4.2 Arranging the elements

The periodic table is a list of all the elements, arranged in order of their atomic numbers. The table's structure allows you to predict the properties of a particular element and what its atoms are likely to do in a chemical reaction.

Columns and rows

The **periodic table** lists all of the known elements in order of increasing atomic number. Having only one proton (atomic number 1), hydrogen is the first element. Ununoctium is the last, with 118 protons (atomic number 118). This list of elements is arranged according to atomic number and in columns depending on the number of electrons in the outer shell.

Periods

The horizontal rows in the periodic table are called **periods**. They are shown in Figure 4.2.1. The period number of an element is the same as the number of shells occupied by the electrons in its atoms.

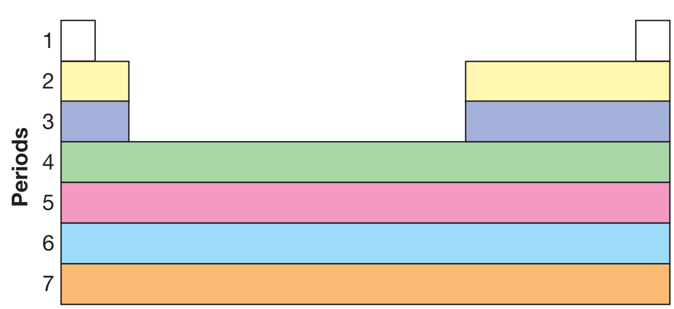


Figure 4.2.1 Periods are the horizontal rows of the periodic table.

Groups

Groups are the vertical columns in the periodic table and are numbered from 1 to 18. An older way of numbering groups was to use the roman numerals I–VIII (Figure 4.2.2). From the group number of an element, you can work out the number of electrons in the outer shell of its atoms.

When using the group numbers 1 to 18, the last digit is usually the number of electrons in the outer shell. So group 12 has two electrons in the outer shell while group 18 has eight.

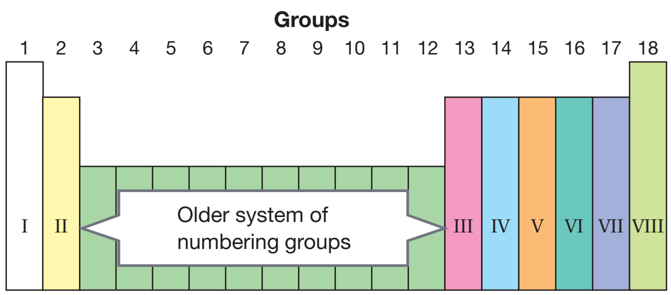


Figure 4.2.2 Groups are the vertical columns of the periodic table.

Placing elements in periods and groups

The electron configuration tells you the period and the group of an element. For example:

- Sodium (element symbol Na) has 11 electrons and an electron configuration of 2,8,1. This configuration shows that sodium atoms have three occupied shells, with one electron in their outer shell. For these reasons, sodium is placed in period 3, group 1.
- Nitrogen (N) has an electron configuration of 2,5. It has two occupied shells and has five outer-shell electrons. It is therefore placed in period 2, group 15.
- Fluorine's (F) electron configuration is 2,7, so it is placed in period 2, group 17.

The connection between electron configuration of the first 18 elements and their position in the periodic table is clear in Table 4.2.1. By arranging the elements this way, the metals are to the left of the periodic table and the non-metals to the right.

Table 4.2.1 Electron configuration of the first 18 elements

Period	Group							
	1	2	13	14	15	16	17	18
1	H 1							He 2
2	Li 2,1	Be 2,2	B 2,3	C 2,4	N 2,5	O 2,6	F 2,7	Ne 2,8
3	Na 2,8,1	Mg 2,8,2	Al 2,8,3	Si 2,8,4	P 2,8,5	S 2,8,6	Cl 2,8,7	Ar 2,8,8

SciFile

The odd one out!

These balloons contain helium atoms. Each helium atom has two electrons occupying a single shell. Helium *should* be in group 2 but is placed in group 18 because of its properties, which are far more like those of group 18 elements than group 2 elements.

4.2

Special blocks

As Figure 4.2.3 shows, the periodic table has three blocks of elements known as the **transition elements**, the **lanthanides** and **actinides**. The placement of elements in these blocks is also based on their electron configuration but is too complex to discuss here.

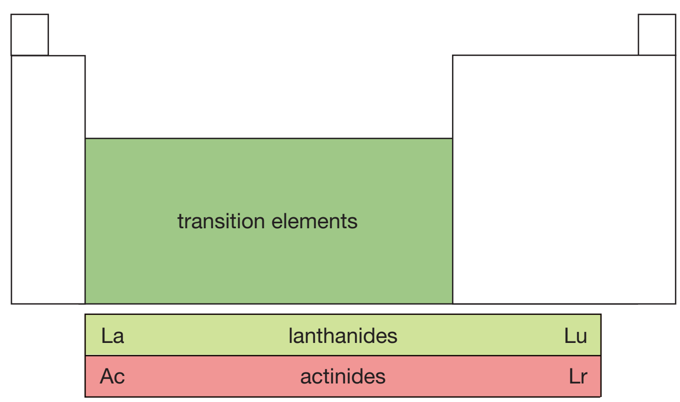


Figure 4.2.3 The transition elements, the lanthanides and actinides, are placed in special blocks. Their placement is also based on electron configuration.

1 p118
2 p119

SCIENCE AS A HUMAN ENDEAVOUR

Nature and development of science

Development of the periodic table

Figure 4.2.4 John Dalton's table of elements and the symbols he gave them

The ancient Greek philosopher Aristotle thought that there were only four elements: earth, water, air and fire. Although scientists now know that these aren't elements, Aristotle's idea began a long search for what substances are made of. As scientific techniques advanced, so did the list of known elements. As more elements were discovered, scientists looked for ways to organise them.

Chemists had long known that certain elements behaved similarly to one another in chemical reactions and in their physical properties. This pattern provided a way of constructing the periodic table that is now used by all chemists worldwide.

The first tables

In 1789, the French chemist Antoine-Laurent de Lavoisier (1743–93) separated the known elements into metals, non-metals and 'earths'. However, his list of 33 elements also included light and a liquid called 'caloric' that was thought to carry heat from hot to cold bodies. Caloric is now known not to exist. Light does exist but is not an element. In 1808, the English chemist John Dalton (1766–1844) went further by giving each of the 36 known elements its own chemical symbol and organising them in order of their mass. You can see his symbols in Figure 4.2.4.

SciFile

Dalton's eyeballs

Dalton was colour blind and suspected that it was because his eyes were filled with blue liquid! He instructed his doctor to dissect his eyes after death to determine if they were. No blue liquid was found but in 1995 the gene that leads to colour-blindness was isolated from his preserved eyes.

Fifty-five elements had been discovered by 1829. Some elements had very similar properties and seemed to be related. For example, chlorine, bromine and iodine (shown in Figure 4.2.5) reacted with other chemicals in a similar way. For this reason, the German chemist Johann Dobereiner (1780–1849) bunched them together into a family of three. Other related elements were also bunched into threes. Dobereiner called his bunches 'triads'. Lithium, sodium and potassium formed a triad, while calcium, strontium and barium formed another.

Figure 4.2.5 The properties of chlorine, bromine and iodine are so similar that Dobereiner placed them in the same triad.

ENDEAVOUR

In 1864, the English scientist John Newlands (1837–98) arranged the 60 known elements in order of their mass, forming a table of seven columns. Every eighth element was placed on a new row and so his arrangement was known as the law of octaves (Figure 4.2.6). However, some boxes in his table ended up with more than one element in them.

Figure 4.2.6 Musical notes repeat every eight notes. Newlands organised his elements in a similar way, with a new 'octave' starting every eight elements.

Success

In 1869, the Russian chemist Dmitri Ivanovich Mendeleev (1834–1907) constructed a table of rows and columns like those of his favourite card game, called solitaire or patience. Each element had its own box. He placed them horizontally according to their atomic masses and vertically according to their properties. He believed in the periodicity (repetition) of the properties of the elements and this arrangement placed elements with similar properties in the same columns. He then realised that some atomic masses were wrong and instead placed those elements where their resemblance to other elements suggested they should go. He also left gaps in his table for undiscovered elements and predicted what their properties might be. For example, he predicted that 'eka-silicon' would eventually be discovered. When isolated in 1886, eka-silicon's properties compared remarkably well with Mendeleev's predictions. Eka-silicon is now known as germanium. Table 4.2.2 compares its predicted and measured properties. Similarly, the properties he predicted for eka-aluminium, eka-boron and eka-manganese were close to those of gallium (isolated in 1875), scandium (1879) and technetium (1937) respectively.

Table 4.2.2 Comparing eka-silicon with germanium

Physical property	Eka-silicon: Mendeleev's predictions	Germanium: properties as measured
Colour	Grey	Grey-white
Atomic mass	72	72.61
Melting point (°C)	High	947
Boiling point (°C)	Below 100	84
Density (g/cm^3)	5.5	5.35

SciFile: An eccentric scientist!

Figure 4.2.7 shows Mendeleev's extravagant hair and beard, which he only trimmed once a year. Mendeleev (and his mother) hitch-hiked over 6000 km in 1848 to get to his first day of university in St Petersburg. While there, he spent some of his time perfecting the perfect vodka!

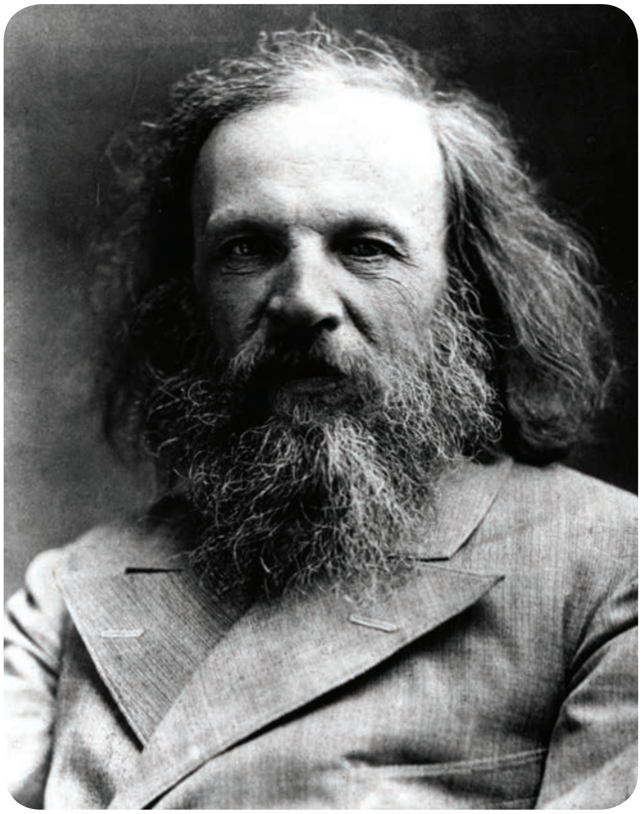


Figure 4.2.7 Dmitri Ivanovich Mendeleev constructed the first modern periodic table.

A similar table to Mendeleev's was constructed in 1868–69 by the German chemist Lothar Meyer (1830–95).

Our modern table

Mendeleev and Meyer arranged the elements in order of increasing atomic mass. To construct his table in 1913, the English physicist Henry Moseley (1887–1915) used atomic numbers instead. An extended version of Moseley's table is still used today.

The discovery of electrons by English scientist J.J. Thomson (1856–1940) in 1897 and the discovery of electron configuration around 1920 eventually led to the periodic table being arranged according to the way electrons were arranged in their shells. Amazingly, this version was identical to the earlier table based on physical and chemical properties! This led scientists to realise that chemistry was largely about electrons and how they behave (particularly the outer-shell electrons).

4.3

4.2 Unit review

Remembering

1 **List** five:
   a group 15 elements
   b period 2 elements
   c common transition elements
   d lanthanides
   e actinides.

