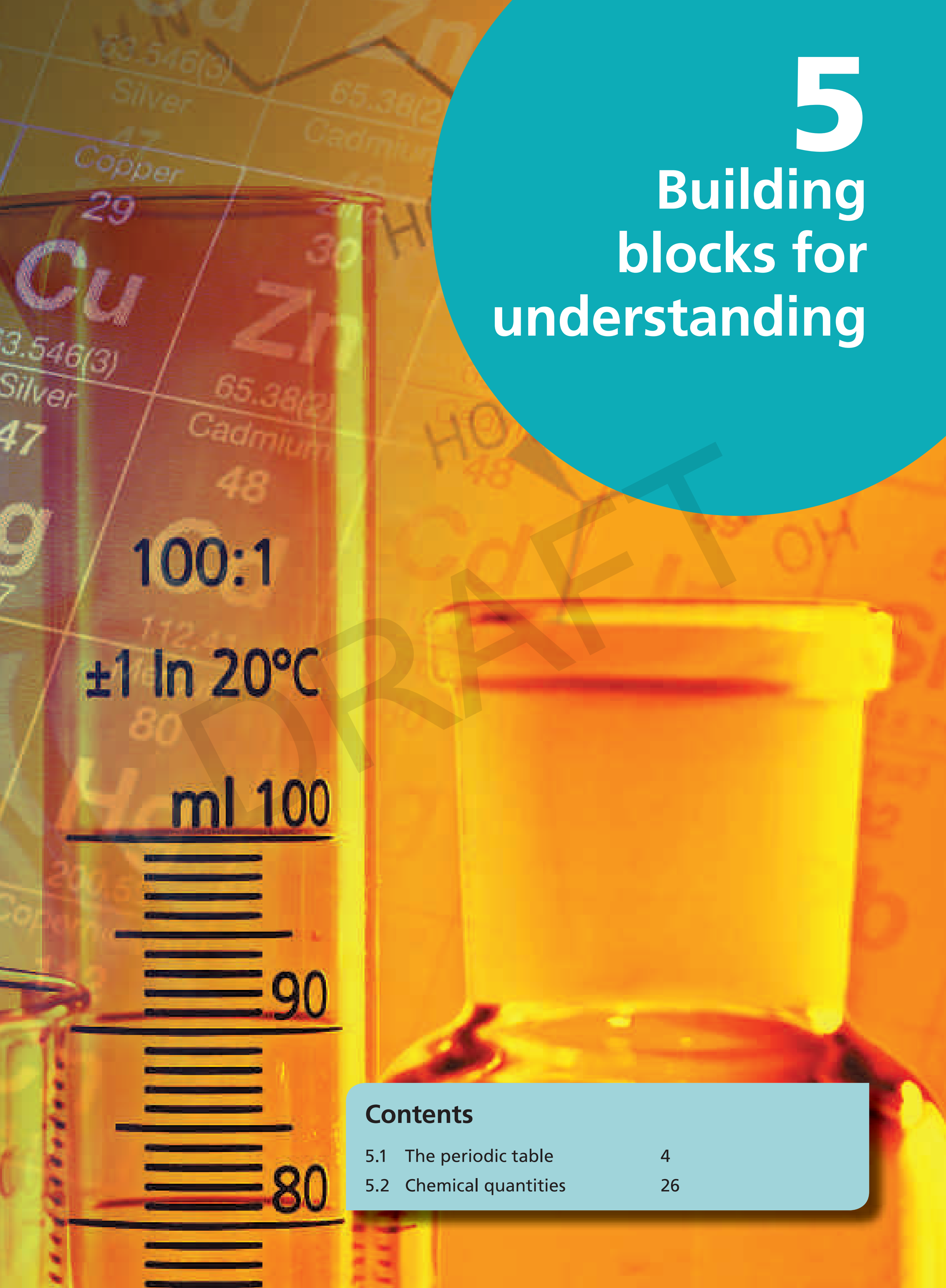
How the ideas in this topic link together

The development of the periodic table is a good example of how scientific theories can change over time. Predictions can be tested, and can demonstrate the limitations of theories. New data can lead to adapting theories, or creating new ones. Mendeleev's periodic table was initially based on known chemical properties and relative atomic masses. As more became known about atomic structure, the early periodic table was modified to take account of the new data. The clear patterns showed how the physical and chemical properties of the elements were related to their atomic structure.

Pattern-seeking is a very important part of working scientifically. The periodic table has many patterns and relationships. Understanding the patterns helps scientists to predict the properties of elements, to understand how they combine to form compounds and to calculate the masses of atoms and molecules. The language chemists use to describe formulae and reactions is used across the world, allowing scientists to communicate their findings and work together to find solutions to problems.

Working Scientifically Focus

- Using models and representational drawings in explanations
- Showing how scientific theories change over time
- Making predictions (of properties or possible reactions) from a scientific model



5 Building blocks for understanding

Contents

THE PERIODIC TABLE

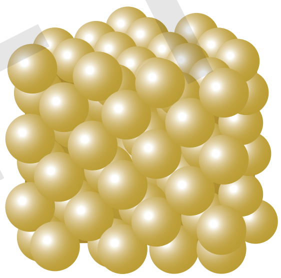
IDEAS YOU HAVE MET BEFORE:

ELEMENTS, MIXTURES AND COMPOUNDS

- Elements cannot be broken down by chemical means.
- Compounds are made from elements chemically combined.
- Chemical symbols are used to show the formulae for elements and compounds.

ATOMS AND ELEMENTS

- A simple model of elements is that they contain building blocks called atoms.
- Different elements are made out of different atoms.
- Elements can be metals, metalloids or non-metals.



THE PERIODIC TABLE

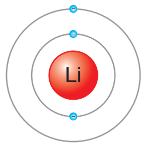
- The periodic table lists elements according to their physical and chemical properties.
- Patterns in properties and reactions can be predicted from the position of different elements in the periodic table.
- The periodic table is arranged into groups and periods (columns and rows).

		H hydrogen 1					He helium 2
Li lithium 3	Be beryllium 4	B boron 5	C carbon 6	N nitrogen 7	O oxygen 8	F fluorine 9	Ne neon 10
Na sodium 11	Mg magnesium 12	Al aluminium 13	Si silicon 14	P phosphorus 15	S sulfur 16	Cl chlorine 17	Ar argon 18
K potassium 19	Ca calcium 20						

5.1

IN THIS CHAPTER YOU WILL FIND OUT ABOUT:

HOW DOES THE MODEL OF THE ATOM HELP US SEE PATTERNS?



- Electrons fill the shells around the nucleus in set patterns.
- Electrons in a shell nearer to the nucleus have a lower energy than electrons in a shell further from the nucleus.
- Electrons can be lost from or gained into the outer shell.

HOW CAN WE EXPLAIN THE ARRANGEMENT OF ELEMENTS IN THE PERIODIC TABLE?

- Mendeleev arranged elements in the first periodic table in order of relative atomic mass.
- In Mendeleev's periodic table, elements with similar properties occur at regular intervals.
- Mendeleev left gaps in his arrangement, predicting that new elements would be discovered whose properties fit the pattern better.
- Scientific models and theories change over time, as new data or observations lead to new theories.
- The elements in the modern periodic table are arranged in order of atomic (proton) number.

WHY DO DIFFERENT GROUPS HAVE DIFFERENT PROPERTIES AND REACTIVITIES?

- Group 0 elements have full outer shells of electrons and do not react.
- Group 1 elements are metals that easily react by losing their one outer electron, to form positive ions.
- Group 7 elements are non-metals that react by gaining an electron to make negative ions.

5.1a Atomic number and the periodic table

Learning objectives:

- explain that the elements in the periodic table are arranged in order of atomic (proton) number
- represent the electronic structure of the first 20 elements
- explain how the electronic structure of atoms follows a pattern.

KEY WORDS

atomic (proton) number
electronic structure
electron shells
energy levels

The electrons of an atom are arranged in patterns. The electrons occupy electron shells around the nucleus in a set order. When a shell is filled, the next electron goes into a shell further from the nucleus with higher energy. These patterns of electrons are the key to the behaviour of atoms.

MAKING LINKS

See topic 1.2 Atomic structure.

The order of elements

All the elements are listed in the periodic table in order of their **atomic number**. The table is laid out so that elements with similar properties are arranged in columns.

Atomic number is the number of protons in an atom. As atoms are neutral, the atomic number also gives the number of electrons in an atom.

For example, the atomic number of hydrogen is 1, carbon is 6 and sodium is 11. This means that hydrogen is the first element in the table, carbon is the sixth and sodium is the eleventh.

		H hydrogen 1					He helium 2
Li lithium 3	Be beryllium 4	B boron 5	C carbon 6	N nitrogen 7	O oxygen 8	F fluorine 9	Ne neon 10
Na sodium 11	Mg magnesium 12	Al aluminium 13	Si silicon 14	P phosphorus 15	S sulfur 16	Cl chlorine 17	Ar argon 18
K potassium 19	Ca calcium 20						

Figure 5.1.1 The elements are listed in order of their atomic number

1 **Which element has the atomic number 13?**

2 **How many protons and electrons are there in a sodium atom?**

The build-up of electrons

Electrons occupy particular **energy levels** called **electron shells** around the nucleus. The lowest energy level shell is nearest the nucleus and takes up to two electrons. This shell is filled first.

Once the first shell is filled, additional electrons go into the second electron shell. The second shell takes up to eight electrons. The next electrons occupy a third shell.

Each of the shells is not a fixed ring, but a three-dimensional zone in which the electrons can orbit.

Oxygen has an atomic number of 8. It has eight protons and so it has eight electrons in the space around the nucleus. The first two electrons go into the first shell. As the first shell is now full, the next six electrons go into the second shell. The **electronic structure** for oxygen is 2,6.

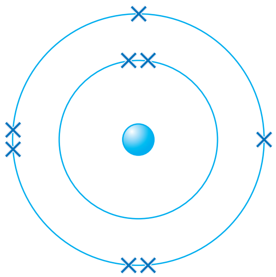


Figure 5.1.2 The electronic structure for oxygen, 2,6. The rings represent the electron shells and the crosses represent electrons. There are 2 electrons on the inner shell and 6 on the outer shell

3 **Draw the electronic structure for hydrogen and for lithium.**

4 **Write down the electronic structure for nitrogen.**

Maximum numbers

The electronic structure of each element in the periodic table can be worked out using:

- the atomic number of the element
- the maximum number of electrons in each shell.

The third shell takes up to eight electrons before the fourth shell starts to fill. Element 19, potassium, has the electronic structure 2,8,8,1.

Use the periodic table to help you to answer these questions.

5 **Work out the electronic structure of argon.**

6 **Use a blank periodic table sheet to draw out the electronic structure of the first 20 elements, putting each structure diagram in the correct box. What do you notice about the number of electrons in the outer shell in each column of the table?**

KEY INFORMATION

Remember:
- the first shell of electrons holds up to 2 electrons
- the second shell holds up to 8 electrons.

DID YOU KNOW?

Sometimes the electrons are represented by dots in diagrams like this, especially if more than one atom is shown. This helps to show which electrons 'belong' to which atom. Electronic structure diagrams are therefore sometimes called 'dot and cross' diagrams.

DID YOU KNOW?

It took many years for scientists to work out the theory of electrons occupying shells. They started by observing the frequencies of light given out by atoms in flame tests and then looked for patterns.

5.1b Electronic structure and groups

Learning objectives:

- explain how the electronic structure of atoms follows a pattern
- recognise that the number of electrons in an element's outer shell corresponds to the element's group number
- explain that elements in a group have similar properties because they have the same number of outer electrons.

KEY WORDS

electronic structure
group

We know that the periodic table is arranged in rows and columns and the elements are written in order of their atomic number. So why are all the elements in the last column all unreactive gases? Why are all the elements in the first column highly reactive metals? The answer lies in the pattern of their electrons.

