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The Free High School Science Texts: Textbooks for High School Students Studying the Sciences Chemistry Grades 10 - 12

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Contents

Chapter 2

What are the objects around us made of? - Grade 10

2.1 Introduction: The atom as the building block of matter

We have now seen that different materials have different properties. Some materials are metals and some are non-metals; some are electrical or thermal conductors, while others are not. Depending on the properties of these materials, they can be used in lots of useful applications. But what is it exactly that makes up these materials? In other words, if we were to break down a material into the parts that make it up, what would we find? And how is it that a material's microscopic structure is able to give it all these different properties?

The answer lies in the smallest building block of matter: the **atom**. It is the *type* of atoms, and the way in which they are *arranged* in a material, that affects the properties of that substance.

It is not often that substances are found in atomic form. Normally, atoms are bonded to other atoms to form **compounds** or **molecules**. It is only in the *noble gases* (e.g. helium, neon and argon) that atoms are found individually and are not bonded to other atoms. We will look at the reasons for this in a later chapter.

2.2 Molecules

Definition: Molecule
A molecule is a group of two or more atoms that are attracted to each other by relatively strong forces or bonds

Almost everything around us is made up of molecules. *Water* is made up of molecules, each of which has two hydrogen atoms joined to one oxygen atom. *Oxygen* is a molecule that is made up of two oxygen atoms that are joined to one another. Even the food that we eat is made up of molecules that contain atoms of elements such as carbon, hydrogen and oxygen that are joined to one another in different ways. All of these are known as **small molecules** because there are only a few atoms in each molecule. **Giant molecules** are those where there may be millions of atoms per molecule. Examples of giant molecules are *diamonds*, which are made up of millions of carbon atoms bonded to each other, and *metals*, which are made up of millions of metal atoms bonded to each other.

2.2.1 Representing molecules

The structure of a molecule can be shown in many different ways. Sometimes it is easiest to show what a molecule looks like by using different types of **diagrams**, but at other times, we may decide to simply represent a molecule using its **chemical formula** or its written name.

1. **Using formulae to show the structure of a molecule**

   A **chemical formula** is an abbreviated (shortened) way of describing a molecule, or some other chemical substance. In chapter 1, we saw how chemical compounds can be represented using element symbols from the Periodic Table. A chemical formula can also tell us the *number* of atoms of each element that are in a molecule, and their *ratio* in that molecule.

   For example, the chemical formula for a molecule of carbon dioxide is:

   $$CO_2$$

   The formula above is called the **molecular formula** of that compound. The formula tells us that in one molecule of carbon dioxide, there is one atom of carbon and two atoms of oxygen. The ratio of carbon atoms to oxygen atoms is 1:2.

   Definition: Molecular formula
   A concise way of expressing information about the atoms that make up a particular chemical compound. The molecular formula gives the exact number of each type of atom in the molecule.

   A molecule of glucose has the molecular formula:

   $$C_6H_{12}O_6$$

   In each glucose molecule, there are six carbon atoms, twelve hydrogen atoms and six oxygen atoms. The ratio of carbon:hydrogen:oxygen is 6:12:6. We can simplify this ratio to write 1:2:1, or if we were to use the element symbols, the formula would be written as CH_2O. This is called the **empirical formula** of the molecule.

   Definition: Empirical formula
   This is a way of expressing the *relative* number of each type of atom in a chemical compound. In most cases, the empirical formula does not show the exact number of atoms, but rather the simplest *ratio* of the atoms in the compound.

   The empirical formula is useful when we want to write the formula for a *giant molecule*. Since giant molecules may consist of millions of atoms, it is impossible to say exactly how many atoms are in each molecule. It makes sense then to represent these molecules using their empirical formula. So, in the case of a metal such as copper, we would simply write Cu, or if we were to represent a molecule of sodium chloride, we would simply write NaCl.

   Chemical formulae therefore tell us something about the *types* of atoms that are in a molecule and the *ratio* in which these atoms occur in the molecule, but they don't give us any idea of what the molecule actually looks like, in other words its *shape*. Another useful way of representing molecules is to use diagrams.