2 **List** the elements that Dobereiner organised into three triads.

3 **State** how many elements were known in:
   a 1789
   b 1808
   c 1829
   d 1864
   e 2011, the year this book was published.

4 **Name** the element that was once called eka-silicon.

Understanding

5 **Outline** how the electron configuration of an element determines its position in the periodic table.

6 **Explain** why Dobereiner organised selected elements into triads.

7 Newland's periodic table ended up having more than one element in some of its boxes. **Explain** why this is not possible.

8 **Explain** why Mendeleev left gaps in his original table.

Applying

9 The electron configurations of different elements are given below. **Identify** in which period and group they should be placed.
   a 2,3
   b 2,8,7
   c 2,8,8,2
   d 2,8,18,6
   e 2,8,18,8,2

10 **Use** the periodic table to **identify** the elements whose electron configurations are listed in Question 9.

11 **Use** the periodic table to help you **predict** the electron configuration for:
   a silicon (Si)
   b helium (He)
   c nitrogen (N)
   d magnesium (Mg).

12 **Identify** the period and group that these atoms would belong to.
   a Ne
   b an atom with atomic number 13
   c an atom with 7 electrons.

13 **Use** the periodic table to **determine** the electron configuration of an atom in:
   a period 2, group 16
   b period 3, group 18.

14 **Use** the periodic table to **determine** how many electrons in an atom:
   a with eight protons
   b with 18 protons
   c with an atomic number of 3
   d with an atomic number of 19
   e in period 2, group 17
   f in period 3, group 2
   g of phosphorus
   h of potassium.

Analysing

15 **Compare** the elements H, Li and Na by listing the similarities and differences in their electron configurations and placement in the periodic table.

16 **Compare** two properties that Mendeleev predicted for eka-silicon with those of germanium.

Evaluating

17 **Propose** reasons why Dalton's symbols for the elements would be difficult to use today.

Creating

18 **Construct** a scale timeline showing important years in the development of the periodic table.

Inquiring

1 Research Mendeleev, Lavoisier or Dalton and construct a biography outlining their lives, achievements and contributions to the development of the periodic table.

2 Find a version of Mendeleev's original periodic table and compare it with the one used today.

3 Find alternative versions of the periodic table that are used today for different purposes.

4.2 Practical activities

1 Investigating a metallic element

Purpose

To make crystals of the metallic element silver.

Materials

- 250 mL beaker
- hotplate or Bunsen burner, tripod, gauze mat and bench mat
- 1 cm × 4 cm clean zinc strip
- 0.3 g silver nitrate (this could be pre-weighed)
- stirring rod
- 0.5 g agar powder (this could be pre-weighed)
- 40 mL distilled or deionised water
- 100 mL measuring cylinder
- Petri dish
- stereomicroscope (optional)

SAFETY

Silver nitrate stains skin, clothing and benchtops and lead nitrate is poisonous.

Wear safety glasses and gloves at all times when dealing with them.

Turn the Bunsen burner to yellow flame or off when not using it.

Tie long hair back.

Procedure

1. Fold a small piece of paper in two and then open it out.
2. Place the paper onto the electronic balance and measure out 0.5 g of agar onto it.
3. Use the measuring cylinder to measure out 40 mL of distilled or deionised water.
4. Pour the water into the beaker and sprinkle the agar into it as shown in step 1 of Figure 4.2.8.
5. Gently warm the beaker over the hotplate or Bunsen burner, stirring until all of the agar is dissolved.
6. Carefully remove the beaker from the hotplate or Bunsen burner and add 0.3 g of silver nitrate to the beaker. Stir until it is dissolved.
7. Pour the agar solution into a petri dish and gently place the zinc strip in the centre as shown in step 2 of Figure 4.2.8.
8. Place the lid on top of the Petri dish and allow the agar to cool in a dark place (perhaps inside a cupboard). The agar should set into a jelly.
9. Inspect the metal crystals that form over the next few days. If available, use a stereomicroscope for a better view.

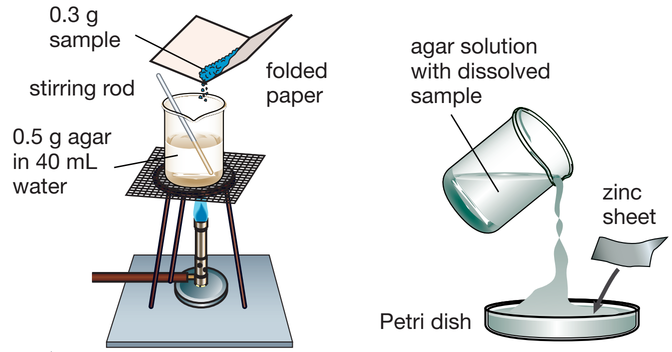


Figure 4.2.8 Making an agar plate

Results

1. Sketch or photograph the pattern produced by the crystals.
2. If you observed the crystals under the stereomicroscope, then draw the shape of an individual crystal.

Discussion

1. **Propose** a reason why the crystals were grown in agar and not a liquid.
2. **Describe** the colour of the agar after the silver nitrate dissolved in it. (This is also what happens if silver nitrate comes into contact with your skin.)
3. **Identify** the atomic number and period for silver.
4. **Identify** which of the following would be the most likely electron configuration of silver.
   A 47
   B 5,11
   C 2,8,18,19
   D 2,8,18,18,1
5. **Name** the special block in which silver is located in the periodic table.

2 Investigating a non-metallic element

Purpose

To determine how much oxygen is in air.

Materials

- non-soapy steel wool (cleaned first with methylated spirits, rinsed in water and dried)
- large (50 mL) test-tube
- plastic container (such as a take-away food container)
- retort stand, bosshead and clamp
- marking pen
- 10 mL measuring cylinder
- access to a calculator

Procedure

1. Add water to the plastic container and place it on the base of the retort stand.
2. Insert a 2–3 cm wad of steel wool in the bottom of the test-tube. Scrunch it up or add a little more so that it stays in place when inverted.
3. Wet the steel wool, then invert the test-tube and clamp so that it as shown in Figure 4.2.9. Make sure that the mouth of the test-tube is well under the surface of the water.
4. Mark where the water level is on the test-tube.
5. Leave the test-tube for at least 2 days. Mark the new water level in the test-tube.
6. Remove the test-tube from the clamp. Leave the steel wool in place and pour water into the test-tube until it reaches the line you marked at the end of the experiment. Empty the water into the measuring cylinder and record its volume in the results table.
7. Pour water into the test-tube until it reaches the line you marked at the start of the experiment. Pour it into an empty measuring cylinder and record its volume (column 2).

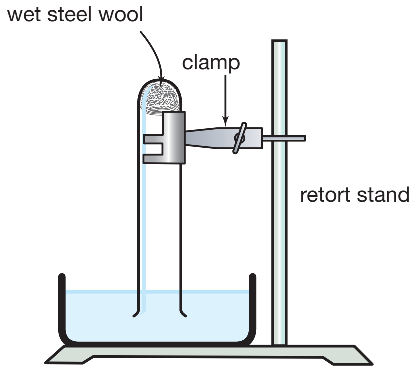


Figure 4.2.9 Inverting the test-tube

Results

1. In your workbook, copy the table below. Record your volumes in columns 1 and 2.
2. Gather the results from another four groups in your class.

Group	Column 1 Volume of air after 2 days (mL)	Column 2 Total volume of air at start (mL)	Column 3 Volume of air 'used up' (mL)	Column 4 Percentage of air 'used up'
1				
2				
3				
4				
5				
Average				

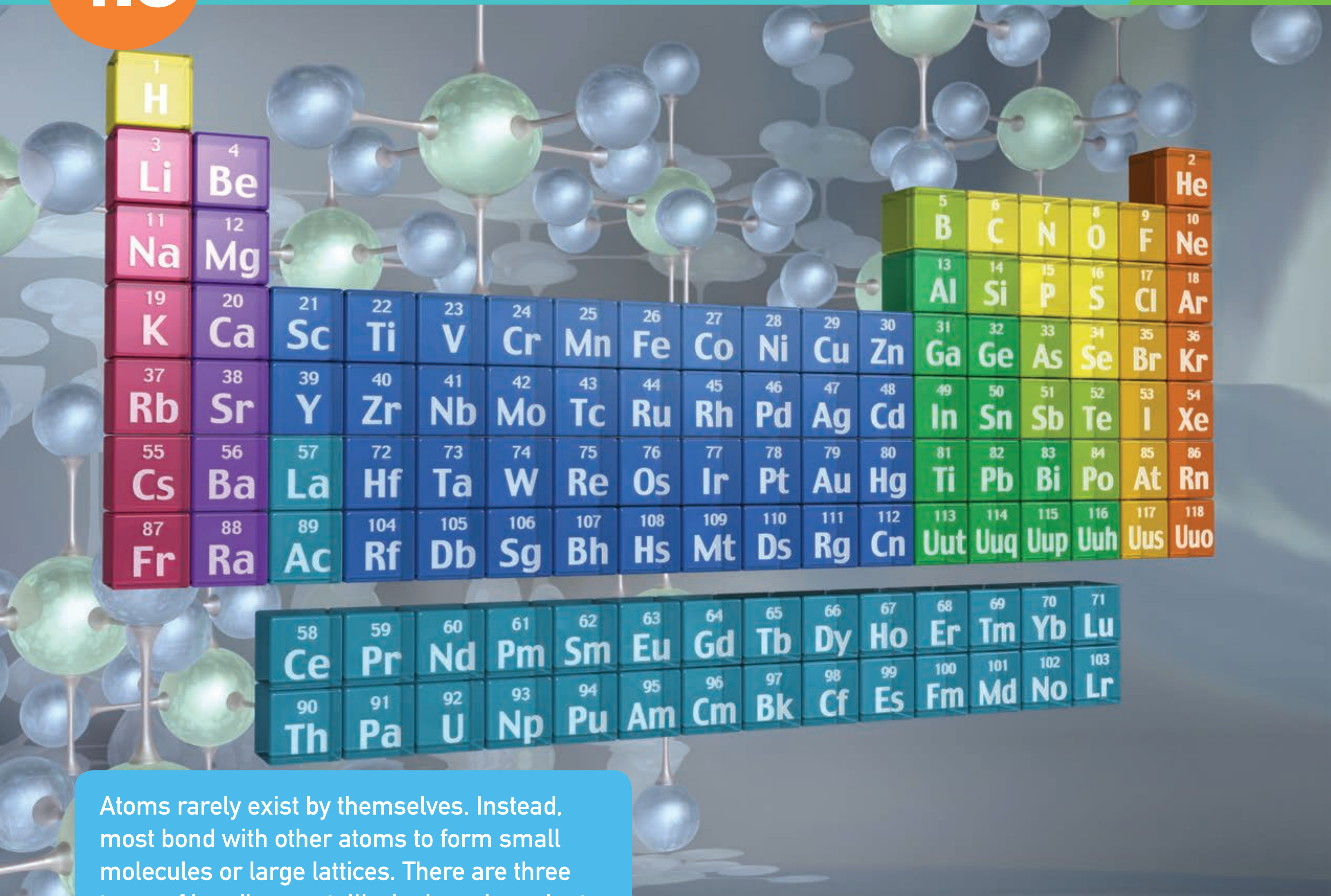
3. Calculate the amount of air 'used up' in the rusting of the steel wool. Calculate this volume by subtracting: column 2 – column 1. Record this volume in column 3.
4. Calculate the percentage of air 'used up' in the experiment.

   $\%\text{ of air used up} = \frac{\text{column 3}}{\text{column 2}} \times 100$
5. Calculate the average of all the groups' measurements.

Discussion

1. **Compare** the results obtained from the different groups.
2. **Identify** the errors in this experiment that will naturally contribute to some variation in the results.
3. **Explain** the advantages of taking multiple measurements in a practical activity.
4. a. Air is about 21% oxygen. **Compare** this percentage with the percentage of air 'used up' in this experiment.
   b. **Assess** whether the two percentages should be the same or not.
   c. **Justify** your answer.
5. For oxygen, **state** its:
   a. atomic number
   b. period and group numbers
   c. electron configuration.