Electron patterns

As we have seen, electrons occupy energy levels (or shells) in atoms. Each element has a pattern of electrons (known as its **electronic structure**) in its atoms that is built up in a particular order.

The electronic structure shows how many electrons are in each shell and how many shells there are.

- 2,7 is the element fluorine, F. It has two electron shells.
- 2,8,5 is the element phosphorus, P. It has three electron shells.

Adding up the numbers in the electronic structure gives the total number of electrons, which is the same as the atomic number of the element.

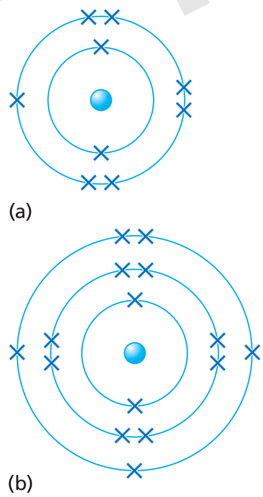


Figure 5.1.3 (a) An atom of fluorine, 2,7; (b) an atom of phosphorus, 2,8,5

1 **An element has the electronic structure 2,8,8,2.**

a **How many electrons do the atoms of the element have?**

b **Use the periodic table to identify the element.**

2 **Work out the electronic structure of the element that has an atomic number 9.**

Electron patterns and groups

There is a very important link between electronic structure and the periodic table. The first row of the periodic table contains the elements hydrogen and helium. These two elements only have electrons in the first electron shell (energy level).

Lithium's third electron goes into the next electron shell. Lithium starts a new row in the periodic table with the electronic structure 2,1.

The third row of the periodic table starts with sodium, Na. We can work out that the electron structure for Na (11 electrons) is 2,8,1. In the same column the fourth row starts with potassium, K, which has 19 electrons. The electron structure for K is 2,8,8,1. All of these elements have one electron in their outside shell. They are all in the first column. This column is known as Group 1.

Let's consider an atom of the element that has the electronic structure of 2,8,2. There are three shells so the element is in the third row of the periodic table. It has two electrons in its outer shell. Its atomic number is 12. (You can work this out by adding up the number of electrons.) Looking at the periodic table the atom is of the element magnesium, 12.

This atom only has two electrons in its outer shell, so it is in the second column. This second column is known as Group 2.

Each column in the periodic table is known as a **group**. For groups 1 to 8, the group number is the same as the number of electrons in the outside shell.

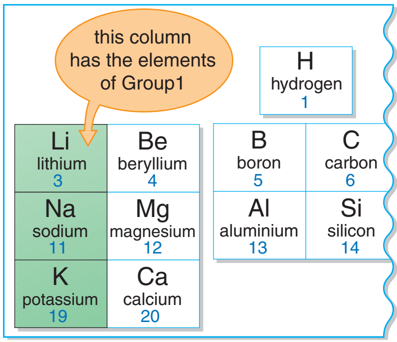


Figure 5.1.4 These elements are in the first column of the periodic table, as they all have one electron in their outer shell

REMEMBER!

The information on this page is true for elements 1–20 in the main groups of the periodic table – the pattern gets more complex after Ca.

3 **What is the pattern of electrons in the atom of the element with atomic number 16? Identify the element and its group.**

4 **To which group does the element chlorine belong?**

Electronic structure and behaviour of elements

Elements in the same group in the periodic table have similar chemical properties.

- The elements in Group 1 are all highly reactive metals
- The elements in Group 7 (the halogens) all react with Group 1 metals to make salts
- The elements in Group 0 are all unreactive (or noble) gases.

This pattern is because all the elements in a group have the same number of electrons in their outer shells.

Use the periodic table at the back of this book to help you to answer these questions.

5 **What are the electronic structures of fluorine and chlorine? Why are they both found in Group 7?**

6 **Identify the elements in Group 6.**

DID YOU KNOW?

The rows in the periodic table are called *periods*. For elements 1–20, each period number is the same as the number of electron shells. Each period contains the elements whose outside shell of electrons is 'filling up'.

5.1c Mendeleev and the periodic table

Learning objectives:

- describe the steps in the development of the periodic table
- explain how Mendeleev left spaces for undiscovered elements
- explain why the element order in the modern periodic table was changed
- explain how testing a prediction can support or refute a new scientific idea.

KEY WORDS

isotope
periodic table
predictions
properties
relative atomic mass

Scientists noticed that some elements had similar properties. They tried to find a way to group them, and began to assemble what is now the periodic table. It was like trying to put together a jigsaw puzzle, but without having all the pieces.

The timeline of ideas

Some elements have been known since ancient times, such as gold, silver, iron and copper. More were discovered over time.

By 1829 a German scientist called Döbereiner had noticed that sometimes three elements had similar **properties**. He noticed patterns with these groups of elements:

lithium sodium potassium	calcium strontium barium	chlorine bromine iodine

Figure 5.1.5 Döbereiner

Döbereiner noticed patterns in the properties of elements but did not make **predictions** of chemical properties.

1 **Find the groups of three elements in the periodic table. What do you notice?**

Allowing for predictions

As we have seen, different attempts had been made to order the elements according to their **relative atomic mass** (the mass shown in the modern **periodic table**). Arranging them in this way meant that some elements seemed to be in the wrong position. In 1869 a Russian scientist, Dimitri Mendeleev, arranged them in a new way. He:

- decided to swap some elements round so that the patterns of chemical behaviour fitted better
- was able to imagine that there were undiscovered elements

- left gaps in his **periodic** table for later discoveries and used the patterns of chemical behaviour to decide where to leave these gaps.

He gave special names to these unknown elements in his gaps. He took the name of the element above the gap in that group and put the prefix 'eka' in front of the name. One unknown was eka-silicon (beyond silicon). This element was discovered in 1886 and was named germanium. Altogether three of these new elements were discovered within Mendeleev's lifetime.

Figure 5.1.6 Mendeleev's periodic table from 1869 was laid out differently from the modern table. He left gaps for 'missing' elements

2 Show that in any of Döbereiner's groups of three elements the relative atomic mass of the middle element is the mean of the values for the other two elements.

3 Explain why some elements appeared to be in the wrong order.

4 Mendeleev left gaps and made predictions for undiscovered elements in the his periodic table. Explain how the discovery of germanium supported this approach.

KEY INFORMATION

If germanium had not fitted into Mendeleev's pattern then the evidence would have *refuted* his predictions, but instead, it did fit his pattern so the new evidence *supported* his predictions.

Discovering the unpredictable

Mendeleev's theory was first proposed in 1869. Before he died in 1907, many new elements had been found to fit into his table, but there were some unanswered questions. In 1913 it was discovered that the order of elements in Mendeleev's periodic table matched the positive charge on the nucleus, the *atomic number*.

In 1932 the discovery of the neutron explained why the order of elements by atomic number did not always match the order by relative atomic mass – the variable numbers of neutrons meant some elements were more massive than their neighbour with a smaller atomic number. The discovery that most elements occurred as **isotopes**, with the same atomic number and identical chemical properties but different masses, confirmed that the elements should be arranged in order of atomic number. So Mendeleev's arrangement was supported by later evidence, which finally was able to explain why the order was correct.

DID YOU KNOW?

Mendeleev did not win a Nobel Prize, as the first Nobel prize for Chemistry was awarded in 1901; however, he does have an element named after him. Look up which number this is.

5 Argon was discovered in 1894. One possibility was to put argon in the same group as lithium and sodium.

a Look at the modern periodic table and use the data to explain why this was a possibility.

b What is wrong with this arrangement?

c Explain why this is not a problem in the modern periodic table.

6 A student stated that 'The modern periodic table is complete'. Briefly explain whether you agree with the student. Use books or the internet to help you with this.

MAKING LINKS

See topic 1.2 Atomic structure.

5.1d Metals and non-metals

Learning objectives:

- identify where metals and non-metals are located on the periodic table
- explain the differences between metals and non-metals based on their physical and chemical properties
- explain that metals react by losing electrons to form positive ions and more reactive non-metals react by gaining electrons to form negative ions.

KEY WORDS

atomic structure
ions

Whether an element is a metal or a non-metal depends on the electronic structure of its atoms. Metals are found on the left-hand side and to the bottom of the periodic table and the non-metals on the right and to the top.

Metals and non-metals in the periodic table

Looking at the periodic table you can see the metals lithium, sodium, magnesium and calcium on the left-hand side.

You can see the non-metals nitrogen, oxygen, sulfur, phosphorus, chlorine and neon on the right-hand side.

From your knowledge of chemistry so far try to find the line of elements that separates the metals from the non-metals. You should notice that the majority of the elements in the whole table are metals.

DID YOU KNOW?

The elements separating the metals from the non-metals is called the metalloids.

Group 1	Group 2											Group 3	4	5	6	7	0
			H hydrogen 1														He helium 2
Li lithium 3	Be beryllium 4											B boron 5	C carbon 6	N nitrogen 7	O oxygen 8	F fluorine 9	Ne neon 10
Na sodium 11	Mg beryllium 12											Al aluminium 13	Si silicon 14	P phosphorus 15	S sulfur 16	Cl chlorine 17	Ar argon 18
K potassium 19	Ca calcium 20	Sc scandium 21	Ti titanium 22	V vanadium 23	Cr chromium 24	Mn manganese 25	Fe iron 26	Co cobalt 27	Ni nickel 28	Cu copper 29	Zn zinc 30	Ga gallium 31	Ge germanium 32	As arsenic 33	Se selenium 34	Br bromine 35	Kr krypton 36
Rb rubidium 37	Sr strontium 38	Y yttrium 39	Zr zirconium 40	Nb niobium 41	Mo molybdenum 42	Tc technetium 43	Ru ruthenium 44	Rh rhodium 45	Pd palladium 46	Ag silver 47	Cd cadmium 48	In indium 49	Sn tin 50	Sb antimony 51	Te tellurium 52	I iodine 53	Xe xenon 54

Figure 5.1.7 Top section of the periodic table

1. **Find the element with atomic number 53. Is this element a metal or a non-metal?**
2. **Is the element with atomic number 26 a metal or a non-metal?**

Metals and non-metals and atomic structure

As we have seen, there is a link between an element's **atomic structure** and its position in the periodic table. So atoms of Group 1 elements, for example, all have just one electron in the outer shell, and Group 2 atoms all have two, and so on.

The metal elements are on the left side of the periodic table, so their atoms all have small numbers of outer electrons. The non-metals are in the groups on the right-hand side of the table - their atoms have outer shells containing more electrons.

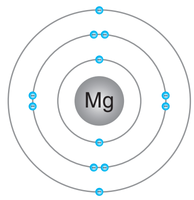


Figure 5.1.8 A magnesium (Mg) atom with two electrons in the outer shell

3. **How many electrons does an uncharged atom of aluminium have in its outer shell?**
4. **Name an element whose uncharged atoms have five electrons in the outer shell.**

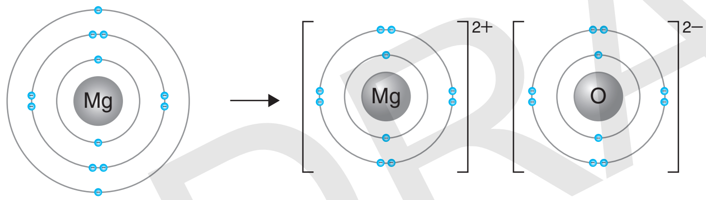
KEY INFORMATION

Atomic structure refers to the whole atom - protons and neutrons in the nucleus as well as the orbiting electrons. Electronic structure describes just the arrangement of electrons in energy levels or 'shells' around the nucleus.

Metal and non-metal ions

When an uncharged atom loses or gains electrons it becomes a charged atom, or **ion**.