   Another type of formula that can be used to describe a molecule is its **structural formula**. A structural formula uses a graphical representation to show a molecule's structure (figure 2.1).

2. **Using diagrams to show the structure of a molecule**

   Diagrams of molecules are very useful because they give us an idea of the *space* that is occupied by the molecule, and they also help us to picture how the atoms are arranged in the molecule. There are two types of diagrams that are commonly used:

(a) C_4H_{10} (b) C_2H_5

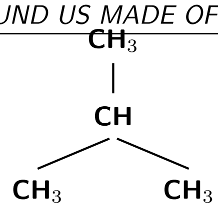

Figure 2.1: Diagram showing (a) the molecular, (b) the empirical and (c) the structural formula of isobutane

- *Ball and stick models*

  This is a 3-dimensional molecular model that uses 'balls' to represent atoms and 'sticks' to represent the bonds between them. The centres of the atoms (the balls) are connected by straight lines which represent the bonds between them. A simplified example is shown in figure 2.2.

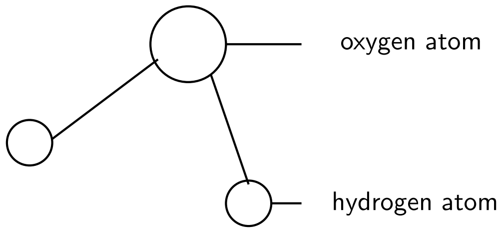

Figure 2.2: A ball and stick model of a water molecule

- *Space-filling model*

  This is also a 3-dimensional molecular model. The atoms are represented by multi-coloured spheres. Space-filling models of water and ammonia are shown in figures 2.3 and 2.4.

Figures 2.3 and 2.4 are some examples of **simple molecules** that are represented in different ways.

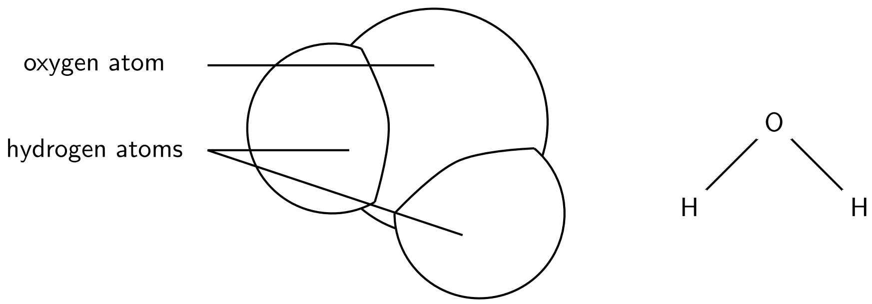

Figure 2.3: A space-filling model and structural formula of a water molecule. Each molecule is made up of two hydrogen atoms that are attached to one oxygen atom. This is a simple molecule.

Figure 2.5 shows the bonds between the carbon atoms in diamond, which is a **giant molecule**. Each carbon atom is joined to four others, and this pattern repeats itself until a complex *lattice* structure is formed. Each black ball in the diagram represents a carbon atom, and each line represents the bond between two carbon atoms.

Interesting Fact

Diamonds are most often thought of in terms of their use in the jewellery industry. However, about 80% of mined diamonds are unsuitable for use as gemstones and are therefore used in industry because of their strength and hardness. These

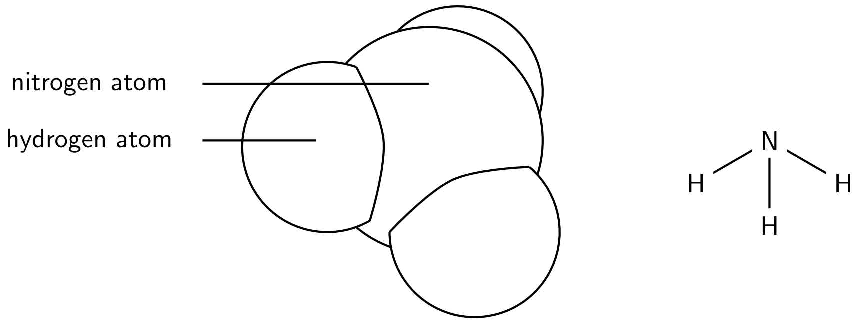

Figure 2.4: A space-filling model and structural formula of a molecule of ammonia. Each molecule is made up of one nitrogen atom and three hydrogen atoms. This is a simple molecule.

