<b>Solids</b>	<b>Liquids</b>	<b>Gases</b>
Amount of order of arrangement of particles	Very orderly	Short-range order, longer-range disorder	Almost complete disorder
Shape	Fixed	Takes shape of container	No shape
Position of particles	Fixed; no movement from place to place	Some movement from place to place	Always moving rapidly from place to place
Spacing of particles	Close ( $\approx 10^{-10}$ m)	Close ( $\approx 10^{-10}$ m)	Far apart ( $\approx 10^{-8}$ m)
Compressibility	Very low	Very low	High
Conduction of heat	Metals and graphite very good; others poor	Metals very good; others poor	Very poor

● **Table 1.1** Comparison of properties of the three states of matter.



passed on from one to another. This requires the molecules to collide, which happens less easily in a gas than in a liquid.

In a solid the particles are held in position by the overall effects of the attractions and repulsions of their neighbours. Even so, the particles do have some movement. They vibrate to and fro, although on average they keep the same position. As the temperature increases, they vibrate more violently, and they pass on the energy of their vibrations to their neighbours. However, the only solids that conduct heat very well are those that have electrons that can move from place to place. Especially, metals have many **free electrons** that can carry their movement energy with them even though the ions themselves are stuck in one place. That is, metals conduct heat well because of their free electrons, not as a result of vibrations of the particles.

Owing to the large amount of empty space in a gas, it is fairly easy to squeeze the molecules into a smaller volume; so gases are easily compressed.

Liquids and solids have their molecules already very close together, so they are very difficult to compress.

## Comparing the melting and boiling points of substances

You will find some representative examples of melting and boiling points in *table 1.2*. A column showing the relative molecular masses ( $M_r$ ) of the molecules has been included in the table. If you look carefully, you will see that there is a *general* rule that governs the values:

The higher the relative molecular mass, the higher the melting point and the higher the boiling point.

One reason why melting and boiling points tend to increase with mass is that, the greater the mass of a molecule, the more electrons it

	Relative molecular mass, $M_r$	Melting pt ( $^{\circ}\text{C}$ )	Boiling pt ( $^{\circ}\text{C}$ )		Relative molecular mass, $M_r$	Melting pt ( $^{\circ}\text{C}$ )	Boiling pt ( $^{\circ}\text{C}$ )
<i>Elements</i>				<i>Compounds</i>			
Helium, He	4		-269	Methane, $\text{CH}_4$	16	-182	-161
Neon, Ne	20	-249	-249	Ethane, $\text{C}_2\text{H}_6$	30	-183	-88
Argon, Ar	37	-189	-186	Propane, $\text{C}_3\text{H}_8$	44	-189	-42
Krypton, Kr	84	-157	-152	Butane, $\text{C}_4\text{H}_{10}$	58	-138	0
Fluorine, $\text{F}_2$	38	-220	-188	Methanol, $\text{CH}_3\text{OH}$	32	-98	65
Chlorine, $\text{Cl}_2$	71	-101	-34	Ethanol, $\text{C}_2\text{H}_5\text{OH}$	46	-68	79
Bromine, $\text{Br}_2$	160	-7	58	Propan-1-ol, $\text{C}_3\text{H}_7\text{OH}$	60	-78	97
Iodine, $\text{I}_2$	230	114	183	Butan-1-ol, $\text{C}_4\text{H}_9\text{OH}$	74	-89	118
Carbon (diamond), C	12	3550	4830	Hydrogen fluoride, HF	20	-83	20
Silicon, Si	28	1410	2680	Hydrogen chloride, HCl	36.5	-114	-85
Germanium, Ge	73	940	2830	Hydrogen bromide, HBr	81	-87	-67
Tin (white), Sn	119	232	2690	Hydrogen iodide, HI	116	-51	-35
Oxygen, $\text{O}_2$	32	-219	-183	Water, $\text{H}_2\text{O}$	18	0	100
Sulphur, S	32	114.5	444.6	Hydrogen sulphide, $\text{H}_2\text{S}$	34	-85	-60
Selenium, Se	79	217	685	Hydrogen selenide, $\text{H}_2\text{Se}$	81	-66	-42
Tellurium, Te	128	450	1390	Hydrogen telluride, $\text{H}_2\text{Te}$	130	-49	-2

● **Table 1.2** Melting points and boiling points of some elements and compounds.