For example, magnesium is a metal, in Group 2 of the periodic table. When an uncharged metal atom like magnesium loses its two outer electrons, it loses two negative charges, but still has the same number of positive charges in its protons, so it becomes a positive ion.



Magnesium atom, Mg

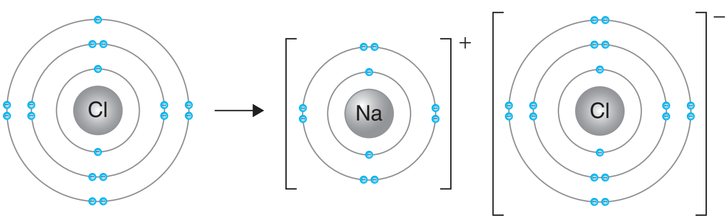
Magnesium ion, Mg^{2+}

Figure 5.1.9 When a metal, such as magnesium, reacts with oxygen, the metal loses electrons and oxygen gains electrons

KEY INFORMATION

Metals make positive ions and non-metals don't make positive ions. An exception to this rule is hydrogen, which is a non-metal that forms positive ions.

Chlorine is a non-metal, in Group 7 of the periodic table. When an uncharged non-metal atom like chlorine gains an outer electron, it has one extra negative charge, but still has the same number of positive charges in its protons, so it becomes a negative ion.



Chlorine atom, Cl

Chlorine ion, Cl^-

Figure 5.1.10 When a non-metal, such as chlorine, reacts with a metal, such as sodium, the non-metal gains an electron from the metal

Elements that form positive ions are metals. Elements that do not form positive ions are non-metals. Some reactive non-metals form negative ions.

5. **Would the element with atomic number 8 form a positive or a negative ion? Explain your answer.**
6. **Write the symbol for the ion of the element with atomic number 3. (Hint: write the element symbol with the charge as a superscript, as in Figure 5.1.9.)**

KEY CONCEPT

5.1e Atoms into ions

Learning objectives:

- recognise the difference between atoms and ions
- explain why metal ions carry a positive charge and non-metal ions carry a negative charge
- represent electronic structure of atoms and ions in diagrams
- explain how the reactions of elements are related to their electronic structure.

KEY WORDS

ion
reactive
electrostatic attraction

The noble gases are all very unreactive. Their atoms are very stable and do not react with other atoms. This is because their outer electron shells are full. Atoms of other elements react by losing or gaining outer electrons.

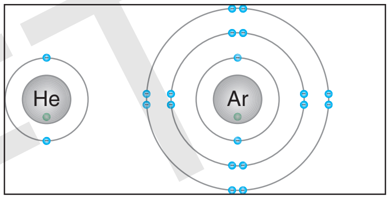


Figure 5.1.11 Atoms of the noble gases helium and argon each have a full outer electron shell. (The inner-most electron shell holds a maximum of two electrons)

Full outer electron shells

Atoms of almost all noble gases have eight electrons in the outer shell. Eight is the maximum number, and the atoms of noble gases are very stable and unreactive.

Atoms of other elements are less stable because they have an unfilled outer electron shell. The atoms react by losing or gaining electrons until the outer shell is full.

The chemical reactions of elements depend on the number of electrons in the outer shell of their atoms.

1. **Apart from argon, name two elements whose atoms have eight electrons in the outer shell.**
2. **The atoms of one noble gas do not have eight electrons in the outer shell, but the outer shell is still full. Which element is this?**

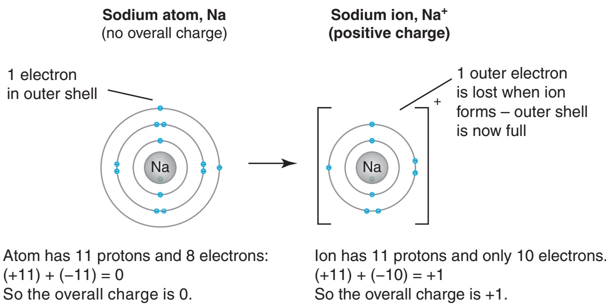


Figure 5.1.12 As for all metals, sodium atoms form positively charged ions

Forming ions

Atoms of most metals have only one, two or three electrons in their outer shell. These atoms react by losing their outer electrons. When this happens, the old outer shell effectively disappears. The next shell down is now the outer shell, and it is full. The 'atom' now has fewer negative charges because it has lost electrons, but it still has the same number of positively charged protons in the nucleus. So overall, it has a positive charge. We call it a positively charged **ion**.

Non-metal atoms typically have four or more electrons in the outer shell. They do not form positive ions. Atoms of more reactive non-metals, like oxygen or chlorine, react by gaining electrons and filling the outer electron shell. The 'atom' now has more negative charges because of the extra electrons, but still has the same number of positively charged protons, so it has become a negatively charged ion.

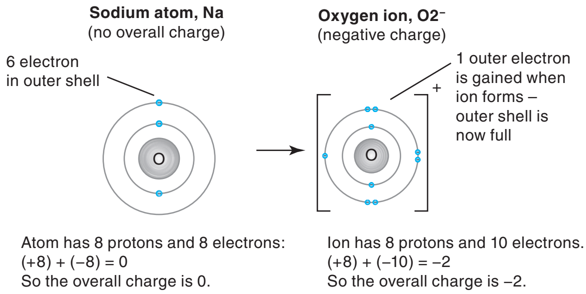


Figure 5.1.13 When non-metal atoms like oxygen form ions, they are negatively charged

3 **How many 'spaces' does an oxygen atom have in the outer shell, which are filled when it forms as ion?**

4 **Predict the charge on a fluoride ion, using your knowledge of fluorine's electronic structure.**

Reactivity and ions

Some elements are more **reactive** than others. In more reactive atoms, it is easier for electrons to be lost or gained in the outer shell.

Potassium and lithium atoms both lose one electron to become positive ions. The **electrostatic attraction** between the negative outer electron and the positive nucleus is weaker in a potassium atom than in a lithium atom, because the outer electron is further from the nucleus, and is 'shielded' by more inner shells. The electron is more easily lost from a potassium atom in reactions, so potassium is more reactive than lithium.

Fluorine and bromine atoms both gain one electron when they become negative ions. The electrostatic attraction between the incoming electron (–) and the nucleus (+) is stronger in a fluorine atom than in a bromine atom, because the outer shell is closer to the nucleus, and is 'shielded' by fewer inner shells. The electron is more easily gained by a fluorine atom in reactions, so fluorine is more reactive than bromine.

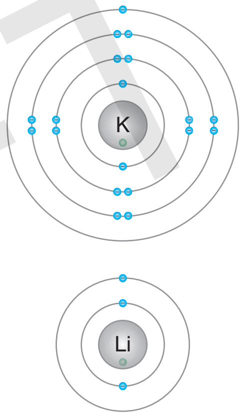


Figure 5.1.14 The outer electron is less strongly held in a potassium atom than in a lithium atom, so potassium atoms are more reactive

5 **Explain why sodium is more reactive than lithium, using ideas about electrostatic attraction.**

6 **Is chlorine likely to be more reactive than fluorine, or less reactive? Explain your answer.**

REMEMBER!

The atoms get larger as you go down each group in the periodic table. So a bromine atom is larger, with more electron shells, than a chlorine or a fluorine atom. The larger the atom, the weaker the electrostatic attraction between the nucleus and the outer electrons.

5.1f Exploring Group 0

Learning objectives:

- describe the unreactivity of the noble gases
- predict the properties of noble gases from trends down the group
- explain how properties of the elements in Group 0 depend on the outer shell of electrons of their atoms.

KEY WORDS

unreactive
trend

The elements in Group 0, the noble gases, all have filled outer electron shells. This make them all very unreactive, so their chemical properties are similar. But their physical properties are different because they have different relative atomic masses.

Figure 5.1.20 Party balloons and weather balloons are filled with helium. Hydrogen is less dense than helium, so would 'float' better, but hydrogen is highly flammable whereas helium is unreactive

Patterns in Group 0

If you find Group 0 on the periodic table, you will see these **elements** in order.

Group 0
Helium (He)
Neon (Ne)
Argon (Ar)
Krypton (Kr)
Xenon (Xe)

All the elements of Group 0 in the periodic table have two things in common.

- They are all **unreactive**.
- They are all gases at room temperature.

The boiling points of the elements in Group 0 show a trend.

Helium has the lowest boiling point. This means that helium remains as a gas at a lower temperature than the other noble gases.

The **trend** is that the boiling points of the gases increase down the group.

Boiling points (°C)	
He	–268
Ne	–246
Ar	–186
Kr	–153
Xe	–108

EXTENSION

Mendeleev did not know of the existence of these noble gases. The gases listed here were discovered much later. If the noble gases had not fitted into Mendeleev's pattern then the evidence would have *refuted* his ideas but instead they did fit his pattern so the new evidence *supported* his ideas.

Link: 5.1f ⟶ 5.1g

1. **Find the relative atomic masses of the Group 0 elements. What do you notice about them, in relation to each element's position in the group?**
2. **There is a noble gas with a bigger relative atomic mass than Xe. Predict its boiling point.**

Why does helium stay as a gas at lower temperatures?

Group 0 elements exist as single atoms, rather than combining with other atoms in molecules or larger structures. The atoms have only weak attractive forces between them, so they move around freely as a gas. All the elements in Group 0 are gases at room temperature.

The attractive forces between the atoms increase with the size of the atom. So elements with larger atoms need more energy to overcome the forces when the element boils and becomes a gas. This is why helium, which has the smallest atoms and the lowest atomic mass of all the noble gases, has the lowest boiling point. The boiling points of the noble gases increase as you go down the group.

3. **Describe how boiling point varies with relative atomic mass for Group 0 elements.**

Why do elements in Group 0 exist as single atoms?

The elements of Group 0 do not make compounds with other elements. Group 0 atoms have eight electrons in their outer shell (except for helium, which has only two electrons). This is a very stable configuration, so there is no electron movement, which makes Group 0 atoms very unreactive.

4. **Draw the electronic structure for helium. Explain why this atom does not join with any other atom.**
5. **Write out the electronic structure (in numbers) for argon. Explain why argon is unreactive but has a higher boiling point than helium.**
6. **One of the first noble gas compounds to be made was xenon tetrafluoride (XeF_4), made in 1962. It is a stable crystalline solid at room temperature.**
   - **a Complete and balance the following equation:**

     __ Xe + __ F_2 → ____

   - **b Suggest why it was a surprise that a noble gas compound had been made.**

DID YOU KNOW?

Helium is also used with neon in the lasers that scan supermarket barcodes. Street lights used in long tunnels contain both neon and sodium. Argon is used for specialist welding and to fill the space between double-glazed windows. Find out what krypton and xenon are used for.

Figure 5.1.21 Neon lights were first used for advertising signs in 1912

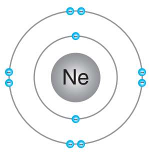


Figure 5.1.22 The electronic structure of neon. The outer shell is full

KEY INFORMATION

The noble gases are very unreactive compared to other elements, but they are not all completely inert. Under certain circumstances, some can react to form compounds, like xenon tetrafluoride.

5.1g Exploring Group 1

Learning objectives:

- explain why Group 1 metals are known as the alkali metals
- predict the properties of Group 1 metals from trends down the group
- relate the properties of the alkali metals to the number of electrons in their outer shell.

KEY WORDS

alkali
hydroxide
ion
oxide
reactivity
stable electronic structure

The outer electron structure determines the physical and chemical properties of an element. The elements in Group 1 all have one electron in their outer shell. They are all soft solid metals at room temperature and react vigorously with water, producing hydrogen.

Properties of Group 1 elements

The Group 1 elements include lithium, sodium, potassium, rubidium and caesium. They are all soft metals with a low density. Lithium, sodium and potassium can be cut with a blade and are less dense than water.