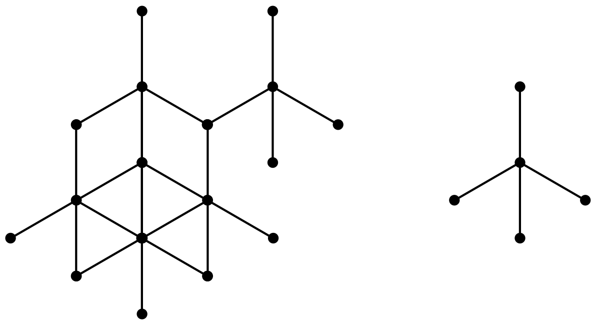

Figure 2.5: Diagrams showing the microscopic structure of diamond. The diagram on the left shows part of a diamond lattice, made up of numerous carbon atoms. The diagram on the right shows how each carbon atom in the lattice is joined to four others. This forms the basis of the lattice structure. Diamond is a giant molecule.

properties of diamonds are due to the strong covalent bonds betwene the carbon atoms in diamond. The most common uses for diamonds in industry are in cutting, drilling, grinding, and polishing.

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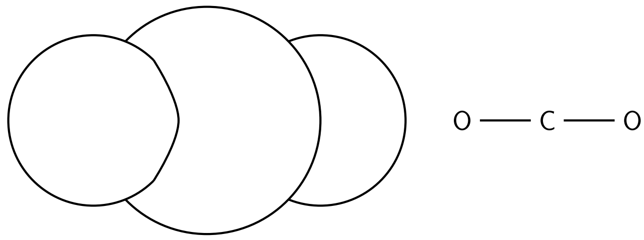
Exercise: Atoms and molecules

1. In each of the following, say whether the chemical substance is made up of single atoms, simple molecules or giant molecules.
   (a) ammonia gas (NH_3)
   (b) zinc metal (Zn)
   (c) graphite (C)
   (d) nitric acid (HNO_3)
   (e) neon gas (Ne_2)
2. Refer to the diagram below and then answer the questions that follow:



(a) Identify the molecule.
(b) Write the molecular formula for the molecule.
(c) Is the molecule a simple or giant molecule?

3. Represent each of the following molecules using its *chemical formula*, *structural formula* and *ball and stick model*.
   (a) H_2
   (b) NH_3
   (c) sulfur dioxide

2.3 Intramolecular and intermolecular forces

When atoms join to form molecules, they are held together by **chemical bonds**. The type of bond, and the strength of the bond, depends on the atoms that are involved. These bonds are called **intramolecular forces** because they are bonding forces *inside* a molecule ('intra' means 'within' or 'inside'). Sometimes we simply call these intramolecular forces chemical bonds.

Definition: Intramolecular force
The force between the atoms of a molecule, which holds them together.

Examples of the types of chemical bonds that can exist between atoms inside a molecule are shown below. These will be looked at in more detail in chapter 4.

- *Covalent bond*

  Covalent bonds exist between non-metal atoms e.g. There are covalent bonds between the carbon and oxygen atoms in a molecule of carbon dioxide.

- *Ionic bond*

  Ionic bonds occur between non-metal and metal atoms e.g. There are ionic bonds between the sodium and chlorine atoms in a molecule of sodium chloride.

- *Metallic bond*

  Metallic bonds join metal atoms e.g. There are metallic bonds between copper atoms in a piece of copper metal.

Intermolecular forces are those bonds that hold *molecules* together. A glass of water for example, contains many molecules of water. These molecules are held together by intermolecular forces. The strength of the intermolecular forces is important because they affect properties such as *melting point* and *boiling point*. For example, the stronger the intermolecular forces, the higher the melting point and boiling point for that substance. The strength of the intermolecular forces increases as the size of the molecule increases.

Definition: Intermolecular force
A force between molecules, which holds them together.

Diagram 2.6 may help you to understand the difference between intramolecular forces and intermolecular forces.