(Values have been measured at standard atmospheric pressure, 100 kPa.)

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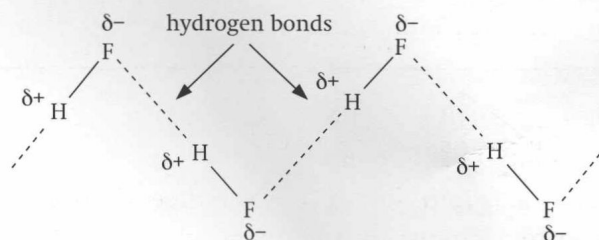
possesses. It is one of the features of large molecules that their electron clouds are more spread out (diffuse), and it is just this type of molecule that has large forces between instantaneous dipoles. These forces are called **instantaneous dipole forces** (more correctly 'instantaneous dipole-induced dipole forces', and are also known as **van der Waals' forces**). Thus, as molecules get heavier, the instantaneous dipole forces become greater, and tend to keep the molecules together.

However, there are many exceptions to the general rule. In particular, you should know that:

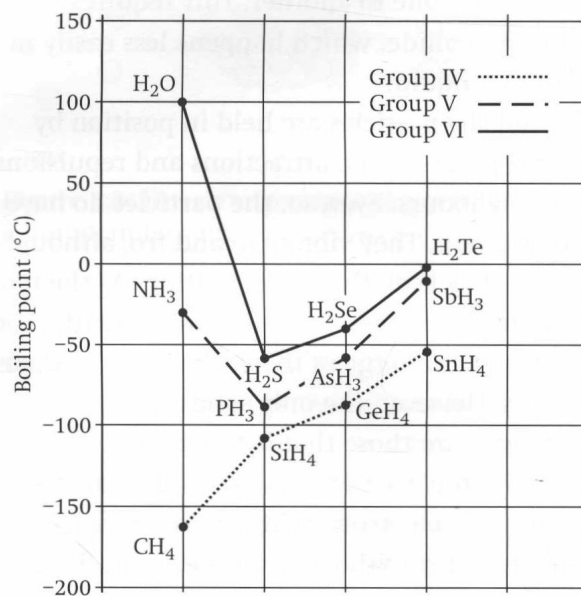
Where melting or boiling points are higher than expected, look for very strong intermolecular forces at work, especially hydrogen bonds.

There are two important examples that you should know about:

- **Hydrogen fluoride**, HF, is rather like water in that its boiling point is far above those of the other hydrides of the halogens. The reason is, again, *hydrogen bonding*. Fluorine is the most electronegative of all the elements, and the hydrogen fluoride molecule is extremely polar. That is, the fluorine atom attracts the pair of electrons in the H-F bond towards itself. The bonding pair spend most of the time nearer the fluorine, thus giving the atom an excess of negative charge. The hydrogen atom has its nucleus (a single proton) only partially surrounded by electrons, and therefore it has an excess positive charge. We show the slight positive and negative charges by the symbols  $\delta^+$  ('delta-plus') and  $\delta^-$  ('delta-minus'). The hydrogen bonds are the forces of attraction between the opposite charges and are shown by the dotted lines in *figure 1.10*.

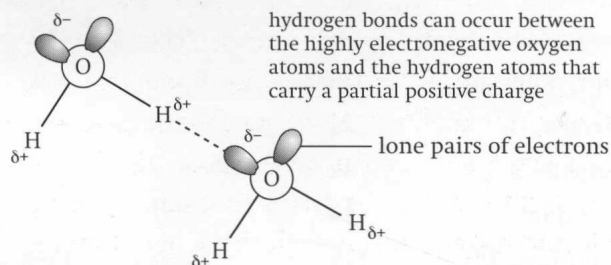


- **Figure 1.10** A representation of the structure of solid hydrogen fluoride, where the molecules take up a zig-zag shape. The molecules are held together by hydrogen bonds.