4.3 Bonding



Atoms rarely exist by themselves. Instead, most bond with other atoms to form small molecules or large lattices. There are three types of bonding: metallic, ionic and covalent. The type of bonding that occurs depends on whether the atoms bonding are metals or non-metals.

Atoms that bond and atoms that don't

Most atoms join with other atoms to form small groupings called molecules or large, regular arrangements called lattices. The links between these atoms are called chemical **bonds**.

However, one group of atoms tends not to bond. These are the atoms of elements in group 18, commonly called the **noble** or **inert gases**. The atoms of noble gases are extremely stable and rarely bond with other atoms. Instead, they are **monatomic**, existing as single atoms. This stability is because of their electron configuration. Helium (He) atoms have two electrons filling their outer shells, while neon (Ne) atoms (like those in the tubes of the advertising sign in Figure 4.3.1) have eight electrons filling their outer-shells. Argon (Ar), krypton (Kr), xenon (Xe) and radon (Rn) also have eight electrons in their outer shells.

All other atoms in the periodic table react, gaining, losing or sharing electrons as they do so. This results in particles with full outer shells or outer shells that hold eight electrons. This gives the particles the same electron configuration and stability as a noble gas. This transfer or sharing of electrons is how bonds form.

Bonding types

The types of bonds formed depend on the type of atoms that are bonding.

- Metallic bonding occurs between metal atoms.
- Ionic bonding occurs between metal atoms and non-metal atoms.
- Covalent bonding occurs between atoms of non-metals.

To understand metallic and ionic bonding, you must first understand what ions are.

Ions

Ions are atoms (or groups of atoms) that have become charged because they have had electrons removed from them or because they have removed electrons from other atoms. Atoms are neutral (no charge) because they have equal numbers of protons and electrons. The transfer of electrons destroys this balance.

Number of electrons in an ion ≠ number of protons

This imbalance gives ions a charge.

- Positively charged ions (+) have more protons than electrons. They form when metal atoms lose their outer-shell electrons.
- Negatively charged ions (–) have more electrons than protons. They form when atoms of non-metals gain electrons.

The ions now have the same electron configuration and stability of noble gases.

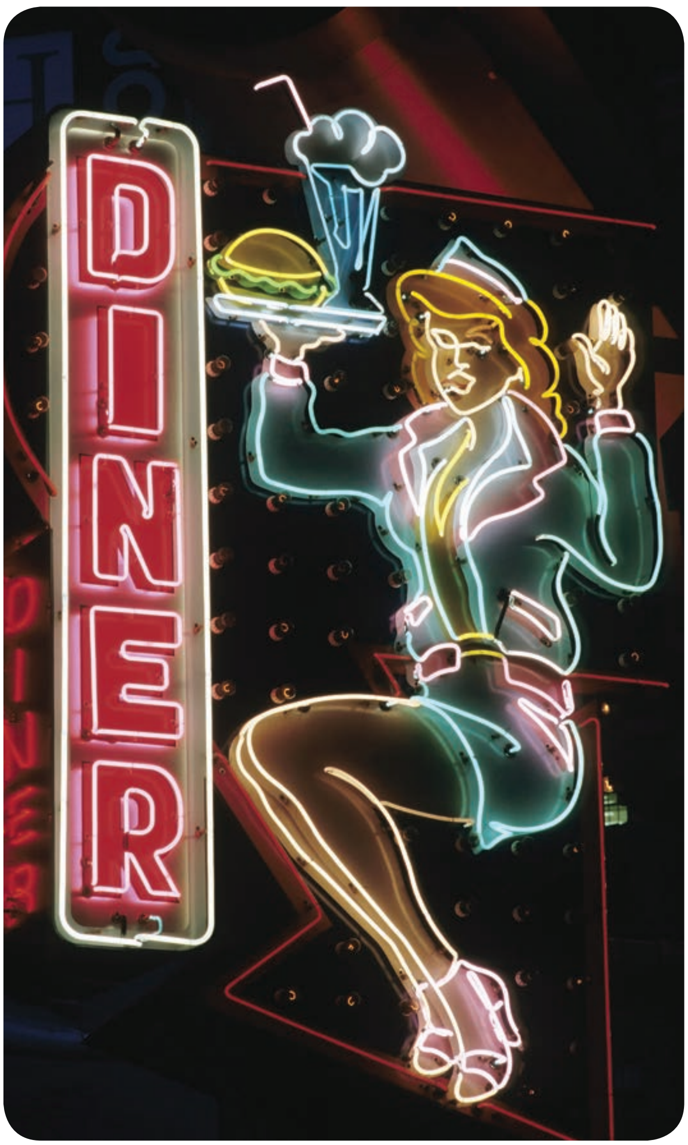


Figure 4.3.1 The neon gas used in this sign is a noble gas. Like all noble gases, neon is stable and so doesn't bond or take part in chemical reactions.

Metallic bonding

Metal atoms have a weak hold on their outer-shell electrons. This gives the outer-shell electrons the freedom to move throughout the metal without being bound to any one atom. Each metal atom becomes a positively charged ion. Opposite charges attract and this electrostatic force provides multidirectional bonding between the positive ions and the 'sea' of loose electrons surrounding them. This bonding holds the metal together and is known as **metallic bonding** (Figure 4.3.2). As Figure 4.3.3 on page 122 shows, metallic bonding explains all the physical properties characteristic of metals.

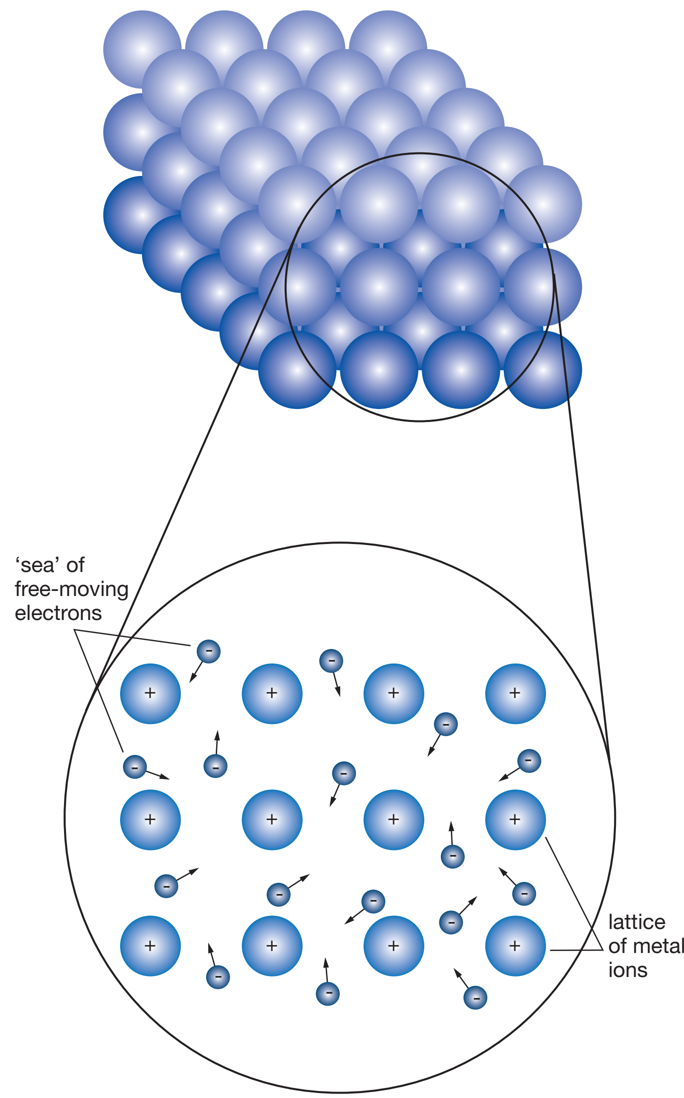


Figure 4.3.2 Metal atoms are bonded to each other because of the attraction between the lattice of positive ions and the electron 'sea' surrounding them.

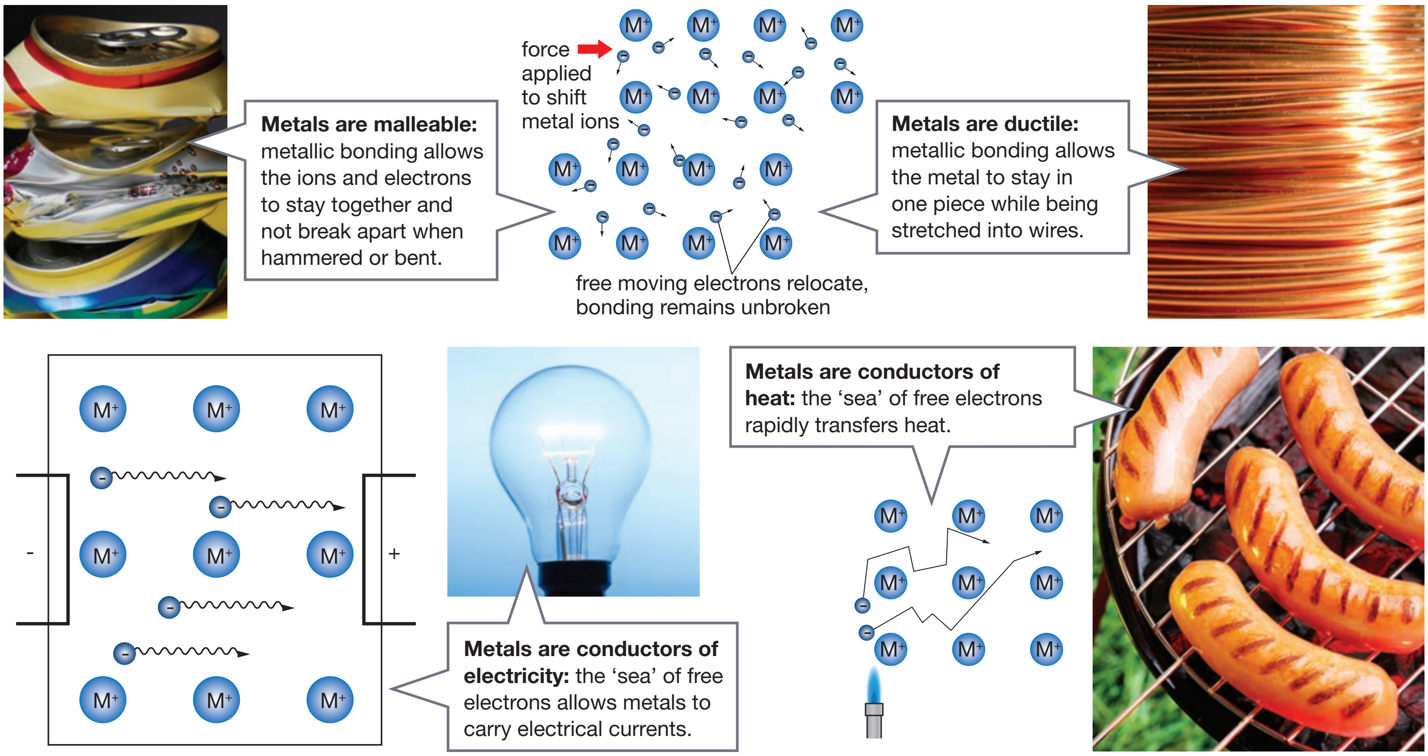


Figure 4.3.3 Each metal ion is attracted to the 'sea' of outer-shell electrons released from all the metal atoms within the lattice. This mutual attraction bonds the metal together.

Ionic bonding

Ionic bonding occurs when metallic elements bond with non-metallic elements. Metal atoms have only a weak hold on their outer-shell electrons. In contrast, non-metallic atoms have a strong hold on their own electrons, and tend to remove outer-shell electrons from any metal atoms nearby. This causes ions to form.