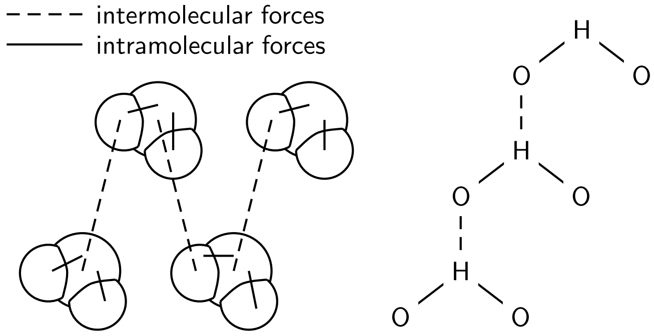


Figure 2.6: Two representations showing the intermolecular and intramolecular forces in water: space-filling model and structural formula.

It should be clearer now that there are two types of forces that hold matter together. In the case of water, there are intramolecular forces that hold the two hydrogen atoms to the oxygen atom *in each molecule of water*. There are also intramolecular forces *between each of these water molecules*. As mentioned earlier, these forces are very important because they affect many of the *properties of matter* such as boiling point, melting point and a number of other properties. Before we go on to look at some of these examples, it is important that we first take a look at the **Kinetic Theory of Matter**.

Exercise: Intramolecular and intermolecular forces

1. Using ammonia gas as an example...
   (a) Explain what is meant by an *intramolecular* force or *chemical bond*.
   (b) Explain what is meant by an *intermolecular* force.
2. Draw a diagram showing three molecules of carbon dioxide. On the diagram, show where the intramolecular and intermolecular forces are.
3. Why is it important to understand the types of forces that exist between atoms and between molecules? Try to use some practical examples in your answer.

2.4 The Kinetic Theory of Matter

The **kinetic theory of matter** is used to explain why matter exists in different *phases* (i.e. solid, liquid and gas), and how matter can change from one phase to the next. The kinetic theory of matter also helps us to understand other properties of matter. It is important to realise that what we will go on to describe is only a *theory*. It cannot be proved beyond doubt, but the fact that it helps us to explain our observations of changes in phase, and other properties of matter, suggests that it probably is more than just a theory.

Broadly, the Kinetic Theory of Matter says that:

- Matter is made up of **particles** that are constantly moving.
- All particles have **energy**, but the energy varies depending on whether the substance is a solid, liquid or gas. Solid particles have the least energy and gas particles have the most amount of energy.
- The **temperature** of a substance is a measure of the *average kinetic energy* of the particles.
- A change in **phase** may occur when the energy of the particles is changed.
- There are **spaces** between the particles of matter.
- There are **attractive forces** between particles and these become stronger as the particles move closer together. These attractive forces will either be intramolecular forces (if the particles are atoms) or intermolecular forces (if the particles are molecules). When the particles are extremely close, repulsive forces start to act.

Table 2.1 summarises the characteristics of the particles that are in each phase of matter.

Table 2.1: Table summarising the general features of solids, liquids and gases.

Property of matter	**Gas**	**Liquid**	**Gas**
Particles	Atoms or molecules	Atoms or molecules	Atoms or molecules
Energy and movement of particles	Particles have high energy and are constantly moving	Particles have less energy than in the gas phase	Low energy - particles vibrate around a fixed point
Spaces between particles	Large spaces because of high energy	Smaller spaces than in gases	Very little space between particles. Particles are tightly packed together
Attractive forces between particles	Weak forces because of the large distance between particles	Stronger forces than in gas. Liquids can be poured.	Very strong forces. Solids have a fixed volume.
Changes in phase	In general a gas becomes a liquid or solid when it is cooled. Particles have less energy and therefore move closer together so that the attractive forces become stronger, and the gas becomes a liquid or a solid	A liquid becomes a gas if its temperature is increased. It becomes a solid if its temperature decreases.	Solids become liquids or gases if their temperature is increased.

Let's look at an example that involves the three phases of water: ice (solid), water (liquid) and water vapour (gas).