Group 1 metals react vigorously with water. The reaction gives off hydrogen gas.

Group 1 metals also burn in oxygen to form **oxides**. Sodium burns to make sodium oxide.

7	Li	lithium	3
23	Na	sodium	11
39	K	potassium	19
85	Rb	rubidium	37
133	Cs	caesium	55

Figure 5.1.23 The Group 1 elements lithium to caesium

Figure 5.1.24 This is sodium on water. Why does it float?

1. **Explain why sodium floats on water.**
2. **Identify the gas given off when potassium reacts with water.**

Reaction trends of alkali metals

Group 1 metals react with non-metals, including oxygen and chlorine, to produce colourless ionic compounds.

They react easily with oxygen to produce oxides, and have to be stored under oil so that they can't react with oxygen in the air. The metals are more reactive with oxygen going down the group.

The word equation for the reaction between lithium and oxygen is:

lithium + oxygen → lithium oxide

The balanced symbol equation for the reaction is:

$4Li(s) + O_2(g) \rightarrow 2Li_2O(s)$

Group 1 metals react with water to produce a metal **hydroxide** and hydrogen gas. The hydroxide dissolves in the water to make an alkaline solution. This is why Group 1 are called the **alkali** metals.

REMEMBER!

You should try to remember the order of reactivity of the alkali metals.

Link: 5.1g → 5.1h

In a demonstration, a teacher placed a small amount of three Group 1 metals in a large trough of water behind a safety screen. The table shows the students' observations.

	Reaction with water – observations
Lithium	• floats on surface, does not melt • gently fizzes as hydrogen gas is given off • colourless solution remains • lithium + water → lithium hydroxide + hydrogen $2Li(s) + 2H_2O(l) \rightarrow 2LiOH(aq) + H_2(g)$
Sodium	• floats on surface, melts to form silvery ball • moves around surface as hydrogen gas is rapidly given off • colourless solution remains
Potassium	• floats on surface, melts to form silvery ball • whizzes around surface as hydrogen gas is very rapidly given off • purple flame observed • colourless solution remains

3 **Suggest how show that the solution remaining after the reaction of a Group 1 metal with water is alkaline.**

4 **Using the table, describe the trend in reactivity with water down Group 1.**

5 **Suggest why the teacher decided not to demonstrate the reaction of rubidium with water, using what you know about the trend in reactivity down Group 1.**

6 **Write balanced symbol equations for the reactions of sodium and potassium with water.**

DID YOU KNOW?

Sodium hydroxide can be used as oven cleaner.

Figure 5.1.25 Sodium hydroxide is a solid. It is often used as a reagent in solution

It is more dangerous to get onto your skin or eyes than acids.
Always wear safety glasses when handling chemicals, especially alkalis.

Making ions

Alkali metals have similar chemical properties. This is because atoms of all the elements in Group 1 have one electron in their outer shell. Like all metals, atoms of the alkali metals all react by losing their outer electrons to form positive ions. The atoms of alkali metals lose one electron to form a full outer shell. This is then a **stable electronic structure.**

When the atom loses one electron it forms a charged **ion**. It has one more positive charge in its nucleus than negative electrons surrounding it. So it is now a positive ion that carries a charge of +1.

Sodium reacts with chlorine to make sodium chloride. The sodium makes an ion that carries a +1 charge. It makes an ionic compound. The compound is a white solid that dissolves in water to form a colourless solution. The other alkali metals also react with chlorine to form chlorides.

7 **Explain why Group 1 atoms lose electrons.**

8 **Draw a 'dot and cross' diagram to show a lithium ion.**

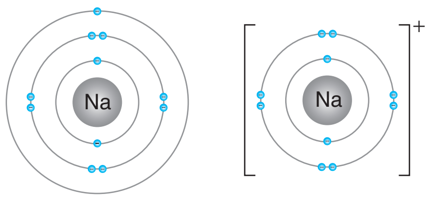


Figure 5.1.26 How alkali metals achieve a stable electronic structure

9 **Potassium reacts with oxygen to form an oxide.**

a **Write a balanced equation for the reaction.**

b **State the electron configurations of potassium and oxygen ions in the oxide.**

c **The oxide dissolves in water to form an alkaline solution. Identify this solution.**

5.1h Exploring Group 7

Learning objectives:

- recall that fluorine, chlorine, bromine and iodine are non-metals called halogens
- relate the properties of the halogens to the number of electrons in their outer shell
- predict the properties of Group 7 elements from trends down the group
- construct balanced symbol equations for the reactions of metals with halogens.

KEY WORDS

displacement reaction
halogen
molecule

Group 7 elements are known as the halogens. These all have one electron missing in their outer electron shells, and react vigorously with Group 1 metals, which have one electron in their outer shells.

The halogens

Group 7 elements, fluorine, chlorine, bromine and iodine, are called the **halogens**.

Halogens are non-metals. Halogens exist as pairs of atoms in **molecules**. Their symbols are F_2, Cl_2, Br_2 and I_2.

They react vigorously with metals such as sodium, potassium and magnesium to form ionic compounds called salts.

For example, chlorine reacts with potassium to make potassium chloride. The word equation is:

potassium + chlorine → potassium chloride

The balanced symbol equation for this reaction is:

$$2K(s) + Cl_2(g) \rightarrow 2KCl(s)$$

The halogens form molecular compounds with other non-metallic elements. These compounds are often gases, or liquids such as acids.

Figure 5.1.27 Chlorine reacting with potassium to make potassium chloride

MAKING LINKS

You will find out more about ionic and molecular compounds in Chapter 6.2.

1. **Give the balanced symbol equation for the reaction between sodium and bromine.**
2. **Identify the product of the reaction between lithium and fluorine.**

Group 7 trends

There is a trend in the physical appearance of the halogens at room temperature. Chlorine is a gas and iodine is a solid.

DID YOU KNOW?

The first use of chlorine was to bleach textiles and it is still used in chemicals for bleaching toilets. Chlorine is used to sterilise water to prevent diseases such as cholera from spreading. Chlorine was also used as a weapon in the First World War. The effects were devastating.

Figure 5.1.28 At room temperature chlorine is a green gas and iodine is a grey solid. What colour is the toxic and volatile bromine?

Halogen	Relative atomic mass	Relative formula mass	Melting point (°C)	Boiling point (°C)	State at room temperature
chlorine, Cl_2		71	–101	–34	gas
bromine, Br_2	80	160	–7	59	
iodine, I_2	127		114	184	

3. **Complete the data table (relative formula mass will be covered again in section 5.2d – for molecules containing two of the same atom, simply double the relative mass of one atom).**
4. **Explain how the state of iodine in Figure 5.1.28 is confirmed by the data in column 5 of the table.**
5. **Describe the trend in boiling point as molecular mass changes.**

Displacement reactions of halogens

The reactivity of the halogens decreases down the group.

If halogens are bubbled through solutions of metal halides there are two possibilities: no reaction, or a **displacement reaction**.

If chlorine is bubbled through potassium bromide solution a displacement reaction occurs. A red-brown colour of bromine is seen. This is because chlorine is more reactive than bromine. Chlorine displaces (takes the place of) bromine. The red-brown colour is the displaced bromine.

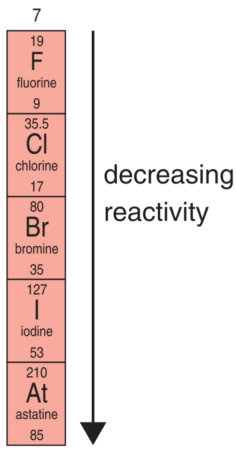


Figure 5.1.29

6. **Identify which halogens will displace iodine from a solution of potassium iodide. Justify your answer.**
7. **Write a balanced symbol equation for the reaction between chlorine and potassium iodide.**
8. **Astatine is at the bottom of Group 7. It is radioactive and extremely rare so its chemistry has not been well studied.**
   a **Predict the state of astatine at room temperature.**
   b **Write a balanced equation for the reaction between chlorine and sodium astatide, NaAt.**
   c **Explain whether astatine, At_2, will react with sodium iodide, NaI.**

Check your progress

You should be able to:

name elements and compounds from their symbols and formulae.	recall the names of the first 20 elements in the periodic table and the elements in Groups 1 and 7.	use symbol equations to describe chemical reactions.
use the periodic table as an example of how scientific theories have changed over time. describe how Mendeleev was able to leave spaces for elements that had not yet been discovered.	decide whether or not given data supports a particular theory. use the electronic structure model to explain why the modern periodic table has the elements in order of atomic number.	explain, using the periodic table, why ordering by relative atomic mass resulted in some elements being grouped wrongly. explain how Mendeleev was able to make predictions of as yet undiscovered elements.
describe the pattern of the electrons in shells for the first 20 elements.	explain how the electronic arrangement of atoms follows a pattern up to the atomic number 20.	explain why elements in the same group have similar chemical properties.
describe the physical and chemical properties of the alkali metals (Group 1) and the halogens (Group 7).	explain that all metals react by losing their outer electrons to form positive ions.	explain that metals react by losing electrons to form positive ions and non-metals do not form positive ions.
explain that non-metals are on the right-hand side of the periodic table.	explain that non-metals do not form positive ions.	explain the differences between the chemical properties and reactions of metals and non-metals in terms of their atomic structure/electronic structure.
describe the unreactivity of the noble gases.	describe the trend down Group 0 of increasing boiling point.	explain the trend down Group 0 of increasing boiling point in terms of relative atomic mass.
predict the reactions with water of Group 1 elements lower than potassium.	predict and explain the relative reactivity down the groups.	explain the trend down the group of increasing reactivity by electron structure.
recall the order of reactivity of chlorine, bromine and iodine.	describe the order of reactivity and explain the displacement reactions of halogens.	predict displacement reaction outcomes of halogens other than chlorine, bromine and iodine.
explain that a stable outer shell of electrons makes noble gases unreactive.	predict the properties of 'unknown' elements from their position in Group 1 or Group 7.	explain the trend of increasing reactivity in terms of electronic structure.

Worked example

Sam and Alex are researching some properties of Group 1 metals.

1 Shade the section of the periodic table where the Group 1 metals are found.

1 H hydrogen 1			
7 Li lithium 3	9 Be beryllium 4		
23 Na sodium 11	24 Mg magnesium 12		
39 K potassium 19	40 Ca calcium 20	45 Sc scandium 21	48 Ti titanium 22
85 Rb rubidium 37	88 Sr strontium 38	89 Y yttrium 39	91 Zr zirconium 40
133 Cs caesium 55	137 Ba barium 56	139 La lanthanum 57	178 Hf hafnium 72
223 Fr francium 87	226 Ra radium 88	227 Ac actinium 89	261 Rf rutherfordium 106

This is incorrect. The first column needs to be shaded.

The two metals they are researching are sodium and potassium.

2 Write down two properties that these metals have.

They are shiny when cut. They are metals so they are dense.

The first property is correct. However, sodium and potassium float on water so have a density less than that of water. This makes them unusual for metals.

Sam and Alex find out that sodium and potassium react with water. They find that sodium reacts with water to make sodium hydroxide and that hydrogen is given off.

3 Write a word equation for the reaction

sodium + water → sodium hydroxide.

The reactants are correct but hydrogen needs to be written as a product on the right-hand side.

Sam says that potassium reacts more vigorously than sodium but Alex says that they are in the same group so they react the same.

4 a Explain why Sam is correct about the trend.

The lower down the group the better they react.

This answer could be expressed more clearly by substituting the word 'better' with 'more vigorously'.

b Explain why Sam is correct using ideas about the structure of atoms.

The bigger the atom the quicker the reaction.

This answer needs more detail. The further away the outer electron is from the nucleus the more easily it is 'lost', as the pull by the positive nucleus on the negative electron is less.