Electrostatic forces pull the positive and negative ions together to form a strong ionic bond. Each ion is surrounded by ions of the opposite charge, building up a three-dimensional structure called a **lattice**. The process is shown in Figure 4.3.4.

The ionic bonding model explains all the important properties of ionic substances, including how they conduct electricity. When solid, ionic substances don't conduct because the ions are bonded within their lattice. When molten or dissolved in water, these ions separate from one another. This allows the ions to conduct an electric current. This current can then light up a globe as in Figure 4.3.5.

Figure 4.3.5 Ionic substances only conduct electricity when molten or dissolved in water.

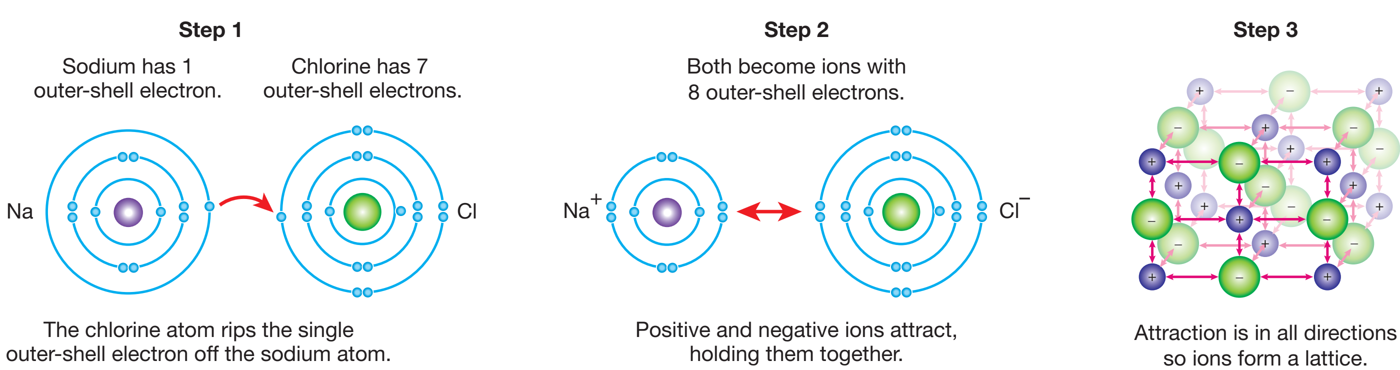


Figure 4.3.4 Ionic bonding holds table salt (sodium chloride, NaCl) together.

Predicting charges of positive ions

Positive ions form when a metal atom loses its outer-shell electrons to obtain the electron configuration of a noble gas. For the metals in groups 1, 2, 13 and 14, their charge is the same as the last digit in the group number.

For example, sodium is in group 1 and so it loses its single outer electron when it becomes an ion. This gives it a single positive charge (+1). Its symbol is Na^{+}. Likewise, calcium is in group 2 and forms Ca^{2+} ions, while tin is in group 14 and forms Sn^{4+} ions. Other examples of metal atoms forming ions are shown in Table 4.3.1.

Table 4.3.1 Predicting charges of metal ions

Element	Group number	Electron configuration	Loses	Charge formed	Ion formed
Lithium Li	1	2,1	1 electron	+1	Lithium ion Li^{+}
Beryllium Be	2	2,2	2 electrons	+2	Beryllium ion Be^{2+}
Sodium Na	1	2,8,1	1 electron	+1	Sodium ion Na^{+}
Magnesium Mg	2	2,8,2	2 electrons	+2	Magnesium ion Mg^{2+}
Aluminium Al	13	2,8,3	3 electrons	+3	Aluminium ion Al^{3+}

The metal atoms of groups 3–12 generally form ions with a charge of +1 or +2. Some, like copper and iron, form multiple charges; for example, Cu^{+} and Cu^{2+}, Fe^{2+} and Fe^{3+}. Unfortunately, these charges cannot be predicted and so you just need to remember them.

Predicting charges of negative ions

Negative ions form when atoms of non-metals remove electrons from metal atoms. The number of electrons they remove is always enough to form an outer shell of 8 electrons.

For example, sulfur is in group 16 and so it gains an additional two electrons, giving it a charge of –2. The ion's new symbol is S^{2-} and is now called sulfide.

Other examples of non-metals forming ions are shown in Table 4.3.2. Figure 4.3.6 summarises the charges formed by various positive and negative ions.

Table 4.3.2 Predicting charges of non-metal ions

Element	Group number	Electron configuration	Gains	Charge formed	Ion formed
Carbon C	14	2,4	4 electrons	–4	Carbide ion C^{4-}
Nitrogen N	15	2,5	3 electrons	–3	Nitride ion N^{3-}
Oxygen O	16	2,6	2 electrons	–2	Oxide ion O^{2-}
Fluorine F	17	2,7	1 electron	–1	Fluoride ion F^{-}
Phosphorus P	15	2,8,5	3 electrons	–3	Phosphide ion P^{3-}
Sulfur S	16	2,8,6	2 electrons	–2	Sulfide ion S^{2-}
Chlorine Cl	17	2,8,7	1 electron	–1	Chloride ion Cl^{-}

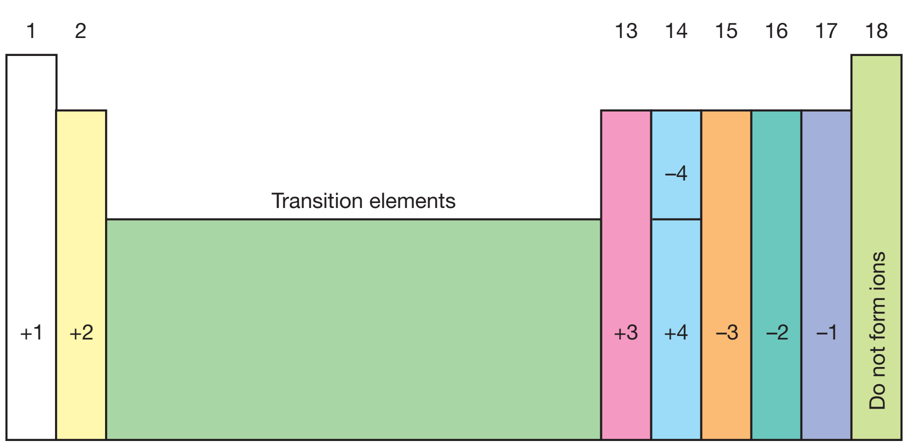


Figure 4.3.6 The group number of an element can be used to predict the charges of any ions formed.

Covalent bonding

Covalent bonding happens when non-metallic atoms bond with each other. Non-metals have the ability to remove electrons from metals but they can't do this to other non-metals. Instead, they share some of their outer-shell electrons. Covalent bonds happen when two non-metals share one or more pairs of outer-shell electrons. If one pair is shared, then one electron from each atom forms the bond. The shared grip on these electrons holds the two atoms together. It's a little like what is happening in Figure 4.3.7.

Figure 4.3.7

Covalent bonding is a bit like two people struggling for the same chair. When both are of the same strength, neither will be able to take the chair off the other. Instead both will continue to hold the chair, indirectly joining them together.

Non-metals only share enough electrons to fill their outer shell or to have eight electrons in it. For example, three additional electrons would fill the outer-shell of nitrogen (electron configuration 2,5). Therefore, a nitrogen atom must pair up three of its electrons with three electrons from other non-metallic atoms. This results in three covalent bonds. Figure 4.3.8 shows that nitrogen can form these three bonds in different ways.

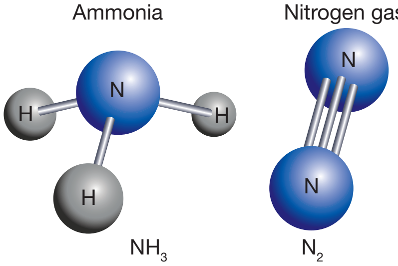


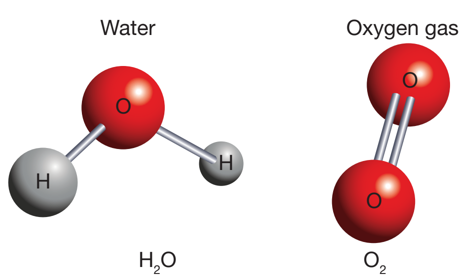
Figure 4.3.8

Nitrogen needs three extra electrons and so forms three single covalent bonds or a triple covalent bond.

Covalent bonding usually results in the formation of discrete groupings of atoms known as **molecules**. Figure 4.3.9 shows just two of the molecules in which oxygen is bonded covalently.

Figure 4.3.9

Oxygen needs two extra electrons and so forms two single covalent bonds or a double covalent bond.



Although carbon forms the basic backbone of millions of different types of molecules, it can also use covalent bonding to form extended lattices of graphite and diamond, shown in Figure 4.3.10.

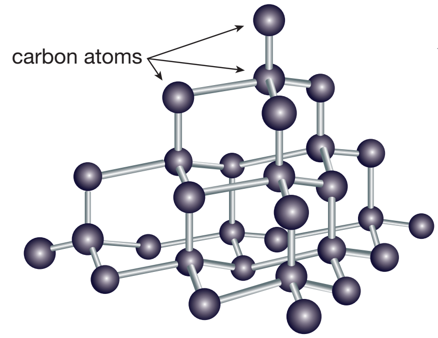


Figure 4.3.10

Carbon's ability to form four covalent bonds allows it to form the incredibly strong structure of diamond.

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INQUIRY science 4 fun

A rough diamond

What does the structure of diamond look like?

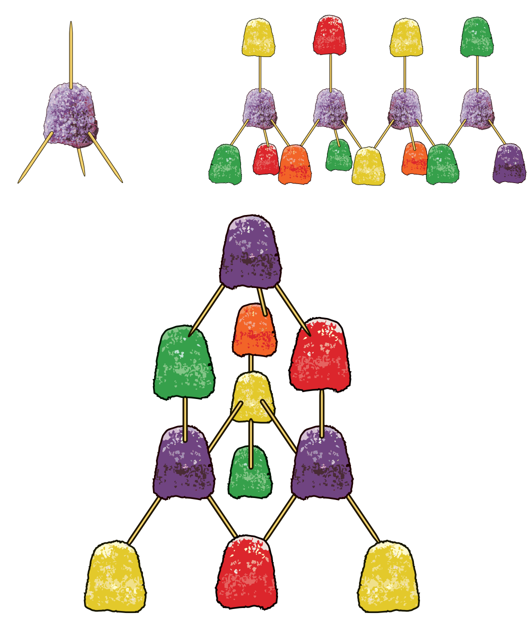
Yes

Collect this …

- lollies such as fruit pastilles or raspberries
- toothpicks
- heavy book

Do this …

1. Insert four toothpicks into a lolly so that they form a tetrahedron or triangular pyramid like that shown below.
2. Attach more lollies and toothpicks until you have diamond!
3. Once built, place the book on top of your structure and push lightly down.



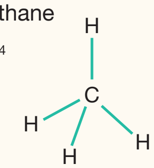
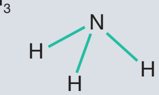
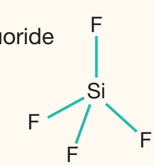
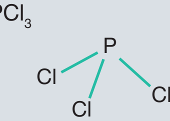
Record this …

Describe how strong your structure was.
Explain why you think this happened.

Predicting the number of covalent bonds

Non-metals form covalent bonds in order to obtain the same electron configuration as a noble gas. Therefore, the number of covalent bonds formed is equal to the number of electrons they need. This is shown in Table 4.3.3.