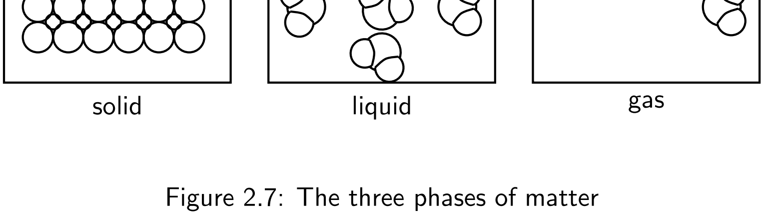


Figure 2.7: The three phases of matter

In a solid (e.g. ice), the water molecules have very little energy and can't move away from each other. The molecules are held close together in a regular pattern called a *lattice*. If the ice is heated, the energy of the molecules increases. This means that some of the water molecules are able to overcome the intermolecular forces that are holding them together, and the molecules move further apart to form *liquid water*. This is why liquid water is able to flow, because the molecules are more free to move than they were in the solid lattice. If the molecules are heated further, the liquid water will become water vapour, which is a gas. Gas particles have lots of energy and are far away from each other. That is why it is difficult to keep a gas in a specific area! The attractive forces between the particles are very weak and they are only loosely held together. Figure 2.8 shows the changes in phase that may occur in matter, and the names that describe these processes.

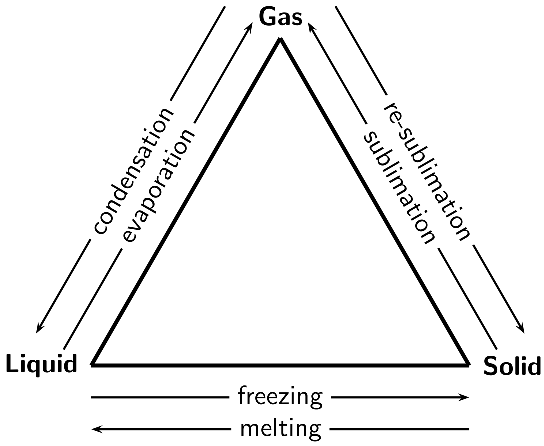


Figure 2.8: Changes in phase

2.5 The Properties of Matter

Let us now look at what we have learned about chemical bonds, intermolecular forces and the kinetic theory of matter, and see whether this can help us to understand some of the macroscopic properties of materials.

1. **Melting point**

Definition: Melting point
The temperature at which a *solid* changes its phase or state to become a *liquid*. The reverse process (change in phase from liquid to solid) is called **freezing**.

In order for a solid to melt, the energy of the particles must increase enough to overcome the bonds that are holding the particles together. It makes sense then that a solid which is held together by strong bonds will have a *higher* melting point than one where the bonds are weak, because more energy (heat) is needed to break the bonds. In the examples we have looked at, metals, ionic solids and some atomic lattices (e.g. diamond) have high melting points, whereas the melting points for molecular solids and other atomic lattices (e.g. graphite) are much lower. Generally, the intermolecular forces between molecular solids are *weaker* than those between ionic and metallic solids.

2. **Boiling point**

Definition: Boiling point
The temperature at which a *liquid* changes its phase to become a *gas*.

When the temperature of a liquid increases, the average kinetic energy of the particles also increases, and they are able to overcome the bonding forces that are holding them in the liquid. When boiling point is reached, *evaporation* takes place and some particles in the liquid become a gas. In other words, the energy of the particles is too great for them to be held in a liquid anymore. The stronger the bonds within a liquid, the higher the boiling point needs to be in order to break these bonds. Metallic and ionic compounds have high boiling points while the boiling point for molecular liquids is lower.

The data in table 2.2 below may help you to understand some of the concepts we have explained. Not all of the substances in the table are solids at room temperature, so for now, let's just focus on the *boiling points* for each of these substances. Of the substances listed, ethanol has the weakest intermolecular forces, and sodium chloride and mercury have the strongest. What do you notice?

Substance	**Melting point (^0C)**	**Boiling point (^0C)**
Ethanol (C_2H_6O)	-114,3	78,4
Water	0	100
Mercury	-38,83	356,73
Sodium chloride	801	1465

Table 2.2: The melting and boiling points for a number of substances

You will have seen that substances such as ethanol, with relatively weak intermolecular forces, have the lowest boiling point, while substances with stronger intermolecular forces such as sodium chloride and mercury, must be heated much more if the particles are to have enough energy to overcome the forces that are holding them together in the liquid or solid phase.