End of chapter questions

Getting started

1. Identify the gas given off when sodium reacts with water. 1 Mark
2. Which of the following is a noble gas? 1 Mark

   A oxygen

   B helium

   C chlorine

   D nitrogen
3. Suggest two differences between a metal and a non-metal. 2 Marks
4. Explain why an atom with an electron pattern of 2,8,1 is in Group 1 of the periodic table. 2 Marks
5. To which group does chlorine belong? 1 Mark

   A 0

   B 1

   C 6

   D 7

Going further

6. When sodium reacts with water hydrogen is given off. Identify the other product. 1 Mark
7. Describe the trend of reactivity of the Group 1 metals. 2 Marks
8. Explain where the atom with an electron pattern of 2,8,2 is positioned in the periodic table. Suggest why it is a metal. 4 Marks
9. Halogen X has a boiling point of 59°C and halogen Y a boiling point of 184°C. 2 Marks

   a Justify which halogen has the lower atomic number.

   b Explain which halogen is more reactive.
10. An element has a mass number of 40 and electronic structure 2,8,8,2. Identify the element and work out the number of neutrons in its nucleus. 2 Marks

More challenging

11 **Explain why the reactivity of Group 1 metals increases down the group.** 2 Marks

12 **Francium is a highly radioactive and rare alkali metal.** 2 Marks

	Melting point (°C)	Density (g/ cm^3)
Rb	39.3	1.53
Cs	28.4	1.93
Fr	**D**	**E**

Predict the melting point, **D**, and density, **E**, of francium using the data in the table.

13 **Chlorine (Cl_2) is bubbled into a solution of potassium iodide (KI).**

a **Describe the reaction.**

b **Explain why this reaction is able to take place.** 4 Marks

Most demanding

14 **Explain why the atom with an electron pattern of 2,8,6 is a non-metal. Explain why this atom is less reactive than the atom with an electron pattern of 2,6 or 2,7.** 4 Marks

15 **The table shows the data and properties for an element.**

Deduce which type of element it is, the group it belongs to and why it reacts so violently with water. 4 Marks

Atomic number	Electronic structure	Reaction with water
37	2,8,8,18,1	Violently, giving off hydrogen and forming an alkali

Total: 34 Marks

CHEMICAL QUANTITIES

IDEAS YOU HAVE MET BEFORE:

ELEMENTS, MIXTURES AND COMPOUNDS

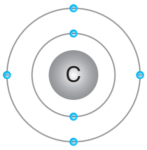
- Elements cannot be broken down by chemical means.
- Compounds are made from elements chemically combined.
- Chemical symbols are used to show the formulae for elements and compounds.

ATOMS ARE REARRANGED DURING CHEMICAL REACTIONS BUT NOT CREATED OR DESTROYED

- Iron and sulfur heated together make iron sulfide.
- Magnesium burns in air to make magnesium oxide.
- Wood burns to make ash, smoke and carbon dioxide.

ATOMS AND THEIR STRUCTURE

- Atoms have a nucleus with a positive charge.
- The nucleus contains neutrons and protons.
- Isotopes of an element have the same number of protons but different numbers of neutrons.



MEASURING QUANTITIES

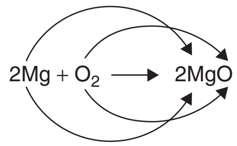
- Balances measure mass in grams.
- Measuring cylinders measure volume in cm^3.
- Stopwatches measure time in minutes and seconds.

5.2

IN THIS CHAPTER YOU WILL FIND OUT ABOUT:

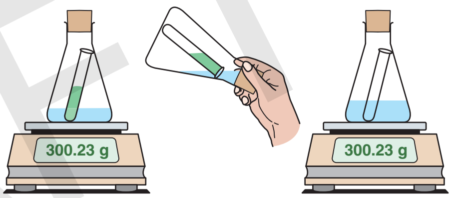
HOW DO WE USE SYMBOLS AND FORMULAE TO SHOW REACTIONS?

- The formula shows the proportions of elements combined in a compound.
- Chemical reactions can be shown by word equations or balanced equations using symbols and formulae.
- State symbols (s, l, g and aq) show the physical state of products and reactants.



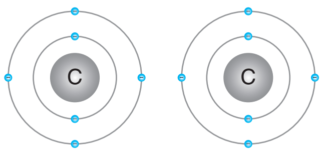
HOW IS MASS CONSERVED IN CHEMICAL REACTIONS?

- The mass of the reactants is the same as the mass of products.
- The same number of atoms is on both sides of an equation.
- An equation needs to be balanced.



WHAT OTHER INFORMATION CAN THE ATOMIC STRUCTURE PROVIDE?

- Relative atomic mass is the average relative mass of an element's isotopes.
- Relative formula mass is the sum of the relative atomic masses of the atoms in the formula.
- The sum of the relative formula masses of the reactants equals the sum of the relative formula masses of the products.



HOW CAN WE MEASURE AMOUNTS OF SUBSTANCES?

- In Chemistry, amounts of substances are measured in moles.
- Masses, in grams, and amounts, in moles, can be interconverted using molar masses.
- Amounts to be made can be predicted using equations.

5.2a Chemical equations

Learning objectives:

- explain that compounds are formed from elements by chemical reactions
- explain how formulae represent elements and compounds
- write equations for simple reactions including the physical states of products and reactants.

KEY WORDS

balanced equation
chemical equation
chemical symbol
compound
formula

Chemical formulae tell us which elements are combined to form a compound and in what proportions. Using chemical formulae, we can form chemical equations to represent chemical reactions.

Elements and compounds

There are about 100 different elements. Each element is shown by a **chemical symbol** in the periodic table. Elements are made up of atoms that are all the same.

A **compound** is a substance that contains atoms of at least two different elements. Compounds are formed by chemical reactions. Compounds can only be separated into elements by a chemical reaction.

The atoms in a compound combine in a fixed ratio. Water has the chemical **formula** H_2O. The formula shows that for every oxygen atom, there are two hydrogen atoms. We say that the ratio of hydrogen to oxygen is 2 : 1.

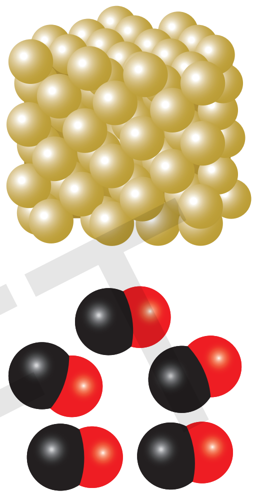


Figure 5.2.1 Gold is an element. Carbon monoxide is a compound. How could you tell just by looking at these diagrams?

1 **Explain whether the following substances are elements or compounds.**

C (carbon) **CO_2 (carbon dioxide)**

Cl_2 (chlorine) **SO_3 (sulfur trioxide)**

2 **For each of the substances below, write down:**

a **the names of the elements**

b **how many different types of atoms they contain**

c **how many atoms are in the molecule overall.**

CO_2 S_8 Cl_2 SO_3 C_{60} C_4H_{10}

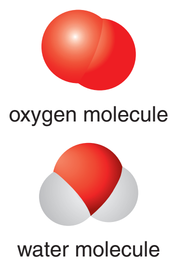


Figure 5.2.2 Which of these is a molecule of an element? How can you tell?

DID YOU KNOW?

The element oxygen exists as a pair of atoms, not as a single atom. Its formula is O_2. The subscript 2 shows that there are 2 atoms of oxygen in its formula.

Formulae and equations

Compounds are formed by chemical reactions. Chemical reactions always involve the formation of one or more new substances, and often involve a detectable energy change.

The **chemical equation** for a reaction tells us what quantities of each chemical react with each other.

The word equation for burning carbon in air is:

carbon + oxygen → carbon dioxide

The symbol equation is:

$C(s) + O_2(g) \rightarrow CO_2(g)$

This reaction releases thermal energy. The carbon dioxide formed has two oxygen atoms for every carbon atom, so on the left-hand side of the equation there must also be two oxygen atoms for every carbon atom.

We show the physical state of each substance using (s) for solid, (l) for liquid, (g) for gas and (aq) for an aqueous solution. The chemical equation for the reaction between solutions of sodium hydroxide and hydrochloric acid is written as:

$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

Figure 5.2.3 Wood contains carbon, and releases energy and carbon dioxide as it combusts

DID YOU KNOW?

Compounds can only be separated into elements by chemical reactions. Sodium chloride can be broken down to sodium and chlorine using electricity in a reaction called electrolysis.

3 **Identify the states of the substances in these equations:**

a $KI(aq) + AgNO_3(aq) \rightarrow KNO_3(aq) + AgI(s)$

b $ZnCO_3(s) \xrightarrow{heat} ZnO(s) + CO_2(g)$

4 **Elements from Group 1 and Group 7 combine to make compounds in the fixed ratio 1 : 1, for example lithium chloride (LiCl). Write the formulae and names of three more examples of such compounds.**

Equations and balancing

When calcium reacts with oxygen it makes magnesium calcium. The word equation is:

calcium + oxygen → calcium oxide

The symbol equation is:

$Ca(s) + O_2(g) \rightarrow CaO(s)$

1 2 1 1 ✗

This is not a **balanced equation**. We need to add '2' to the front of the formulae of calcium and calcium oxide:

$2Ca(s) + O_2(g) \rightarrow 2CaO(s)$

2 2 2 2 ✓

5 **Write a balanced equation for the formation of sodium chloride from sodium, Na, and chlorine, Cl_2, including state symbols.**

6 **Write a balanced equation for the formation of aluminium oxide, Al_2O_3, from its elements.**

7 **Complete and balance the following equation by suggesting values for d, e and f:**

$dN_2 + eO_2 \rightarrow fNO_2$

5.2b Conservation of mass and balanced equations

Learning objectives:

- explain the law of conservation of mass
- explain why a multiplier appears as a subscript in a formula
- use ratio when writing and balancing equations.

KEY WORDS

balanced equation
conservation of mass
products
ratio
reactants

When a solid changes state to become a gas, it *appears* to lose mass, but this isn't the case. When the product of a reaction is a gas, some of the gas might escape, but if you could collect and measure the mass of all the gas and other reaction products, the mass of the products would be the same as the mass of the reactants.

The law of conservation of mass

The law of **conservation of mass** states that 'no atoms are lost or made during a chemical reaction'. This means that the mass of the **products** equals the mass of the **reactants**. In a chemical reaction, all the atoms in the reactants are *rearranged* into products. For example, for one type of reaction:

AB + CD → AD + CB

the mass of atoms at the start equals the mass of atoms at the finish. This is consistent with the 'conservation of mass'.

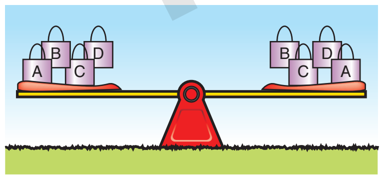


Figure 5.2.4 Masses of reactants and products are balanced

1 Identify the missing product to ensure conservation of mass.

XY + ZRT → XR + ? + T

Balanced equations and formulae

Many compounds contain more than one atom of an element in a formula. The number of atoms of the element is written as a subscript that is written after the element symbol.

For example, the formulae in the table have more than one atom of some of their elements. If there is no number after the element symbol it means 1 atom.