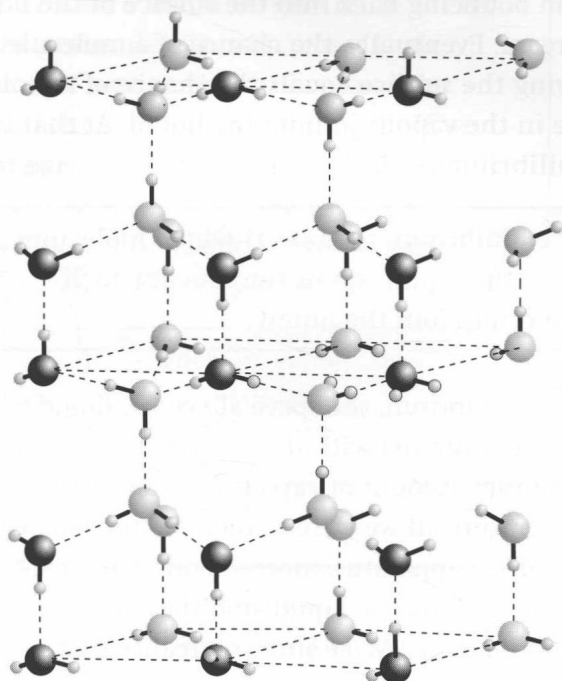


- **Figure 1.11** The boiling points of the Group IV, V and VI hydrides.

- **The Group VI hydrides**, especially water, H<sub>2</sub>O. Compared to the other hydrides of Group VI, the melting and boiling points of water are remarkably high. (The values are shown in *table 1.2*, and a graph of the data is included in *figure 1.11*.) The reason for this lies in hydrogen bonding (*figure 1.12*). In every one of its states, water molecules can hydrogen bond together. In ice the regular arrangement of the lattice leaves a large amount of free space. *Figure 1.13* illustrates the extent of the free space. Because the water molecules in ice are not so close together as in liquid water, ice is less dense than liquid water. In liquid water there is a tremendous amount of order compared to other liquids. Although the pattern of hydrogen bonding is always changing, water molecules are held together much more tightly than are molecules in most other liquids.



- **Figure 1.12** The origin of hydrogen bonding in water.



● **Figure 1.13** The structure of ice.

### SAQ 1.1

As shown in figure 1.11, water and ammonia have boiling points much higher than those of the other hydrides of the elements in their Groups. However, the boiling point of methane,  $\text{CH}_4$ , is lower than those of the other hydrides of Group IV. What is the reason for the difference?

## Why gases liquefy, and solids melt

When two molecules are far apart, they move completely independently; neither will feel the presence of the other. However, if they come closer, then intermolecular forces get to work. The two molecules will attract one another. You will have learned about attractive intermolecular forces when you studied instantaneous dipole forces in *Chemistry 1*, chapter 3. Also, you should have come across hydrogen bonding and dipole–dipole interactions as intermolecular forces that tend to bring molecules together. However, think about molecules coming *very* close together. The ‘outside’ of a molecule is really a layer of negatively charged electrons: the **electron cloud**. When molecules

approach closely, the electron clouds repel one another. It is the great strength of the repulsion that puts a limit on how close the atoms can get.

If two molecules collide with a great deal of energy, the negatively charged electron clouds get squeezed together and the resulting repulsion pushes them apart. Indeed, in a gas the force is so great that it overcomes the (attractive) intermolecular forces. Thus the molecules return to their life of rushing round at random in the body of the gas.

On the other hand, at lower temperatures the speeds of the molecules are lower and the force of collisions can be much less. There is a better chance of the intermolecular forces equalling, and indeed being greater than, the repulsive forces as the molecules collide. When this happens the molecules will not spring apart. Rather, they will remain close together and we see the gas turning to a liquid.

The molecules of different gases have their own characteristic intermolecular forces, and repulsive forces between their electron clouds. Therefore the temperature at which the forces between colliding molecules become low enough for the instantaneous dipole forces to win is different for every gas; i.e. different gases liquefy at different temperatures.

We can turn this line of argument on its head, and explain the change of liquid to gas by discussing the two opposing forces as the temperature of a liquid increases to its boiling point (see SAQ 1.2).