Table 4.3.3 Predicting number of covalent bonds

Element	Group number	Electron configuration	Number of electrons needed	Number of covalent bonds formed	Example
Carbon C	14	2,4	4	4	Methane CH_4 
Nitrogen N	15	2,5	3	3	Ammonia NH_3 
Oxygen O	16	2,6	2	2	Oxygen gas O_2 
Fluorine F	17	2,7	1	1	Fluorine gas F_2 
Silicon Si	14	2,8,4	4	4	Silicon tetrafluoride SiF_4 
Phosphorus P	15	2,8,5	3	3	Phosphorus trichloride PCl_3 
Sulfur S	16	2,8,6	2	2	Sulfur monoxide SO 
Chlorine Cl	17	2,8,7	1	1	Hydrogen chloride (hydrochloric acid) HCl 

4.3 Unit review

Remembering

1 **State** whether the following elements have a high or low attraction for outer-shell electrons.
- a metals
- b non-metals

2 **Recall** the different types of bonding by matching them with the correct combination of elements.

a	metallic	i	metal/non-metal
b	ionic	ii	non-metal/non-metal
c	covalent	iii	metal/metal

3 **Recall** the different types of bonding by matching them with the term that best identifies them:

a	metallic	i	shared electrons
b	ionic	ii	electron sea
c	covalent	iii	charged atoms

4 Below are eight fragments of sentences. Arrange them to form two complete sentences that **recall** the formation of ions.

Metal atoms ...
Non-metal atoms ...

... gain electrons ...
... lose electrons ...

... to form positive ions, ...
... to form negative ions, ...

... which have more protons than electrons.
... which have more electrons than protons.

5 **Name** the ions formed from the following atoms.
- a sodium
- b chlorine
- c oxygen

6 **Name** two forms of carbon that are lattice structures.

Understanding

7 **Define** the term *monatomic*.

8 **Explain** why the noble gases tend not to form bonds.

9 **Explain** why metals are good electrical conductors.

10 **Define** the terms:
- a malleable
- b ductile.

11 **Outline** how sodium and chlorine atoms eventually form a lattice of sodium chloride

12 **Explain** why molten and dissolved sodium chloride conduct electricity but solid sodium chloride doesn't.

13 **Explain** why sodium chloride is neutral with no overall charge despite it being constructed of charged ions.

14 **Explain** why ionic substances form lattices instead of molecules.

Applying

15 **Use** the metallic bonding model to **explain** why metals don't tend to break when bent.

16 **Use** the structure of diamond to **explain** why it is so strong.

17 **Use** the following electron configurations to **predict** the likely charges of ions formed from the following atoms.
- a Mg (2,8,2)
- b fluorine (2,7)
- c lithium (2,1)
- d phosphorus (2,8,5)

18 **Use** the periodic table to **predict** the likely charges of ions formed by atoms of the following elements.
- a Br
- b Sr
- c Se
- d Fr

19 **Identify** the missing information to complete the following table.

Number of protons	Number of neutrons	Number of electrons	Overall charge	Atom or ion?	Symbol
8	6	10			
10	10	10			
13	15	10			
	18	18	–1		
19	20		+1		K^+

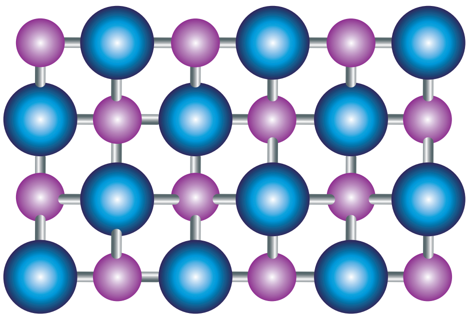
20 **Use** the ionic bonding model to **explain** why ionic substances conduct electricity when molten but not when solid.

21 **Use** the periodic table and group number to **predict** the number of covalent bonds formed by the following non-metals.
- a O
- b N
- c Cl
- d Si

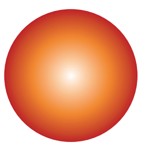
Analysing

22 **Classify** the following diagrams as representing an atom, a molecule or a lattice.

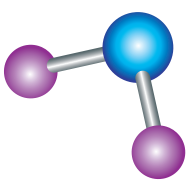
a



b



c

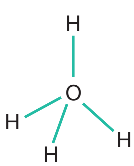


23 Molecules of hydrogen, oxygen and nitrogen all have two atoms in them. Their formulas are H_2, O_2 and N_2. **Compare** the bonding in these molecules by listing their similarities and differences.

Evaluating

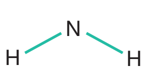
24 **Assess** whether the following molecules are likely or not by counting the number of bonds in them.

a

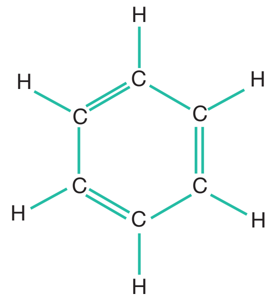


b Br—Br

c



d



e Cl═Cl

25 You don't need to worry about the number of neutrons when calculating the charge of an ion. **Propose** a reason why.

Inquiring

1 Research the noble gases, their uses and their discovery.
2 Within the shells are subshells named s, p, d and f. They explain the existence of the transition elements, lanthanides and actinides. Find diagrams of what these subshells look like.
3 Research the bonding that takes place in graphite and use it to explain why graphite is an electrical conductor.

Figure 4.3.11 These three electrodes are made up of graphite and have just been removed from an electric arc furnace used to make steel.

4.3 Practical activities

1 Model building

Purpose

To construct models of molecules and lattices.

Materials

- chemistry atomic model kit

Procedure

1 Use the model kit to construct lattices of the following. Then test the structure as suggested.

Diamond, C

- Use identical pieces that have four 'pegs' arranged in a tetrahedron.
- Test how rigid the structure is by lightly squashing it.

Graphite, C

- Use identical pieces that have three 'pegs' arranged as a three-point star.
- Test how easy it is to slide one layer over another.

Sodium chloride, NaCl

- Use two colours and pieces that have six 'pegs'.
- Test how rigid the structure is by lightly squashing it.

2 Use the model kit to construct molecules of the following.

Water, H_2O

- For the O atom, use a piece with four 'pegs', but only use two of them.
- For the H atoms, use a piece with one 'peg' only.

Ammonia, NH_3

- For the N atom, use a piece with four 'pegs', but only use three of them.
- For the H atoms, use a piece with one 'peg' only.

Methane, CH_4

- For the C atom, use a piece with four 'pegs'.
- For the H atoms, use a piece with one 'peg' only.

Extension

3 Use the model kit to design and construct different molecules. Make sure you use every peg or hole in every piece you use.

4 Write the chemical formula for each of the molecules you construct.

Discussion

1 **Describe** how the lattice structure of diamond makes it incredibly strong.

2 Use the periodic table to **propose** another element that could be expected to form a similarly shaped and strong lattice as diamond.

3 Atoms bond to obtain the same electron configuration as a noble gas. To do this they form bonds and electron pairs. **State** how many electron pairs should be around most atoms in the molecules you constructed.

4 The pieces chosen to represent hydrogen only had one 'peg' and not the four suggested for other pieces in the molecules. Use electron configuration to **explain** why.

4.4 Family groupings

Look at your family carefully and you'll probably notice some similarities. It might be hair colour, nose or shape of their chins. There will also be many differences. The elements in each group of the periodic table have family resemblances too in the way they look and the way they react.

Similar but different

The elements of any particular group in the periodic table all have the same number of outer-shell electrons. For this reason, they tend to form ions of the same charge or have the same number of covalent bonds when forming a molecule. This causes most compounds constructed from them to be similar.

For example, every element in group 2 is a solid metal. When bonding with a non-metal, each forms an ion carrying a charge of +2. For this reason, they all form similar ionic compounds with chlorine, with the same 1:2 ratio of metal to chlorine. This is shown in their formulas $MgCl_2$, $CaCl_2$, $SrCl_2$ and $BaCl_2$.

Likewise, each of the elements of group 17 form the same number of covalent bonds and therefore form similar molecules, such as HF, HCl, HBr and HI.

However, the elements in each group are not the same. This is because every new period represents a new shell being added to the atom. Therefore, the outer electrons are a little further out from the nucleus and held less tightly. Hence, a gradual change in physical and chemical properties is observed as you move down through the atoms in each group.

Predicting chemical equations

Elements of the same group tend to react in very similar ways and so a balanced chemical equation for one element can be used to predict the reactions for other elements in the group. For example, the balanced formula equation for the reaction of fluorine with hydrogen sulfide is:

$$F_2 + H_2S \rightarrow S + 2HF$$

The other group 17 elements will react in similar ways and so it can be expected that their balanced formula equations will be similar too. Hence, you can predict that their reactions will most likely be:

$$Cl_2 + H_2S \rightarrow S + 2HCl$$
$$Br_2 + H_2S \rightarrow S + 2HBr$$
$$I_2 + H_2S \rightarrow S + 2HI$$

Groups 1 and 2—alkali metals and alkaline earths

The group 1 elements shown in Figure 4.4.1 are known as the **alkali metals**. The group 2 elements are known as the **alkaline earths**. All the alkali metals:

- form +1 ions
- are far too reactive to be found naturally in their pure forms
- have typical metallic properties
- display similar extreme chemical behaviour.

Figure 4.4.1 The alkali metals are group 1 and alkaline earths are group 2.

Lithium, sodium and potassium are light enough to float on water and are so soft that they can be cut with a knife. They all react with chlorine gas (and the other group 17 elements) and produce similar white salts. For example, lithium reacts with chlorine gas to form lithium chloride. This can be written as the balanced formula equation:

$$2Li + Cl_2 \rightarrow 2LiCl$$

From this equation, you can predict the equations for the other alkali metals and other group 17 elements. These equations are:

$$2Na + Cl_2 \rightarrow 2NaCl$$
$$2K + Br_2 \rightarrow 2KBr$$
$$2Li + F_2 \rightarrow 2LiF$$

All alkali metals react violently with water, producing an alkaline or basic solution and hydrogen gas, which sometimes ignites due to the heat produced. This is what has happened in Figure 4.4.2. For sodium, the balanced formula equation is:

$$2Na + 2H_2O \rightarrow 2NaOH + H_2$$

The other alkali metals will do much the same:

$$2Li + 2H_2O \rightarrow 2LiOH + H_2$$
$$2K + 2H_2O \rightarrow 2KOH + H_2$$

Figure 4.4.2 An accident in which sodium was accidentally dropped into water. Reactions of this type become even more violent as you move down group 1.

Trends in chemical activity of metals

Metal atoms have a weak hold on their outer-shell electrons. As you move down a group, extra shells are added and so this hold gets even weaker. This makes the metal more unstable and more reactive. Lithium atoms are the smallest in group 1 and lithium metal fizzes when put in water. Sodium atoms are a little bigger and can react explosively with moisture in the air. Reactions with moisture become even more violent as you move down group 1. To avoid this happening accidentally, alkali metals are usually stored in paraffin oil to keep them moisture-free.

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Group 2 metals all act in a similar, but slightly less reactive, way to group 1.

Group 14

The elements of group 14 display a wide range of properties (Figure 4.4.3). The group begins with the non-metal carbon, moves through the metalloids silicon and germanium and finishes with the metallic elements tin and lead.

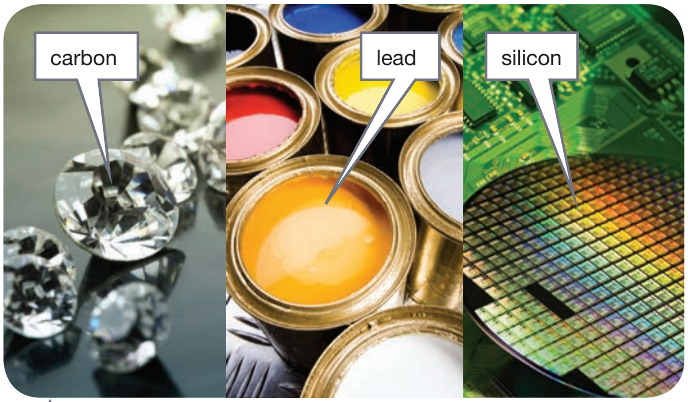

Figure 4.4.3 Group 14 elements

Pure carbon exists in several different forms or **allotropes**, the most common being amorphous carbon (charcoal), diamond, graphite and buckyballs. The molecular structures of the four allotropes of carbon are all different.