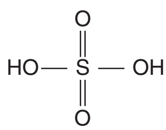
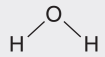
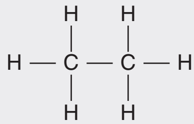
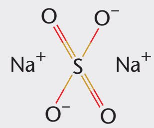
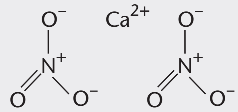


Figure 5.2.5 The molecule has 2 H atoms, 1 S atom and 4 O atoms. Its formula is H_2SO_4

carbon dioxide	CO_2	O=C=O
water	H_2O	
ethane	C_2H_6	
magnesium chloride	$MgCl_2$	Cl^- Mg^{++} Cl^-
sodium sulfate	Na_2SO_4	
calcium nitrate	$Ca(NO_3)_2$	

2 **Determine the number of hydrogen atoms in the formula for ethanoic acid, CH_3COOH.**

3 **How many atoms of each element are in the formula for sodium thiosulfate, $Na_2S_2O_3$?**

4 **How many atoms of each element are in the formula for aluminium sulfate, $Al_2(SO_4)_3$?**

Balancing equations

This equation is balanced.

$$NaCl + AgNO_3 \rightarrow NaNO_3 + AgCl$$

Here is another example:

$$Mg + O_2 \rightarrow MgO$$

This equation is not balanced.

The two atoms of oxygen will each join with a magnesium atom to make two formula units of MgO.

$$Mg + O_2 \rightarrow 2MgO$$

This equation is still not balanced.

If we make 2 formula units of MgO, there must have been 2 atoms of Mg as reactants to 'conserve mass'.

$$2Mg + O_2 \rightarrow 2MgO$$

There are the same number of atoms of each type on the two sides and 'mass is conserved'. A balanced chemical equation can tell us the ratio of the number of atoms or molecules of each type reacting.

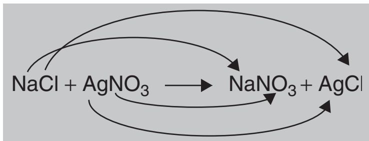
5 **Balance the following equations by determining the numerical values for d, e, f and g.**

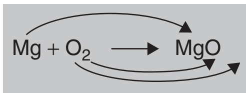
a) $dNaOH + eH_2SO_4 \rightarrow fNa_2SO_4 + gH_2O$

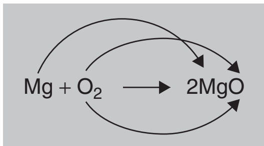
b) $dC_3H_8 + eO_2 \rightarrow fCO_2 + gH_2O$

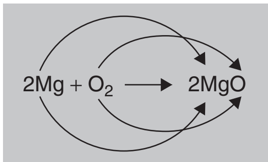
DID YOU KNOW?

You must write the equation of the reaction that actually happened even if it is difficult to balance. You cannot alter the formulae of compounds even if this makes balancing easier.









5.2c Explaining observed changes in mass

Learning objectives:

- explain any observed changes in mass in a chemical reaction
- identify the mass changes using a balanced symbol equation
- explain these changes in terms of the particle model.

KEY WORDS

gas
mass
particles

Burning – combustion – is an oxidation reaction. When iron wool burns, the iron reacts with the oxygen in the air, making iron oxide. The iron oxide has more mass than the original iron wool. By conservation of mass, the mass of the iron oxide is equal to the mass of the iron and the oxygen before the reaction.

Losing mass

If baking powder is heated, it gives off carbon dioxide, which makes cakes rise. It seems as if the mass of baking powder is more before it is heated than after. This is because it rearranges and loses the carbon dioxide.

Some reactions may seem to involve a change in **mass**. This can usually be explained because a reactant or product is a **gas**. The mass of the gas is often not measured.

A gas can be driven off as a product or taken in as a reactant.

For example: when copper carbonate is heated, it reacts to make copper oxide and carbon dioxide. Where does the carbon dioxide gas go?

Jaheer heats 5.0 g of copper carbonate.

After heating, Jaheer measured the mass of the copper oxide.

It had a mass of 3.2 g.

How much gas was made?

The answer is 1.8 g.

Figure 5.2.6 Heating copper carbonate

copper carbonate → copper oxide + carbon dioxide

$CuCO_3(s) \rightarrow CuO(s) + CO_2(g)$

Figure 5.2.7 Solid copper carbonate when heated turns to solid copper oxide and the gas carbon dioxide

1. Sarita **heats 2 g of zinc carbonate.**

   zinc carbonate → zinc oxide + carbon dioxide

   2 g of zinc carbonate makes 1.3 g of zinc oxide. Explain why the mass of zinc oxide is less than the mass of zinc carbonate.

2. **In the reaction in question 1, what mass of carbon dioxide is produced?**

3. **Copper sulfate crystals change into copper sulfate powder when heated. Water is given off. If 2.5 g of copper sulfate crystals make 1.6 g of powder, how much water is produced?**

KEY INFORMATION

The conservation of mass means that the mass of the products must equal the mass of the reactants.

Gaining mass

Sometimes one of the reactants is a gas.

For example: when a metal reacts with oxygen the mass of the oxide produced is greater than the mass of the metal.

Alesha heats 4.0 g of copper. When it cools the solid product has a mass of 5.0 g. Why is there an increase?

The answer is that the copper reacted with 1.0 g of oxygen to form copper oxide.

4 **When magnesium is heated in air it gains mass. If 2.4 g of magnesium makes 4.0 g of magnesium oxide, how much oxygen was added from the air?**

Figure 5.2.8 Blue, hydrated copper sulfate becomes white when heated and the water is driven off

HIGHER TIER ONLY

Other reactions and limiting reactants

If magnesium carbonate is put into acid in a flask and the flask is on top of a digital balance, what is observed? The mass of the flask and its contents decrease.

This is because carbon dioxide is being given off.

$$MgCO_3 + 2HCl \rightarrow MgCl_2 + H_2O + CO_2$$

If the mass decrease is measured every 2 minutes, a graph such as Figure 5.2.11 could be drawn to use in an investigation.

It is not only carbon dioxide or water vapour that will be produced in a reaction. Other reactions will produce other gases such as hydrogen. These reactions will produce a decrease in mass also.

$$Mg + 2HCl \rightarrow MgCl_2 + H_2$$

If the acid in this reaction is in excess, then the magnesium is the limiting reactant. Once the magnesium has been used up no more hydrogen can be made.

5 **Look at the graph in Figure 5.2.10. Explain when the reaction stopped and why it stopped.**

6 **The reaction between 8.4 g of $MgCO_3$ and excess hydrochloric acid solution stops after 7 minutes. The mass of the flask and contents has decreased by 4.4 g. Sketch the graph you might obtain when monitoring this reaction.**

7 **6.54 g of zinc was added to 6.97 g of nitric acid. At the end of the reaction, 2.92 g of zinc remained.**

a **Explain which of the reactants was the limiting reactant.**

b **State the mass of zinc that reacted.**

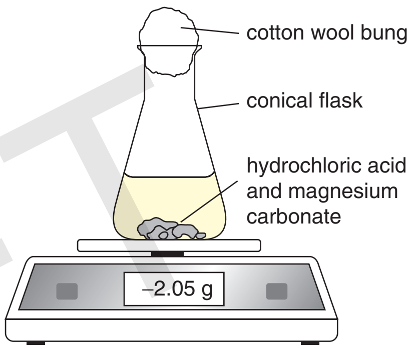


Figure 5.2.9 Measuring the mass of carbon dioxide given off

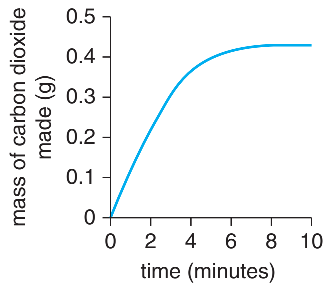


Figure 5.2.10 Graph showing mass of carbon dioxide given off

5.2d Relative formula mass

Learning objectives:

- identify the relative atomic mass of an element from the periodic table
- calculate relative formula masses from relative atomic masses
- verify the law of conservation of mass in a balanced equation.

KEY WORDS

relative atomic mass
relative formula mass

When atoms join together to form a compound, the compound has the same mass as the individual atoms added together. The formula of a chemical substance shows the fixed ratio in which the atoms combine. So by adding together the masses of all the atoms in the formula, we can find the mass of the chemical substance.

Mass number

Atoms are made up of protons, neutrons and electrons. Protons and neutrons make up the nucleus with the electrons orbiting the nucleus in 'shells'. Most of the atom is empty space with almost all the mass in the nucleus.

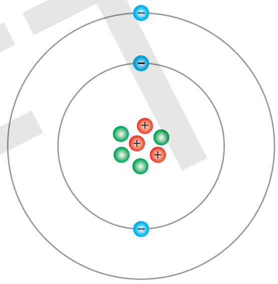


Figure 5.2.11 The lithium atom has three protons and four neutrons

A lithium atom has three protons, so has an *atomic number* of 3.

A lithium atom has four neutrons, so has a *mass number* of 3 + 4 = 7.

We can show this in symbols as $^{7}_{3}Li$.

Some versions of the periodic table lists the mass numbers of the elements. We saw in Chapter 5.1 that some elements have *isotopes*, where atoms of an element have the same number of protons but different numbers of neutrons. This means the different isotopes have different mass numbers. Chlorine has two main isotopes, ^{35}Cl and ^{37}Cl.

1. **Sodium atoms contain 11 protons and 12 neutrons. Cobalt atoms contain 27 protons and 32 neutrons. Aluminium atoms contain 13 protons and 14 neutrons. What are the mass numbers for sodium, cobalt and aluminium?**
2. **The mass number of argon is 40 and its atomic number is 18. How many neutrons does it have?**

Relative atomic mass

Atoms are very very small, so it is more useful to know how the mass of one compares to the mass of another than the actual values. Chemists first compared other atoms to hydrogen, giving hydrogen the value 1. Later they found that it was better to use the isotope of carbon that has mass number 12.

The **relative atomic mass** (A_r) of an element is the mass of the atoms of the element compared with the carbon-12 isotope. It is an average value for all the isotopes of the element.

3. **Show the atomic structure of the three isotopes of carbon, $^{12}_{6}C$, $^{13}_{6}C$ and $^{14}_{6}C$, in terms of protons and neutrons.**

Relative formula mass

The **relative formula mass** (M_r) of a compound is the sum of the relative atomic masses of the atoms in the numbers shown in the formula.

For example: M_r of MgO	= A_r of Mg	+ A_r of O		
	= 24	+ 16		= 40
M_r Na_2O	= (23 × 2)	+ 16		= 62
M_r CO_2	= 12	+ (16 × 2)		= 44
M_r $CuCO_3$	= 63.5	+ 12 + (16 × 3)		= 123.5
M_r $Mg(OH)_2$	= 24	+ (16 × 2) + (1 × 2)		= 58

In a balanced chemical equation, the sum of the relative formula masses of the reactants in the quantities shown equals the sum of the relative formula masses of the products in the quantities shown.

	$ZnCO_3$	→	ZnO	+ CO_2
M_r are:	65 + 12 + (16 × 3)		65 + 16	12 + (16 × 2)
	65 + 12 + 48	→	81 + 12 + 32	
	125	→	125	

4. **Work out the relative formula mass of $MgSO_4$.**
5. **Calculate the relative formula mass of $Cu(NO_3)_2$.**
6. **Show that the sums of the relative formula masses of reactants and products are equal in this reaction:**

$$MgBr_2 + 2AgNO_3 \rightarrow Mg(NO_3)_2 + 2AgBr$$

7. **X is a molecule containing carbon and hydrogen. Work out the molecular formula and relative formula mass of X.**

$$X + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

KEY INFORMATION

The relative formula mass of a formula with brackets can be worked out in two ways.
Either M_r $Mg(OH)_2$ = 24 + (16 × 2) + (1 × 2) = 58
or = 24 + (16 + 1) × 2 = 58

DID YOU KNOW?

This holds true for any quantity, so if the relative formula mass of the reactants is 94, then the relative formula mass of the products will be 94. This can be applied to any quantity, so 94 g of reactants will produce 94 g of products, or 94 tonnes of reactants will produce 94 tonnes of products.

KEY INFORMATION

The full size '2' before a formula applies to all of the formula. The subscript '$_2$' after an element symbol applies to just the one element. Brackets are used if the '$_2$' applies to more than just one element symbol.