### SAQ 1.2

- Use your knowledge of intermolecular attractions and repulsions to explain why liquids turn into gases as the temperature increases.
- Why does every substance have its own particular boiling point?

## Explaining changes of state

Everyone is familiar with the change of liquid water to a vapour. This happens when water evaporates from a puddle, or when washing dries on a windy or sunny day, or when water boils in a kettle. Likewise, most people in industrialised countries convert liquid water into a solid, ice, by



cooling water in a freezer. To understand why, and how, a substance changes state you need to know about intermolecular forces that attract molecules to each other, and the repulsions between the electron clouds when molecules get very close together. However, you also need to understand two further ideas: the first is that of *equilibrium*, and the second is the idea of *vapour pressure*.

We can bring these together by thinking about an experiment to measure the vapour pressure of a liquid, illustrated in *figure 1.14*. (You are not expected to know the details of the experiment for your examinations.) The idea is to introduce a small quantity of the liquid into a tube filled with mercury. (Owing to the high density of mercury, the liquid will float to the top.) Without the liquid, there would be a vacuum above the mercury, and, at standard conditions, the column of mercury would be 760 mm tall. With the liquid present, some of the molecules escape into the space. Once in the vapour, they exert a pressure, and the mercury is pushed down slightly; i.e. the height of the mercury column is reduced. The difference between the heights (once some corrections are made for the presence of the liquid) is the vapour pressure of the liquid.

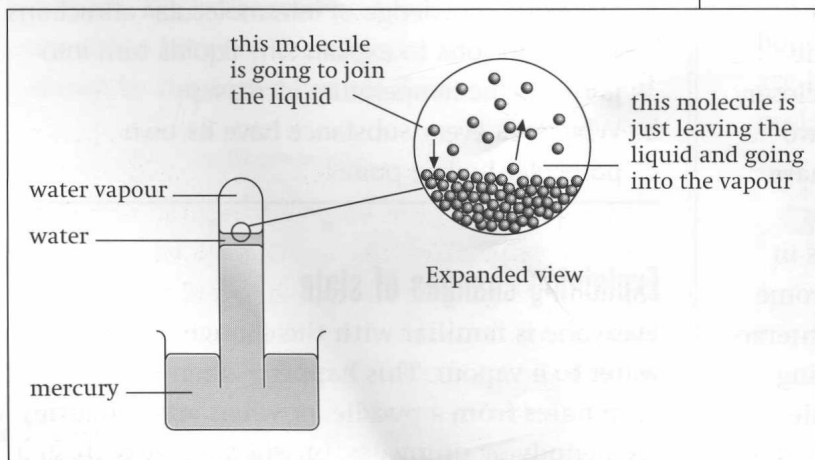
The molecules that escape from the surface of the liquid tend to have higher than average energies – that is why they escape. If a molecule has lower than, or about the same as, the average energy it is unlikely to escape the clutches of the other molecules – the intermolecular forces will hold it back. However, as more molecules escape into the space above the mercury, the chances of

them bouncing back into the surface of the liquid increase. Eventually, the chance of a molecule leaving the surface equals the chance of a molecule in the vapour joining the liquid. At that time, **equilibrium** is reached (see *Chemistry 1*, page 180):

At equilibrium, the rate at which molecules leave the liquid equals the rate at which molecules join the liquid.

At equilibrium, the space above the liquid has become saturated with the vapour – it contains the maximum amount of vapour possible at the given temperature. (If we were to increase the temperature of the apparatus, more vapour could exist in the space above the liquid, and the vapour pressure would increase.) Make sure you realise that equilibrium is a *dynamic* process: there is a great deal of change going on with molecules constantly leaving and joining the liquid. However, they do so at the same rate (many millions per second). *Figure 1.14* shows how evaporation of a liquid can take place under equilibrium conditions.

However, equilibrium will not always be achieved. For example, on a warm, windy day, wet clothes dry very quickly because the atmosphere is not saturated with water vapour as it would be on a cold, wet day. There are (relatively) so few water molecules in the atmosphere that they have little chance of going back on to the clothes once they have left the surface, or of water molecules already in the air sticking to the clothes. That is, wet clothes drying on a line will not reach equilibrium like the liquid in *figure 1.14*.