?

Exercise: Forces and boiling point

The table below gives the molecular formula and the boiling point for a number of organic compounds called *alkanes*. Refer to the table and then answer the questions that follow.

Organic compound	**Molecular formula**	**Boiling point (^{0}C)**
Methane	CH_2	-161.6
Ethane	C_2H_6	-88.6
Propane	C_3H_8	-45
Butane	C_4H_{10}	-0.5
Pentane	C_5H_{12}	36.1
Hexane	C_6H_{14}	69
Heptane	C_7H_{16}	98.42
Octane	C_8H_{18}	125.52

Data from: http://www.wikipedia.com

(a) Draw a graph to show the relationship between the number of carbon atoms in each alkane, and its boiling point (Number of carbon atoms will go on the x-axis and boiling point on the y-axis).
(b) Describe what you see.
(c) Suggest a reason for what you have observed.
(d) Why was it enough for us to use 'number of carbon atoms' as a measure of the molecular weight of the molecules?

3. **Density and viscosity**

Density is a measure of the mass of a substance per unit volume. The density of a solid is generally higher than that of a liquid because the particles are hold much more closely together and therefore there are more particles packed together in a particular volume. In other words, there is a greater mass of the substance in a particular volume. In general, density increases as the strength of the intermolecular forces increases. **Viscosity** is a measure of how resistant a liquid is to changing its form. Viscosity is also sometimes described as the 'thickness' of a fluid. Think for example of syrup and how slowly it pours from one container into another. Now compare this to how easy it is to pour water. The viscosity of syrup is greater than the viscosity of water. Once again, the stronger the intermolecular forces in the liquid, the greater its viscosity.

It should be clear now that we can explain a lot of the **macroscopic properties** of matter (i.e. the characteristics we can see or observe) by understanding their **microscopic structure** and the way in which the atoms and molecules that make up matter are held together.

Activity :: Investigation : Determining the density of liquids:

Density is a very important property because it helps us to identify different materials. Every material, depending on the elements that make it up, and the arrangement of its atoms, will have a different density.

The equation for density is:

$$\text{Density} = \text{Mass/Volume}$$

Discussion questions:

To calculate the density of liquids and solids, we need to be able to first determine their mass and volume. As a group, think about the following questions:

- How would you determine the mass of a liquid?
- How would you determine the volume of an irregular solid?

Apparatus:

Laboratory mass balance, 10 ml and 100 ml graduated cylinders, thread, distilled water, two different liquids.

Method:

Determine the density of the distilled water and two liquids as follows:

1. Measure and record the mass of a 10 ml graduated cyclinder.
2. Pour an amount of distilled water into the cylinder.
3. Measure and record the combined mass of the water and cylinder.
4. Record the volume of distilled water in the cylinder
5. Empty, clean and dry the graduated cylinder.
6. Repeat the above steps for the other two liquids you have.
7. Complete the table below.

Liquid	Mass (g)	Volume (ml)	Density (g/ml)
Distilled water			
Liquid 1			
Liquid 2			

Activity :: Investigation : Determining the density of irregular solids:

Apparatus:

Use the same materials and equipment as before (for the liquids). Also find a number of solids that have an irregular shape.

Method:

Determine the density of irregular solids as follows:

1. Measure and record the mass of one of the irregular solids.
2. Tie a piece of thread around the solid.
3. Pour some water into a 100 ml graduated cylinder and record the volume.
4. Gently lower the solid into the water, keeping hold of the thread. Record the combined volume of the solid and the water.
5. Dtermine the volume of the solid by subtracting the combined volume from the original volume of the water only.
6. Repeat these steps for the second object.
7. Complete the table below.