Carbon is also in molecules in every living thing on Earth, and anything that was once living such as wood.

A carbon atom forms four covalent bonds when it joins with other carbon atoms or the atoms of other non-metals. This gives carbon the ability to form an amazing range of molecules. Substances that have carbon skeletons like those in Figure 4.4.4 are known as **organic** substances and their molecules are **organic molecules**. Organic molecules make up all living things, fossil fuels and plastics (more correctly known as polymers).

2 p136 4.4 4.5 4.6

Silicon is found as silicon dioxide and metal silicates. Together they make up 75% of the Earth's crust—sand, clay, asbestos and quartz contain silicon as do many gemstones (like the one in Figure 4.4.5). Silicon is the major component of glass.

Mendeleev predicted the existence of germanium 15 years before its discovery, naming it eka-silicon. Germanium is used as the catalyst in fluorescent lights and its oxides are used in the production of lenses for optical instruments such as microscopes. Both silicon and germanium are semiconductors and are widely used in electronic components.

Tin and lead are metals and have characteristic metallic properties. They are malleable (able to be bent), ductile (able to be stretched into wires), lustrous (shiny when polished) and electrical and thermal conductors.

Figure 4.4.5 Silicon dioxide is the main component of many precious and semiprecious gemstones such as this amethyst.

Figure 4.4.4 There are more compounds of carbon than any other element on Earth. Organic compounds form the basis for all life on Earth, all fossil fuels and all plastics.

Group 17—the halogens

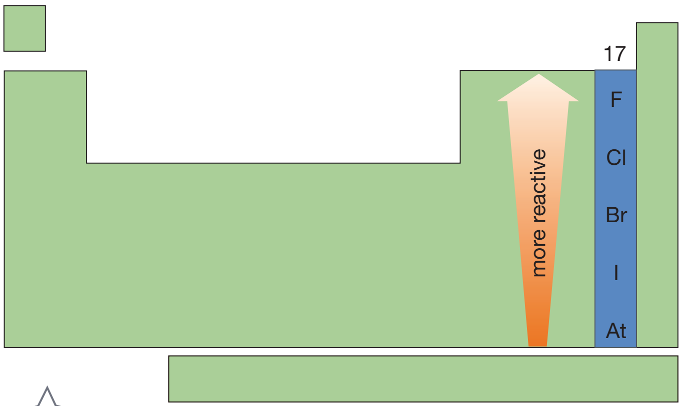

Figure 4.4.6 The halogens

The group 17 elements shown in Figure 4.4.6 are known as the **halogens**. Atoms of halogens:

- form ions with a charge of –1
- are not found in nature in their pure form but are found in various types of salts, including sea salt
- get bigger and become less reactive as you move down the group
- all form molecules, each being made up of two atoms (F_2, Cl_2, Br_2 and I_2). You can see this in Figure 4.4.7
- have coloured and poisonous vapours (Figure 4.4.8).

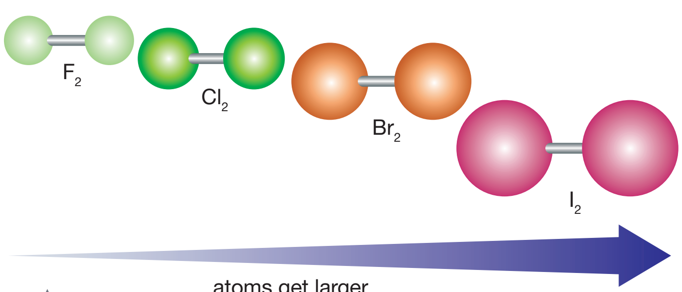

Figure 4.4.7 Although the halogens have similarities, their different sizes result in some differences in their properties.

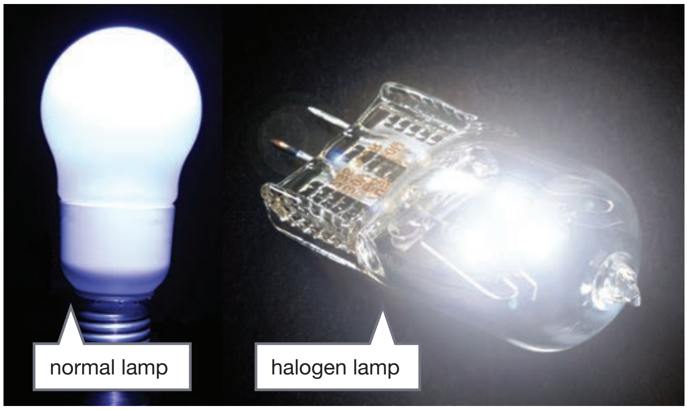

Figure 4.4.8 Halogen lamps are filled with a noble gas with a trace of a halogen such as iodine or bromine. Halogen lights are much brighter than normal globes.

All of the halogens convert hydrogen sulfide (H_2S, known as 'rotten egg' gas) into sulfur (S). These reactions also form very similar gases composed of hydrogen and the halogen, such as hydrogen chloride. These reactions can be written as balanced formula equations:

$$F_2 + H_2S \rightarrow S + 2HF$$

$$Cl_2 + H_2S \rightarrow S + 2HCl$$

As Figure 4.4.9 shows, they also react in a similar way with iron.

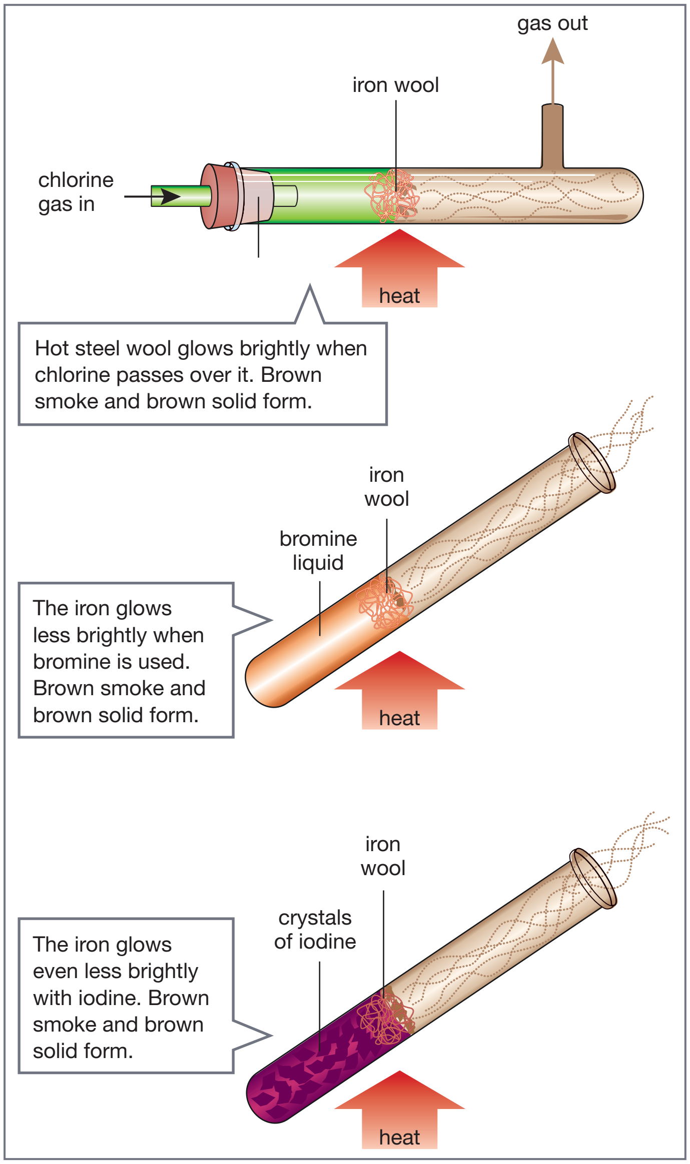

Figure 4.4.9 Each of the halogens reacts with iron in a very similar way, forming very similar products.

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SciFile

Quiet children!

Bromine forms the bromide ion, the salts of which were used as sedatives in the 19th century. Bromide salts were given to some children by providing them with their own saltshaker at dinner time. That kept them quiet!

Group 18—the noble gases

Figure 4.4.10 The noble or inert gases

The noble gases of group 18 (shown in Figure 4.4.10) are colourless and occur naturally in the atmosphere. Distillation is used to separate them from liquid air. They are incredibly stable and react only under rare and extreme circumstances. Helium is safe and light enough to be used for party balloons and for airships (unlike the alternative hydrogen, which is explosive). Balloons of the other noble gases get progressively heavier. This is shown in Figure 4.4.11.

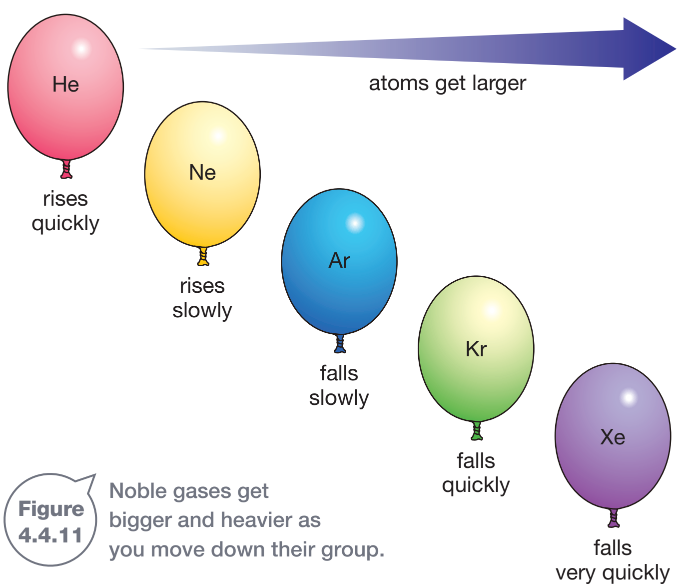


Figure 4.4.11 Noble gases get bigger and heavier as you move down their group.

SciFile

Squeaky voices

Your voice goes high and squeaky when you breathe in helium from a party balloon. Your vocal cords vibrate more quickly in helium because it is lighter than air, making the pitch go higher. Don't try this, though. People have died from breathing in gases such as helium.

The transition metals

The transition elements (groups 3–12) are all metals and include many of our most useful, colourful and valuable ones such as iron, copper, zinc, gold and silver. Their salts and solutions are also colourful—some are shown in Figure 4.4.12. The transition metals have very similar properties. For example, they all tend to be relatively hard with high melting points. Likewise, the period 4 metals iron, cobalt and nickel are all magnetic, despite being in different groups.

SciFile

Scandium, agent Sc-46!

The East German secret police (the Stasi) regularly sprayed opponents of the government (dissidents) with radioactive scandium, $^{46}_{21}Sc$. The unknowing dissidents were then traced with a Geiger counter strapped under the armpits of Stasi agents. Vibrations alerted the agent that their trace was nearby. The Stasi also used radioactive silver bullets that could be safely shot into the tyres of cars they wanted to track.

Figure 4.4.12 These solutions of transition metals show how colourful transition metals can be.

4.4 Unit review

Remembering

1 **State** the group number of the following 'families' of elements.
- **a** alkali metals
- **b** alkaline earth metals
- **c** halogens
- **d** noble gases

2 **Name** the alkali metal that would be the:
- **a** smallest atom
- **b** most reactive.

3 **Name** an element that has similar properties to:
- **a** potassium (K)
- **b** calcium (Ca)
- **c** oxygen (O).

4 **Name** three allotropes of carbon.

5 **State** the chemical formulas for molecules of fluorine and molecules of chlorine.

6 **Name** the separation method that is used to separate noble gases from air.

Understanding

7 **Explain** why helium is used instead of hydrogen in airships.