● **Figure 1.14** Measuring the vapour pressure of a liquid (note that the diagram is not to scale).



● **Figure 1.15** Water boiling under reduced pressure. Water will boil even at room temperature if the pressure is low enough.

If we heat a liquid we give more energy to the molecules. This increases their chance of leaving the surface, and the liquid will evaporate more quickly. If we continue to heat the liquid, it will eventually boil (*figure 1.15*). We shall not prove it, but the condition for a liquid to boil is that:

A liquid boils when its vapour pressure equals atmospheric pressure.

We shall return to this point later (see page 19).

### SAQ 1.3

If you put a little alcohol (ethanol) or propanone on the back of your hand, the liquid will evaporate and you will feel the back of your hand get cold. Why does evaporation lead to cooling? [*Hint: Think about the range of energies the molecules possess, and why even the less energetic particles eventually evaporate.*]

### SAQ 1.4

What are clouds made of? Briefly explain why clouds form, and why they often lead to rain falling.

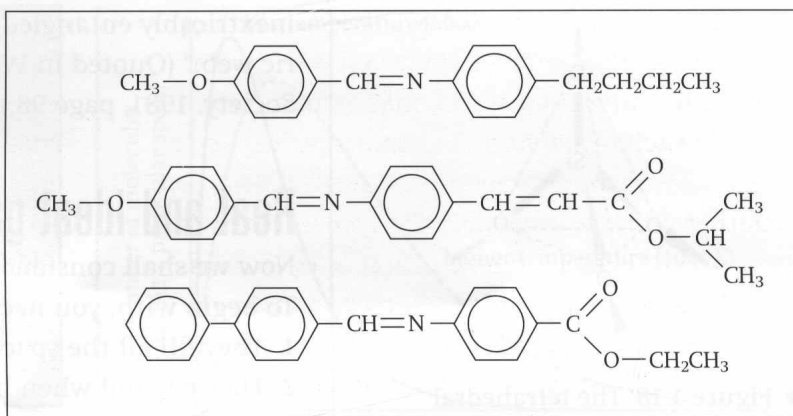
## Some remarkable substances

In this section we shall briefly consider some substances that are difficult to classify as a solid, liquid or gas.

### Liquid crystals

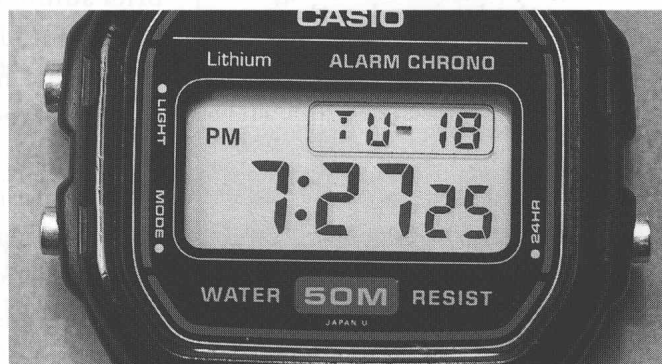
It seems a contradiction to call a crystal 'liquid'. We expect crystals to be solids, and certainly not liquids. Essentially, **liquid crystals** are liquids that have sufficient long-range order in them to make them behave like a solid. However, they will only behave like a solid over a certain range of temperatures. Usually a liquid crystal is made from molecules that are long, thin and not very symmetrical. You will find some examples in *figure 1.16*.

The intermolecular forces must be strong enough to hold the molecules together, but not so strong as to restrict their movement too much. The unsymmetrical nature of the molecules leads



● **Figure 1.16** Examples of molecules that make liquid crystals.

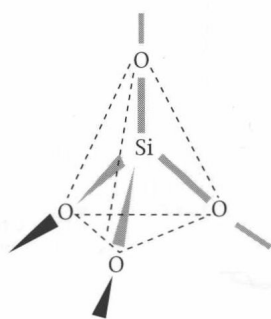
to an unsymmetrical packing of the molecules. The very useful property of liquid crystals is that the arrangement of the molecules can be upset by very slight changes in their surroundings. Especially, in the liquid crystals used in calculators, digital watches and computer displays (*figure 1.17*), the molecules rearrange themselves when the crystal is subjected to a small electric field. The rearrangement changes the way the crystal absorbs light.