Solid	**Mass (g)**	**Volume (ml)**	**Density (g/ml)**
Solid 1			
Solid 2			
Solid 3			

2.6 Summary

- The smallest unit of matter is the **atom**. Atoms can combine to form **molecules**.
- A **molecule** is a group of two or more atoms that are attracted to each other by chemical bonds.
- A **small molecule** consists of a few atoms per molecule. A **giant molecule** consists of millions of atoms per molecule, for example metals and diamonds.
- The structure of a molecule can be represented in a number of ways.
- The **chemical formula** of a molecule is an abbreviated way of showing a molecule, using the symbols for the elements in the molecule. There are two types of chemical formulae: molecular and empirical formula.
- The **molecular formula** of a molecule gives the exact number of atoms of each element that are in the molecule.
- The **empirical formula** of a molecule gives the relative number of atoms of each element in the molecule.
- Molecules can also be represented using **diagrams**.
- A **ball and stick** diagram is a 3-dimensional molecular model that uses 'balls' to represent atoms and 'sticks' to represent the bonds between them.
- A **space-filling model** is also a 3-dimensional molecular model. The atoms are represented by multi-coloured spheres.
- In a molecule, atoms are held together by **chemical bonds** or **intramolecular forces**. Covalent bonds, ionic bonds and metallic bonds are examples of chemical bonds.
- A **covalent bond** exists between non-metal atoms. An **ionic bond** exists between non-metal and metal atoms, and a **metallic bond** exists between metal atoms.
- **Intermolecular forces** are the bonds that hold *molecules* together.
- The **kinetic theory of matter** attempts to explain the behaviour of matter in different phases.
- The theory says that all matter is composed of **particles** which have a certain amount of **energy** which allows them to **move** at different speeds depending on the temperature (energy). There are **spaces** between the particles, and also **attractive forces** between particles when they come close together.

- Understanding chemical bonds, intermolecular forces and the kinetic theory of matter, can help to explain many of the **macroscopic properties** of matter.
- **Melting point** is the temperature at which a *solid* changes its phase to become a *liquid*. The reverse process (change in phase from liquid to solid) is called **freezing**. The stronger the chemical bonds and intermolecular forces in a substance, the higher the melting point will be.
- **Boiling point** is the temperature at which a liquid changes phase to become a gas. The stronger the chemical bonds and intermolecular forces in a substance, the higher the boiling point will be.
- **Density** is a measure of the mass of a substance per unit volume.
- **Viscosity** is a measure of how resistant a liquid is to changing its form.

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Exercise: Summary exercise

1. Give one word or term for each of the following descriptions.
   (a) The property that determines how easily a liquid flows.
   (b) The change in phase from liquid to gas.
   (c) A composition of two or more atoms that act as a unit.
   (d) Chemical formula that gives the relative number of atoms of each element that are in a molecule.
2. For each of the following questions, choose the one correct answer from the list provided.
   A Ammonia, an ingredient in household cleaners, can be broken down to form one part nitrogen (N) and three parts hydrogen (H). This means that ammonia...
      i. is a colourless gas
      ii. is not a compound
      iii. cannot be an element
      iv. has the formula N_3H
   B If one substance A has a melting point that is *lower* than the melting point of substance B, this suggests that...
      i. A will be a liquid at room temperature.
      ii. The chemical bonds in substance A are weaker than those in substance B.
      iii. The chemical bonds in substance A are stronger than those in substance B.
      iv. B will be a gas at room temperature.
3. Boiling point is an important concept to understand.
   a Define 'boiling point'.
   b What change in phase takes place when a liquid reaches its boiling point?
   c What is the boiling point of water?
   d Use the kinetic theory of matter and your knowledge of intermolecular forces, to explain why water changes phase at this temperature.
4. Refer to the table below which gives the melting and boiling points of a number of elements, and then answer the questions that follow. (*Data from http://www.chemicalelements.com*)

Element	**Melting point**	**Boiling point (^{0}C)**
copper	1083	2567
magnesium	650	1107
oxygen	-218.4	-183
carbon	3500	4827
helium	-272	-268.6
sulfur	112.8	444.6

a What state of matter (i.e. solid, liquid or gas) will each of these elements be in at room temperature?

b Which of these elements has the strongest forces between its atoms? Give a reason for your answer.

c Which of these elements has the weakest forces between its atoms? Give a reason for your answer.

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