5.2e Amounts in moles

Learning objectives:

- describe the measurement of amounts of substances in moles and be able to define the Avogadro constant
- calculate the number of moles in a given mass
- calculate the mass of a given number of moles.

KEY WORDS

Avogadro constant
molar mass
mole
relative formula mass

The relative formula mass of a substance tells you the mass of one mole of it in grams. To find the mass of a single molecule, divide this mass by 6.02×10^{23}.

HIGHER TIER ONLY

Measurement of amounts

Atoms and molecules are so tiny, it's impossible to experiment with them individually. Instead, we deal with them in set amounts, called **moles** (mol). 1 mol of a substance contains the same number of particles as there are atoms in 12 g of carbon-12 (the isotope of carbon with relative atomic mass, A_r = 12).

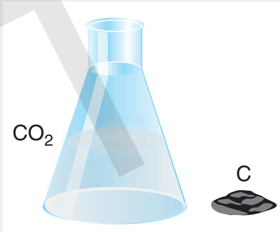


Figure 5.2.12 This lump of carbon (C) and jar of carbon dioxide (CO_2) may not look the same but they have the same amount of substance. There is the same number of atoms in the carbon lump as the number of molecules of carbon dioxide in the jar

The measurement of amounts in moles can apply to:

atoms	ions	formulae
molecules	electrons	equations

For example:

In one mole of carbon (C) the number of atoms is the same as the number of molecules in one mole of carbon dioxide (CO_2).

The mass of one mole of a substance in grams is numerically equal to its **relative formula mass**.

Number of atoms in 1 mole C	= number of molecules in 1 mole CO_2
Carbon has a formula of C	Carbon dioxide has a formula of CO_2
C has a relative atomic mass of 12	CO_2 has a relative molecular mass of 12 + (2 × 16) = 44
Mass of 1 mol = 12 g	Mass of 1 mol = 44 g

1. **Work out the mass of one mole of H_2O.**

   (Relative atomic mass of H is 1, relative atomic mass of O is 16.)

2. **Work out the mass of 3 moles of KBr. (Hint: Use the periodic table to help you.)**

The Avogadro constant

The number of atoms, molecules or ions in a mole of a given substance is the Avogadro constant.

The value of the **Avogadro constant** is 6.02×10^{23} per mole (or 6.02×10^{23} mol^{-1}).

3 **Calculate the number of particles in two moles of helium, He.**

DID YOU KNOW?

Avogadro's hypothesis was developed in 1810, over 200 years ago.

Calculating molar mass

The name for the mass of one mole is the molar mass and its unit is g/mol or g mol^{-1}.

A_r is the abbreviation used for relative atomic mass. The mass of one mole of an element that consists of atoms is numerically equal to the relative atomic mass, A_r on the periodic table. For example:

- Ne has a relative atomic mass, A_r, of 20
- 1 mole of neon atoms, Ne has a molar mass of 20 g/mol

The mass of one mole of an element that exists as a molecule is found by multiplying the relative atomic mass by the number of atoms in the molecular formula.

Oxygen gas has the formula O_2.

- The relative atomic mass, A_r for oxygen is 16.
- In O_2, there are two oxygen atoms, so the relative formula mass is $2 \times 16 = 32$
- The molar mass for O_2 is 32 g/mol.

The mass of one mole of a compound is found by adding up all the relative atomic masses in the formula.

- Calcium carbonate has the formula $CaCO_3$.
- Relative atomic masses A_r are Ca = 40, C = 12, O = 16.
- The relative formula mass of $CaCO_3$ is $40 + 12 + (3 \times 16) = 100$
- The molar mass is 100 g/mol.

If the formula contains brackets, these must be considered in the calculation:

- Magnesium nitrate has the formula $Mg(NO_3)_2$.
- Relative atomic masses, A_r of Mg = 24, N = 14, O = 16.
- Relative formula mass of $Mg(NO_3)_2 = 24 + (2 \times 14) + 2\,(3 \times 16) = 148$
- The molar mass is 148 g/mol.

4 **Work out the molar mass of:**

a **nitrogen gas, N_2**

b **zinc oxide, ZnO**

c **magnesium carbonate, $MgCO_3$**

d **ammonium sulfate, $(NH_4)_2SO_4$**

5 **Work out the amount, in moles, of molecules in 72 g of water.**

6 $2H_2 + O_2 \rightarrow 2H_2O$

a **Work out the number of moles of O_2 needed to make four moles of water.**

b **Calculate the number of grams of H_2 needed to make four moles of water.**

5.2f Using balanced equations to calculate masses

Learning objectives:

- calculate the masses of substances in a balanced symbol equation
- calculate the masses of reactants and products from balanced symbol equations
- calculate the mass of a given reactant or product.

KEY WORDS

mass
mole
product
reactant

Chemists measure the amounts of substances in *moles*. A balanced equation for a reaction shows us how many moles of reactant will produce how many moles of product. From this, we can use molar masses to work out what mass of reactant is needed to produce a certain mass of product.

HIGHER TIER ONLY

Masses of substance from an equation

The **masses** of **reactants** and **products** can be calculated from balanced symbol equations. For example, looking at the reaction between magnesium and oxygen to form magnesium oxide:

$$2Mg + O_2 \rightarrow 2MgO$$

we can see that:

2 moles of Mg react with 1 mole of O_2 to produce 2 moles of MgO.
We know that A_r of Mg = 24, and A_r of O = 16, so the molar masses are as follows.

Molar mass of Mg is 24 g/mol.
Molar mass of O_2 is 32 g/mol.
Molar mass of MgO is 40 g/mol.

Now we can work out the masses of reactants and products in the reaction like this:

mass of substance (g) = amount (mol) x molar mass (g/mol)

So (2 x 24) g Mg reacts with (1 x 32) g O_2, to give (2 x 40) g MgO.

Or,

48 g Mg + 32 g O_2 → 80 g MgO

If we only need to produce, say, 8 g MgO, we could scale down the masses of reactants by the same factor, and use 4.8 g Mg and 3.2 g O_2.

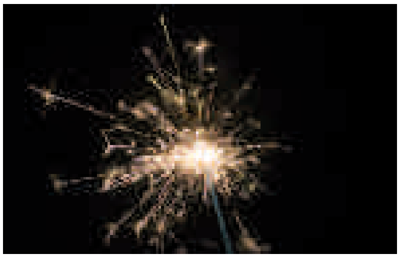


Figure 5.2.13 Magnesium reacting with oxygen to form magnesium oxide

Link: 5.2f → 5.2g

1. **Calculate the mass of MgO made from 6.0 g of Mg.**
2. **Calculate the mass of Mg needed to make 2.0 g of MgO.**

Different ways of measuring amounts in moles

Chemical equations can be interpreted in terms of **moles**. Another example is:

$Mg + 2HCl \rightarrow MgCl_2 + H_2$

This shows that one mole of magnesium reacts with two moles of hydrochloric acid to produce one mole of magnesium chloride and one mole of hydrogen gas. These are the ratios in which reactants and products relate. The actual amount of moles made will be in these ratios but can be measured in terms of mass, concentration or volume.

3. **Determine the number of moles of H_2O that will be made from six moles of propane on combustion with O_2.**

   $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$

DID YOU KNOW?

The number of moles of Mg that is actually used can be found from the mass of Mg used measured in grams. The number of moles of HCl can be found from the concentration of the acid in g/dm³. The number of moles of H_2 made can be found from the volume collected compared with the molar volume.

Predicting masses

Moles can also be used to predict masses from equations.

This is the equation for burning a fuel called heptane, C_7H_{16}:

$$C_7H_{16} + 11O_2 \rightarrow 7CO_2 + 8H_2O$$

What mass of carbon dioxide, CO_2, is formed when 100 g of C_7H_{16} is burned?

Stage 1: find the **molar masses**

Molar mass of C_7H_{16} is (7 × 12 g/mol) + (16 × 1 g/mol) = 100 g/mol

Molar mass of CO_2 is (1 × 12 g/mol) + (2 × 16 g/mol) = 44 g/mol

Stage 2: use the number of moles from the equation to find the ratio

From the equation, 1 mole of heptane produces 7 moles of CO_2.

The mass of 7 moles of CO_2 is 7 mol × 44 g/mol = 308 g

So 100 g of heptane will produce 308 g of carbon dioxide.

Use the relative atomic masses in the periodic table to help you answer Questions 4 and 5.

4. **Find the mass of ZnO made when 1.25 g of $ZnCO_3$ are thermally decomposed**

   $ZnCO_3 \rightarrow ZnO + CO_2$

5. **Find the mass of $CuCO_3$ needed to make 7.95 g of CuO.**

5.2g Balancing equations

Learning objectives:

- convert masses in grams to amounts in moles
- balance an equation given the masses of reactants and products
- change the subject of a mathematical equation.

KEY WORDS

balanced equations
molar mass
moles
thermal decomposition

We do not always know the balanced equation for a reaction. If we can measure the masses of the reactants and products, we can work out the number of moles of each reactant required and how many moles of product are made. This gives us another way to form and balance chemical equations.

HIGHER TIER ONLY

Measuring amounts in moles

The numbers in a balanced equation can be calculated from the masses of reactants and products by converting the masses in grams into amounts in moles and converting these to simple whole number ratios.

Let's look at the **thermal decomposition** of magnesium carbonate that produces magnesium oxide (MgO) and carbon dioxide (CO_2). We find that if 42.0 g of $MgCO_3$ is heated then 20.0 g of MgO is produced.

We have already learned that:

mass of substance (g) = **molar mass** (g/mol) × amount of substance (mol)

So this equation can be rearranged to find the number of moles:

$$\text{amount of substance (mol)} = \frac{\text{mass of substance (g)}}{\text{molar mass (g/mol)}}$$

So to calculate the number of moles in 20.0 g of MgO:

First calculate the molar mass of MgO:

(A_r Mg = 24, A_r O = 16) M_r MgO = 24 + 16 = 40.

Then use the equation above: Molar mass MgO is 40 g/mol.

$$\text{amount of MgO (mol)} = \frac{20.0\text{ g}}{40\text{ g/mol}} = 0.5\text{ mol}$$

Next, let's calculate how many moles of $MgCO_3$ there are in 42.0 g.

Figure 5.2.14 The thermal decomposition of $MgCO_3$

KEY INFORMATION

In this case 22.0 g of CO_2 would have been driven off: remember the law of conservation of mass.

First we find the molar mass of $MgCO_3$: (A_r Mg = 24, A_r C = 12 A_r O = 16)

$$M_r\ MgCO_3 = 24 + 12 + (3 \times 16) = 84$$

Molar mass $MgCO_3$ is 84 g/mol.

Then use the equation above:

$$\text{amount of } MgCO_3 \text{ (mol)} = \frac{42.0\ \text{g}}{84\ \text{g/mol}} = 0.5\ \text{mol}$$

1 **What is the mass of MgO that would be produced by 84 tonnes of $MgCO_3$?**

Balancing the equation

$$MgCO_3 \rightarrow MgO$$

So this calculation tells us that:

0.5 mol $MgCO_3$ → 0.5 mol MgO

thus 1 mol of $MgCO_3$ makes 1 mol of MgO.

The ratio is 1 : 1.

The ratio of moles in the equation is therefore:

$$MgCO_3 \rightarrow MgO + ?\ CO_2$$

But what about the number of moles of CO_2? Looking at the formula there must be 1 mol of CO_2 produced because there was 1 mol carbon in the reactant.

We need to check by mass. To conserve mass, 22 g of CO_2 would have been produced. Molar mass of CO_2 is 44 g/mol.

$$\text{amount of } CO_2 \text{ (mol)} = \frac{\text{mass of } CO_2 \text{ (g)}}{\text{molar mass of } CO_2 \text{ (g/mol)}} = \frac{22.0\ \text{g}}{44\ \text{g/mol}} = 0.5\ \text{mol}$$

So 0.5 mol of $MgCO_3$ makes 0.5 mol of MgO and 0.5 mol of CO_2, which is a 1 : 1 : 1 ratio.