8 Carbon has a unique ability to form the backbone of many long chain-like molecules. **Describe** the feature of carbon atoms that enables it to do this.

Applying

9 **Use** the electron configuration of oxygen and sulfur to **explain** why they form similar molecules such as H_2O and H_2S.

10 **Use** family resemblances and the balanced formula equations given in this unit to **predict** the reactions of:
- **a** sodium (Na) and water (H_2O)
- **b** rubidium (Rb) with water (H_2O)
- **c** lithium (Li) with iodine (I_2)
- **d** sodium (Na) with bromine (Br_2).

11 **Identify** the chemical formulas missing from each of the following balanced formula equations.
- **a** $F_2 + H_2S \rightarrow \ldots + 2HF$
- **b** $2Na + Br_2 \rightarrow 2\ldots$
- **c** $2Na + 2\ldots \rightarrow 2NaOH + H_2$

Analysing

12 The noble gases and halogens are non-metals yet are very different from each other. **Contrast** these two groups.

13 **Compare** the alkali metals with the alkaline earths by listing their similarities and differences.

Evaluating

14 All the noble gases except radon (which is radioactive) could be used in party balloons but helium is the best. **Propose** a reason why.

15 Tin acts like a non-metal below 13°C. In 1913 Captain Robert Scott and two fellow explorers froze to death in Antarctica after they ran out of heating fuel that was stored in tins. **Use** the properties of metals and non-metals to **propose** reasons why they unexpectedly ran out.

Creating

16 The melting and boiling points for each of the halogens are shown below.
- **a** **Identify** which halogens would be solid, liquid or gas at the following temperatures.
  - **i** 20°C
  - **ii** 100°C
  - **iii** –199°C
  - **iv** 150°C
- **b** **Use** a spreadsheet or graph paper to **construct** accurate line graphs of the:
  - **i** melting point versus period number
  - **ii** boiling point versus period number.

Period number	Group 17 element	Melting point (°C)	Boiling point (°C)
2	Fluorine F	–220	–188
3	Chlorine Cl	–101	–35
4	Bromine Br	–7	59
5	Iodine I	114	185

Inquiring

1 **a** Research the uses for the alkali metals, alkaline earths, halogens and/or noble gases.
b Construct a table to summarise your findings.

2 Some metals such as lead and mercury are known as cumulative poisons. Research what this means, the symptoms of cumulative poisoning and how you might come into contact with these metals.

4.4 Practical activities

1 The alkaline earths

Purpose

To investigate the reactivity of group 2, the alkaline earth elements.

SAFETY

Do not handle calcium metal with your bare hands because it reacts with moisture.

Wear safety glasses at all times.

Materials

- 2 test-tubes and rack
- 1 rubber stopper with hole and glass tubing
- 1 rubber stopper with no hole
- 250 mL beaker
- Bunsen burner, tripod, bench mat and matches
- distilled or de-ionised water
- 1 × 5 cm strip of magnesium
- steel wool or emery paper
- phenolphthalein solution
- small sample of calcium
- tweezers

Procedure

Part A

1. Clean the magnesium strip with steel wool or emery paper and then form it into a coil.
2. Place the coil in a test-tube and cover it with distilled water.
3. Watch *very* carefully over the next 5 minutes. Look for bubbles.
4. If nothing happens, heat it gently over a yellow flame (Figure 4.4.13).
5. When finished, add 1 drop of phenolphthalein to the solution. Record the colour.

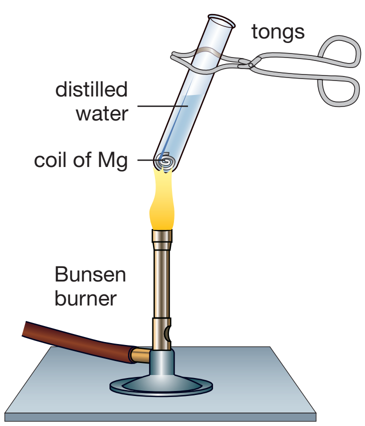


Figure 4.4.13

Part B

1. Put about 5 cm of distilled water into the other test-tube.
2. Using tweezers, add a piece of calcium. Immediately stopper the test-tube with the single-holed rubber stopper and collect any gas generated as shown in Figure 4.4.14.
3. Once the inverted test-tube is filled with gas, remove and stopper it.
4. Light a match, remove the stopper and place the match near the opening of the test-tube filled with gas. Record what happens and use the following guide to identify the gas.
   - If the match pops, then the gas is hydrogen.
   - If the match goes out, then the gas is carbon dioxide.
   - If the match flares up, then the gas is oxygen.
5. Add one drop of phenolphthalein to the test-tube that contained the water and calcium. Record the colour.

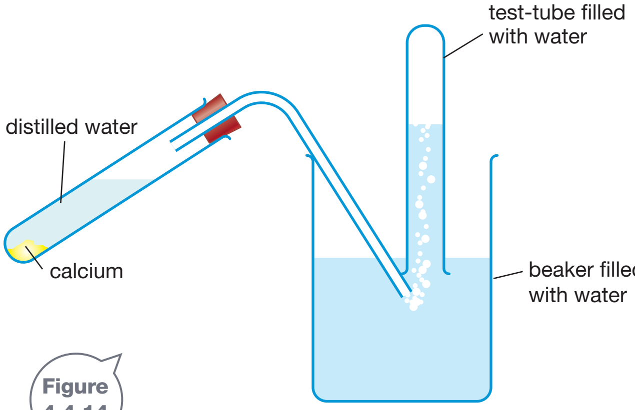


Figure 4.4.14

Discussion

1. From your observation, **state** whether Mg or Ca is more reactive.
2. **Describe** what happens to reactivity as you move down group 2.
3. Group 1 metals are more reactive than group 2 metals. **Propose** a reason why reactions of group 1 metals are rarely performed in class except as teacher demonstrations.

4.4 Practical activities

2 Making casein plastic

Casein was an early plastic that is still used for buttons and some wood glues. It is hardened industrially with formalin.

Purpose

To make a simple plastic and test its properties.

Materials

- full-cream milk
- vinegar
- hotplate or Bunsen burner, bench mat, tripod and gauze mat
- 100 mL measuring cylinder
- 2 × 250 mL beakers
- thermometer
- glass stirring rod
- elastic band
- coarse cloth for straining
- paper towel/filter paper
- assorted moulds (such as bottle caps, moulded chocolate trays)
- fine sandpaper
- tongs

SAFETY

Turn the Bunsen burner to the yellow flame or off when not using it.

Tie long hair back.

Procedure

1 Place 100 mL of milk in one of the beakers. Warm it gently over the hotplate or Bunsen burner until the milk reaches 50°C as shown in Figure 4.4.15. Do not overheat it.

2 Add 10 mL of vinegar and stir it with the stirring rod. Record what happens.

3 The milk should curdle to form white lumps of (casein) and yellowish liquid called whey. Record what it looks and smells like.

4 Secure the piece of cloth tightly over the other beaker and strain through the curds and whey.

5 Carefully remove the cloth and squeeze it to remove as much liquid as you can.

6 Empty the curds (casein) onto the paper towel/filter paper. Pat them dry, then firmly press into moulds. Leave the casein to dry in the sun.

7 After a couple of days, remove the mould and polish the casein with the sandpaper.

8 Use tongs to hold a small amount of the dry casein in a Bunsen flame. Record whether it melts, burns or chars (goes black).

Extension

9 Casein can be used to make a glue that will stick wood together. Search the internet for a procedure that uses casein to make wood glue. Show your teacher your procedure. Once you obtain their permission to proceed, carefully follow your procedure, make your glue and use it to stick two pieces of timber together (for example, two icy-pole sticks).

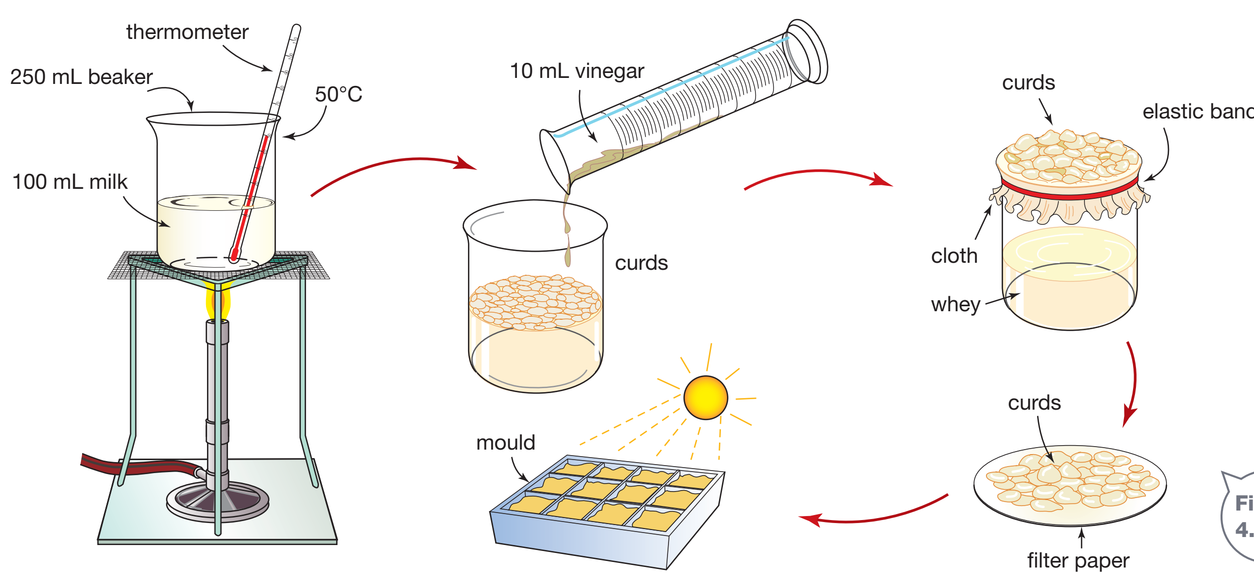


Figure 4.4.15

Results

Record your observations at each step of the practical activity.

Discussion

1 **Propose** a reason for adding the vinegar to the milk in step 2.

2 Plastics can be classified as thermoplastic or thermosetting. Thermoplastics melt when heated while thermosetting plastics burn. **Classify** casein as thermoplastic or thermosetting.

3 Plastics such as casein are organic compounds that contain carbon. **Propose** possible sources of the carbon in this experiment.

3 Halogen precipitates

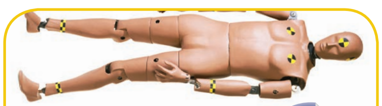
Precipitates are insoluble solids that sometimes form when solutions are mixed.

Purpose

To compare the precipitates formed when lead combines with different halogens.

Materials

- disposable gloves
- white tile or spotting tray (tray with small indents)
- dropping bottles of saturated solutions of lead nitrate, potassium fluoride (KF), potassium chloride (KCl), potassium bromide (KBr) and potassium iodide (KI)
- marker pen



SAFETY

Lead is toxic. Do not get any of its compounds on you.

Wear disposable gloves and safety glasses at all times.

Wash your hands well after the activity.

Procedure

1 Put your disposable gloves and safety glasses on.

2 Place 3 drops of potassium fluoride solution in one indent of the spotting tray or on one corner of the white tile.

3 Use the marker pen to label this sample with KF.

4 In the each of the other indents or corners, place 3 drops of potassium chloride, potassium bromide or potassium iodide solutions. Label them KCl, KBr or KI respectively.