● **Figure 1.17** A liquid crystal display.

### Glass

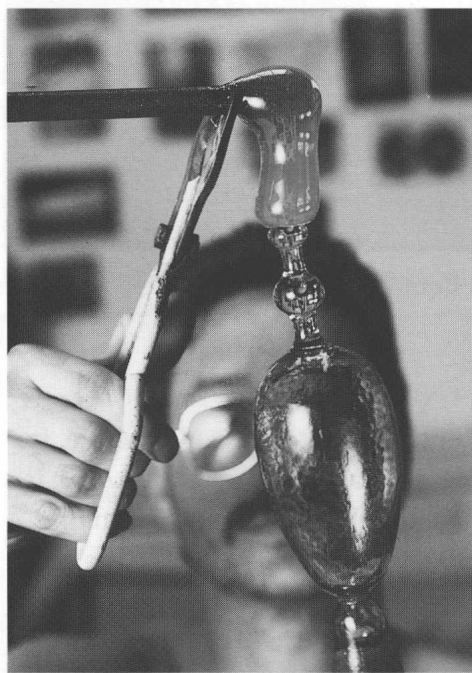
Glass is a most unusual material. For example, it allows light to pass through it very easily. Also, it melts over a range of temperatures and remains viscous (rather like treacle). This allows glass to be 'blown into many different shapes (*figure 1.19*), or to be rolled into sheets for use in windows. The basic building block of ordinary glass is a tetrahedron built from a silicon atom with four oxygen atoms around it (*figure 1.18* on page 10). The tetrahedra join to give a three-dimensional interlocking structure that gives glass its high



- **Figure 1.18** The tetrahedral group of a central silicon atom and four oxygen atoms that is the basis for the structure of glass.

viscosity. However, unlike a normal solid, glass has no long-range order in its structure.

In 1880 the Irish physicist John Tyndall (who had a 'professional interest' in ice, as he was an extremely keen mountaineer) compared ice and glass in this way: "The ice is music, the glass is noise – the ice is order, the glass is confusion. In the glass, molecular forces constitute an



- **Figure 1.19** A glass blower. Glass remains viscous over a large range of temperatures, which allows it to be blown into an amazing variety of shapes.

inextricably entangled skein, in the ice they are woven to a symmetric web." (Quoted in W. H. Brock et al., *John Tyndall*, Royal Dublin Society, 1981, page 98.)

## Real and ideal gases

Now we shall consider the properties of gases in greater depth.

To begin with, you need to know that gases show these properties:

- 1 They fill all the space open to them.
- 2 They expand when heated.
- 3 They exert a pressure on the walls of their containers.
- 4 The pressure changes as the temperature changes.

Shortly, we shall deal with each of these in more detail; but you will find that we shall spend only part of the time discussing the properties of real gases such as hydrogen, oxygen, methane and so on. Real gases are complicated things, and it can be helpful at first to use a simplified model of a gas. In fact, much of this unit will be about 'gases' that do not exist in the real world – these are gases that we call *ideal gases*.

## The behaviour of ideal gases

Ideal gases have some, but not all, of the properties of real gases. A brief summary of the characteristics of an **ideal gas** is given in *box 1A*. An ideal gas is a gas in which there are no intermolecular forces, and in which the molecules don't take up any space themselves (we regard them as points). Also, we assume that the molecules do not change their total kinetic energy when they bump into each other; this is what we mean if we say that the collisions are 'perfectly elastic'. No real gas is ideal, although some come close to ideal behaviour, e.g. helium. You will find that we can compare real and ideal gases, and from their different characteristics we can learn a great deal about real gases. The key idea is to explain why real gases are different from ideal gases – but more of this later.

### Box 1A Key assumptions about ideal gases

*In an ideal gas:*

- the molecules have mass, but negligible size;
- there are no intermolecular forces;
- the collisions between the molecules are perfectly elastic.

## The kinetic theory of gases

You will already know that the molecules in a gas are in a constant state of random motion. This feature of gases is one of the main foundations of the **kinetic theory of gases**. A statement of the main features of the kinetic theory of gases is given in *box 1B*. As far as we know, the kinetic theory of gases is an extremely well-supported theory. There is a great deal of evidence to show that, to all intents



### Box 1B The kinetic theory of gases

#### Main idea:

- Gases consist of molecules in a constant state of random motion.

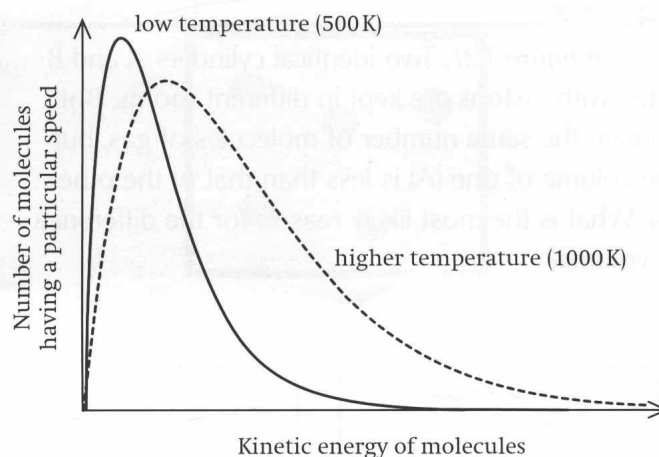
#### Related ideas:

- The pressure of a gas is due to the collisions of the molecules with the walls of the container.
- The molecules travel in straight lines until they collide with one another, or with the walls of the container.
- In these collisions, the total kinetic energy of the molecules does not change.

and purposes, the motion of gas molecules is random. This means that in any gas, on average, there will be as many molecules moving in one direction as in any other direction.

You might like to know that the average speed of gas molecules is of the order of  $500 \text{ m s}^{-1}$  at room temperature. The lighter the molecule, the greater the average speed (and vice versa). For example, hydrogen molecules have an average speed somewhat above  $1500 \text{ m s}^{-1}$ , and carbon dioxide molecules have an average speed nearer to  $350 \text{ m s}^{-1}$ .

There is a wide range of energies among the molecules in a gas. Some move very rapidly, and much faster than the average, and some move very much more slowly than the average. When a gas is heated, *on average* all the molecules increase their kinetic energies (i.e. move faster); but this does not mean that they *all* increase. Always, some will pick up more energy than others. Indeed, during a collision between two molecules, one of them may move off with a greater speed, and one with a lower speed than before. However, the majority move near to the average speed. The way the kinetic energy of the molecules varies with temperature is shown in figure 1.20. As the temperature goes up, the average energy of all the molecules increases, but the distribution of speeds, and therefore kinetic energies, spreads out. Especially, the proportion of molecules with high kinetic energies increases. If you look carefully at the shapes of the graphs, you will see that they are not quite symmetrical – the curve stretches out more at higher than at lower energies. (This is an important point when explaining how changes in temperature influence the rates of chemical reactions: see *Chemistry 1*, page 169.)



● **Figure 1.20** How the distribution of the kinetic energy of the molecules in a gas changes with temperature.

### SAQ 1.5

Use ideas from the kinetic theory of gases to answer these two questions:

- What happens to the average kinetic energy of the molecules in a gas as the temperature increases?
- What might happen to the kinetic energy of any individual molecule in a gas as the temperature increases?

### The pressure and volume of an ideal gas

The pressure of a gas is caused by the collisions of the molecules with the walls of its container. By doing some mathematics, it is possible to show that the pressure of an ideal gas depends on three factors:

- 1 the number of molecules per unit volume (i.e. the concentration of the gas);
- 2 the mass of the molecules;
- 3 their speed.

This should make sense to you because, if there are more molecules present in a given volume, there should be more collisions with the walls, so the pressure should increase. Likewise, if the molecules have a greater momentum (mass times speed), the harder they will bounce off the walls. Therefore, they exert a greater force on the walls, and cause the pressure to increase.