5 Add 3 drops of lead nitrate to each of the samples as shown in Figure 4.4.16. Record your observations.

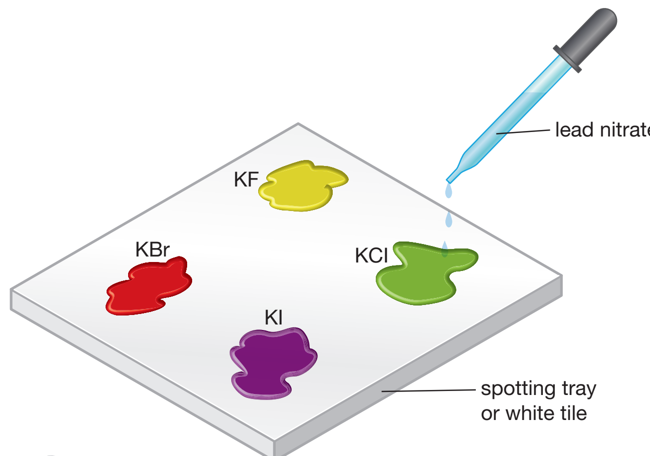


Figure 4.4.16

Results

Construct a table to display your observations.

Discussion

1 The solid formed was a precipitate. **Describe** what a precipitate is.

2 The precipitates formed were lead fluoride, lead chloride, lead bromide and lead iodide. **Compare t**he precipitates formed by listing their similarities and differences.

3 **Describe** the changes in the colour of the precipitates as you moved down group 17 from fluorine to iodine.

4 Chapter review

Remembering

1 **State** how many different elements are currently known.

2 **Name** the following elements.
a F b Ca c Na d Pb

3 **Name** the ions formed from the following atoms.
a sulfur
b aluminium
c phosphorus

4 **List** the symbols for the group 18 elements.

5 **State** the likely charges of the ions that belong in groups 1, 2 and 13–18.

Understanding

6 **Explain** why atoms are neutral, despite containing particles with positive and negative charges.

7 **Define** the following terms.
a malleable b lattice
c allotrope d organic molecule

8 a Helium could be placed in group 2. **Explain** why.
b **Explain** why helium is normally placed in group 18 instead.

9 **Outline** how a sodium ion forms.

10 **Explain** why atoms in the same group have similar properties.

Applying

11 **Identify** the following metals.
a the only metal that is a liquid at 25°C
b those in period 3
c those in group 14

12 **Identify** a non-metallic element that:
a is in group 15
b is in period 2
c has similar properties to chlorine
d would have atoms of a larger diameter than those of oxygen.

13 **Identify** which type of bonding would most likely happen between:
a identical metal atoms
b metal and non-metal atoms
c atoms of non-metals
d Na and O
e N and F
f Mg and Mg.

14 **Use** the group number of the following elements to **predict** the likely charges of the ions formed by:
a sulfur b aluminium
c potassium d nitrogen.

15 **Use** the periodic table to **predict** the number of covalent bonds formed by:
a P b S c Si d Br.

16 Carbon is a group 14 element that forms a compound CH_4. **Use** this information to **predict** the formula of compounds formed from hydrogen and the other four elements of group IV:
a silicon b germanium
c tin d lead.

17 Chlorine reacts with hydrogen sulfide in the reaction:

$$Cl_2 + H_2S \rightarrow S + 2HCl$$

Use this equation to **predict** the balanced formula equation for the reaction of the following halogens with hydrogen sulfide (H_2S).
a bromine (Br_2) b iodine (I_2)

Analysing

18 **Contrast** ionic bonding with covalent bonding.

19 **Analyse** whether the following ions are likely to exist or not.
a Br^{2-} b Sr^{2+} c Se^{3-} d Fr^{-}

Evaluating

20 Plumbing pipes were once made of lead. **Use** the symbols of the periodic table to **propose** where the words 'plumber' and 'plumbing' came from.

Creating

21 **Construct** a simple outline of the periodic table, labelling the direction the groups and periods go, the important 'family' groups and the important other blocks.

22 **Use** the following ten key terms to **construct** a visual summary of the information presented in this chapter.
row
period
group
column
electron configuration
atomic number
protons
electrons
shells
periodic table

4.7

Thinking scientifically

The periodic table below will assist you in answering these questions.

Q1 State which of the following is the correct symbol for sodium.

A So **B** Sm
C Na **D** NA

Q2 State which of the following statements is incorrect.

A Chlorine is in period 3.
B Chlorine is in group 17.
C Chlorine has the atomic number 17.
D Chlorine would react in a similar way to O, N, C and B.

Q3 Every new period represents another electron shell being added to an atom. From this information, identify the statement below that is *definitely* true.

A Helium is the biggest of all the group 18 elements.
B A sulfur atom is bigger than an oxygen atom.
C A sulfur atom is bigger than a phosphorus atom.
D A sodium atom is bigger than a potassium atom.

Q4 Atoms in the same group of a periodic table tend to form similar molecules. Carbon and chlorine form the molecule CCl_4. Use this information and the periodic table to identify which of the following molecules could not occur.

A CN_4 **B** CBr_4
C CF_4 **D** $SiCl_4$

Q5 Atoms in the same group of the periodic table tend to react in a very similar way to each other. The reaction between lithium and chlorine is:

$$2Li + Cl_2 \rightarrow 2LiCl$$

Use this information and the periodic table to identify which of the following reactions could not occur.

A $2Mg + Cl_2 \rightarrow 2MgCl$
B $2Na + Cl_2 \rightarrow 2NaCl$
C $2Na + Br_2 \rightarrow 2NaBr$
D $2K + I_2 \rightarrow 2KI$

Group

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
Period 1	1 **H** hydrogen																	2 **He** helium
Period 2	3 **Li** lithium	4 **Be** beryllium											5 **B** boron	6 **C** carbon	7 **N** nitrogen	8 **O** oxygen	9 **F** fluorine	10 **Ne** neon
Period 3	11 **Na** sodium	12 **Mg** magnesium											13 **Al** aluminium	14 **Si** silicon	15 **P** phosphorus	16 **S** sulfur	17 **Cl** chlorine	18 **Ar** argon
Period 4	19 **K** potassium	20 **Ca** calcium	21 **Sc** scandium	22 **Ti** titanium	23 **V** vanadium	24 **Cr** chromium	25 **Mn** manganese	26 **Fe** iron	27 **Co** cobalt	28 **Ni** nickel	29 **Cu** copper	30 **Zn** zinc	31 **Ga** gallium	32 **Ge** germanium	33 **As** arsenic	34 **Se** selenium	35 **Br** bromine	36 **Kr** krypton
Period 5	37 **Rb** rubidium	38 **Sr** strontium	39 **Y** yttrium	40 **Zr** zirconium	41 **Nb** niobium	42 **Mo** molybdenum	43 **Tc** technetium	44 **Ru** ruthenium	45 **Rh** rhodium	46 **Pd** palladium	47 **Ag** silver	48 **Cd** cadmium	49 **In** indium	50 **Sn** tin	51 **Sb** antimony	52 **Te** tellurium	53 **I** iodine	54 **Xe** xenon
Period 6	55 **Cs** caesium	56 **Ba** barium	57–71 lanthanides	72 **Hf** hafnium	73 **Ta** tantalum	74 **W** tungsten	75 **Re** rhenium	76 **Os** osmium	77 **Ir** iridium	78 **Pt** platinum	79 **Au** gold	80 **Hg** mercury	81 **Tl** thallium	82 **Pb** lead	83 **Bi** bismuth	84 **Po** polonium	85 **At** astatine	86 **Rn** radon
Period 7	87 **Fr** francium	88 **Ra** radium	89–103 actinides	104 **Rf** rutherfordium	105 **Db** dubnium	106 **Sg** seaborgium	107 **Bh** bohrium	108 **Hs** hassium	109 **Mt** meitnerium	110 **Ds** darmstadium	111 **Rg** roentgenium	112 **Cn** copernicium	113 **Uut** ununtrium	114 **Uuq** ununquadium	115 **Uup** ununpentium	116 **Uuh** ununhexium	117 **Uus** ununseptium	118 **Uuo** ununoctium

Lanthanides	57 **La** lanthanum	58 **Ce** cerium	59 **Pr** praseodymium	60 **Nd** neodymium	61 **Pm** promethium	62 **Sm** samarium	63 **Eu** europium	64 **Gd** gadolinium	65 **Tb** trebium	66 **Dy** dysprosium	67 **Ho** holmium	68 **Er** erbium	69 **Tm** thulium	70 **Yb** ytterbium	71 **Lu** lutetium
Actinides	89 **Ac** actinium	90 **Th** thorium	91 **Pa** protactinium	92 **U** uranium	93 **Np** neptunium	94 **Pu** plutonium	95 **Am** americium	96 **Cm** curium	97 **Bk** berkelium	98 **Cf** californium	99 **Es** einsteinium	100 **Fm** fremium	101 **Md** mendelevium	102 **No** nobelium	103 **Lr** lawrencium

Glossary

Unit 4.1

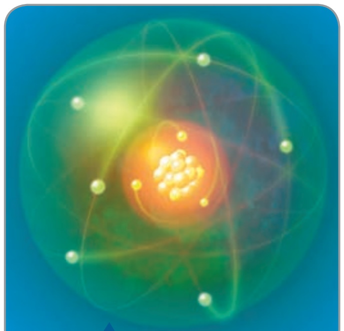
Atomic number: number of protons in an atom

Atoms: the particles that make up all materials; the smallest part of an element that can take part in a chemical reaction

Electron configuration: arrangement of electrons in electron shells

Electron shells: also known as energy levels, the regions surrounding the nucleus where electrons may be found

Element: a substance composed of atoms with the same atomic number; 118 are known to exist



Electron shells

Energy levels: also known as electron shells, the regions surrounding the nucleus where electrons may be found

Ground state: the lowest energy arrangement of an atom's electrons in energy levels (shells)

Indirect evidence: evidence that does not involve direct observation

Nucleus: heavy core at the centre of the atom, made of protons and neutrons

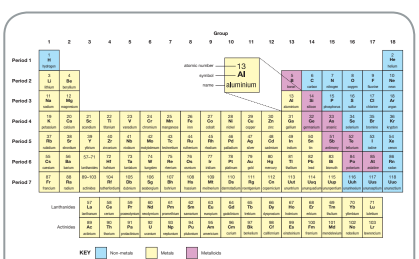
Unit 4.2

Actinides: a special block of metallic elements with atomic numbers 90–103

Groups: vertical columns of the periodic table; group number is equal to the number of electrons in the outer shell of atoms of the elements in that group

Lanthanides: a special block of metallic elements with atomic numbers 58–71

Periodic table: a list of all the known elements, arranged horizontally in order of increasing atomic number and vertically according to the number of electrons in the outer shell



Periodic table

Periods: horizontal rows of the periodic table; period number of an element is equal to the number of electron shells

Transition elements: a special block of metallic elements covering elements from groups 3–12

Unit 4.3

Bonds: links that join atoms together

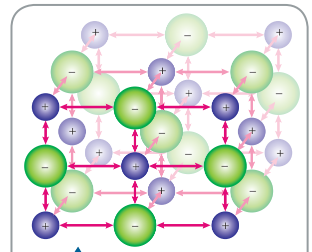
Covalent bonding: the sharing of electrons between atoms of non-metals

Inert gases: also known as noble gases, group 18 elements known for their stability (lack of reactivity)

Ionic bonding: attraction of positive and negative ions formed from the transfer of electrons from metallic to non-metallic atoms

Ions: 'charged atoms' (or groups of atoms) formed by electrons being transferred from one atom to another

Lattices: a regular arrangement of particles. In ionic lattices, the particles are ions; in solid molecular lattices, the particles are molecules; and in diamond and graphite, the particles are atoms



Lattice

Metallic bonding: a mutual attraction between a lattice of positive metal ions and a sea of delocalised, outer-shell electrons

Molecules: discrete groupings of atoms covalently bonded together

Monatomic: atoms that exist on their own, without bonding with others

Noble gases: also known as inert gases, group 18 elements known for their stability (lack of reactivity)

Noble gases

Unit 4.4

Alkali metals: group 1 elements

Alkaline earths: group 2 elements

Allotropes: forms of the same element that have different molecular structures and therefore different properties

Halogens: group 17 elements

Organic: compound that is or was part of a living thing; contains carbon

Organic molecules: molecules that have a backbone of carbon

Allotropes