The **balanced equation** is therefore:

$$MgCO_3\ (s) \rightarrow MgO\ (s) + CO_2\ (g)$$

Applying the calculations

Aluminium oxide, Al_2O_3 produces aluminium, Al and oxygen, O_2.

If 204 g of Al_2O_3 produce 108 g of Al, work out the number of moles of Al_2O_3, Al and O_2 involved and hence write out the full balanced equation.

2 **Explain how you worked out the number of moles of Al_2O_3 and Al from the masses given.**

3 **Explain how you worked out the number of moles of O_2 and how you deduced the mole ratio of the three substances.**

DID YOU KNOW?

In this example we found that 0.5 moles of $MgCO_3$ produced 0.5 moles of MgO. So we can say that 1 mole of $MgCO_3$ would produce 1 mole of MgO.

KEY CONCEPT

5.2h Amounts of reactants and products

Learning objectives:

- identify which reactant is in excess
- explain the effect of a limiting quantity of a reactant on the amount of products
- calculate amount of products in moles or masses in grams.

KEY WORDS

directly proportional
excess
limiting

Chemicals react with each other in fixed ratios. In a reaction between two chemicals, if there is more than is needed of one reactant, then it will be left over at the end. Using an excess of one reactant makes sure the other reactant is used up.

MAKING LINKS

See topic 7.3c Making salts and topic 7.4a Measuring rates for more information.

HIGHER TIER ONLY

Reactants in excess

In a reaction between magnesium and dilute hydrochloric acid a small amount of one of the reactants is used and a lot more of the other reactant is used. We need to decide:

- when to use a lot more of one reactant
- what is 'a lot more'.

For example, when making the salt, magnesium chloride, all the acid needs to be used up, as it is hard to separate the magnesium chloride as a solid from the solution if there is still acid present.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

It is much easier to have too much magnesium metal, use up all the acid reacting with the metal, then filter off the solution from the **excess** magnesium solid.

In this case the **limiting** reagent is the hydrochloric acid and the excess is magnesium.

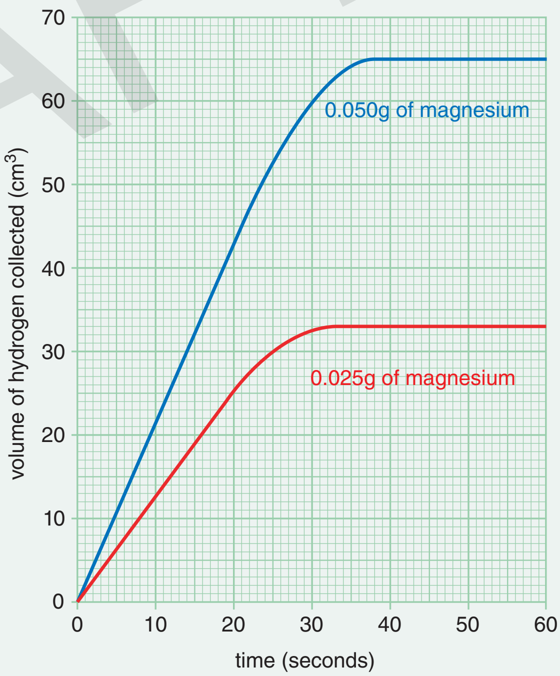


Figure 5.2.15 How much more hydrogen is formed if the amount of magnesium is doubled?

On the other hand if we are investigating the rate of reaction between magnesium and hydrochloric acid, we will want to use up all of the magnesium, as it is easier to measure the mass of the solid than adjust the concentration and volume of acid. So we use excess acid and a known amount of magnesium.

Magnesium is used up by the end of the reaction so it is known as the *limiting reactant*.

Look at the graph in Figure 5.2.15.

In this reaction, both Mg and HCl are the reactants.
To find the rate of reaction we limit the amount of magnesium.

1. **Use construction lines to find the rate of reaction when 0.050 g of magnesium is used (blue line).**
2. **Determine the volume of hydrogen formed when:**
   - **a 0.025 g of magnesium is used.**
   - **b 0.050 g of magnesium is used.**

   Describe the relationship between the amount of reactant used and the amount of product formed.
3. **Predict the volume of hydrogen if 0.0125 g of magnesium is used. Sketch this reaction line on a copy of the graph.**

ADVICE

The limiting reactant *limits* the amount of product that can be made.

How much product?

The amount of product is **directly proportional** to the amount of limiting reactant used. So, if the amount of this reactant doubles, the amount of product doubles. The more reactant that is used, the more product it forms.

4. **Explain what is meant by 'limiting reactant'.**
5. **Predict what will happen to the amount of product if the limiting reactant is quadrupled.**

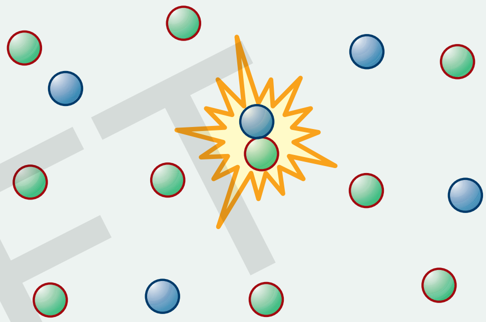


Figure 5.2.16 In this model of colliding particles, is it the green or the blue particles that represent the limiting reactant?

Molar masses and moles again

Mg(s) + 2HCl(aq) → $MgCl_2$(aq) + H_2(g)

Magnesium is the limiting reactant.

(*Ar* Mg = 24, *Ar* H = 1, *Ar* Cl = 25.5)

6. **In the reaction above, if 2 mol of $MgCl_2$ are needed,**
   - **a determine how many moles of Mg must be used at the start**
   - **b determine the minimum number of moles of HCl that must be used for the reaction.**
7. **Work out the molar masses of the reactants and products.**
8. **Work out the mass of magnesium chloride that could be made from 3 g of magnesium.**

REMEMBER

If you looked up the relative atomic mass of an element and measured it out in grams, this would be the same as the molar mass.

5.2i Concentrations of solutions

Learning objectives:

- know that concentration is mass per given volume of solution
- calculate the mass of solute in solution
- explain how concentration is related to mass and volume.

KEY WORDS

concentration
solute
solution
solvent

Many chemical reactions take place in solutions. A solution is formed when a solute is dissolved in a solvent, and solutions can be dilute or concentrated. Dilutions are often critical, such as making the correct formulations of medicines or baby milk formula. We use the mass per volume to make sure we have the correct concentration.

Concentration

Many chemical reactions take place in **solutions**. A solution with a greater **concentration** has more particles in a fixed volume than a less concentrated solution.

In a solution, the particles are called the **solute** and the liquid is called the **solvent**.

The concentration of a solution can be measured as mass of solute per given volume of solution.

We can write this as: $\text{concentration} = \frac{\text{mass of solute}}{\text{volume of solution}}$

Concentration is often given as grams per dm^3 (g/dm^3).

A solution of sodium chloride, NaCl, with concentration 2.5 $g/100\ cm^3$ has 2.5 g of NaCl in 100 cm^3 of solution. So 200 cm^3 of this solution has 2 × 2.5 = 5.0 g of NaCl, and 1000 cm^3 (1.0 dm^3) has 10 × 2.5 g = 25.0 g of NaCl. The concentration is 25.0 g/dm^3.

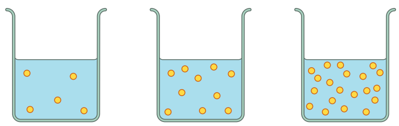


Figure 5.2.17 The concentration increases as the number of solute particles in a fixed volume increases

1. **Put the following solutions in order with the most concentrated first:**
   20 $g/100\ cm^3$ 20 $g/1000\ cm^3$ 8 $g/50\ cm^3$
2. **Calculate the concentration of the following solutions in g/dm^3.**
   3.2 g in 100 cm^3 3.2 g in 250 cm^3 6.4 g in 500 cm^3

KEY INFORMATION

1 dm^3 is commonly called 1 litre but we use 1 dm^3 when using units in scientific contexts. There are 1000 cm^3 in 1 dm^3.

Calculating the mass of solute

When you carry out reactions you often use a measured volume of solution. If you know the concentration of a solution and the volume of the solution, you can work out the mass of solute in that volume.

Example: A solution has a concentration of 30 g/dm^3. How much solute is in 150 cm^3?

There are 30 g in 1 dm^3, which is 1000 cm^3. So in 150 cm^3 the amount of solute is: $\frac{150\ cm^3}{1000\ cm^3} \times 30\ g = 4.5\ g$

So there are 4.5 g of solute in 150 cm^3.

3 A solution has a concentration of 4.2 g/dm^3. Calculate the mass of solute dissolved in 250 cm^3 of solution.

4 A solution has a concentration of 5.4 g/100 cm^3. Calculate the mass of solute dissolved in 35 cm^3 of solution.

HIGHER TIER ONLY

Relating mass, volume and concentration

Mass, volume and concentration in g/dm^3 are related by the equation:

$$\text{concentration (in g/dm}^3\text{)} = \frac{\text{mass of solute (in g)}}{\text{volume of solution (in dm}^3\text{)}}$$

For an experiment you might need to make up a solution with a given concentration. Using this equation you can work out how much solute is needed.

Example: A solution of 37.5 g/dm^3 of NaCl is needed. The volume of solution required is 200 cm^3. How much NaCl is needed?

Rearranging the equation for concentration:

mass of solute (in g) = concentration (in g/dm^3) × volume of solution (in dm^3)

$$\text{mass of NaCl} = 37.5\ g/dm^3 \times \frac{200}{1000}\ dm^3 = 7.5\ g$$

Concentrations can also be given as moles per dm^3 (mol/dm^3).

$$\text{concentration (in mol/dm}^3\text{)} = \frac{\text{amount of solute (in mol)}}{\text{volume of solution (in dm}^3\text{)}}$$

Dividing the mass of solute by the molar mass gives the number of moles.

Example: A solution contains 20 g of sodium hydroxide (NaOH) in 1 dm^3. What is the amount of NaOH in moles?

Relative atomic masses (A_r): Na = 23, O = 16, H = 1

The molar mass of NaOH = 40 g/mol, so the amount of NaOH in 20 g is $\frac{20\ g}{40\ g/mol} = 0.5$ mol

5 What would happen to the concentration of the solution if the mass of the solute added were increased?

6 Give two ways of decreasing the concentration.

7 A solution has a concentration of 86.0 g/dm^3. What volume of solution is needed to provide 21.5 g of solute?

5.2j Amounts in chemistry

Learning objectives:

- use relative atomic masses to calculate relative formula mass
- explain how relative formula mass relates to number of moles
- explain how number of moles relate to other quantities.

KEY WORDS

Avogadro constant
molar mass
mole

We have learned that there are different ways of measuring chemical quantities. Chemists measure amounts in *moles*. We have learned how each quantity relates to the mole. Let's draw these altogether here in one place.

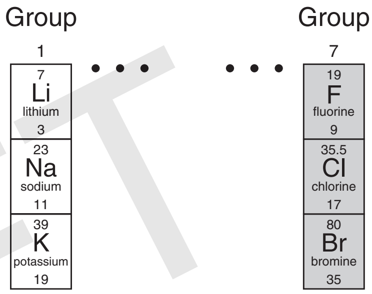


Figure 5.2.18 The periodic table shows the relative atomic mass of each element

Relative formula mass

The relative atomic mass of each element is written on the periodic table.

A relative formula mass is the sum of the relative atomic masses of elements in a substance.

A subscript number 3 by an element symbol, for example, shows that there are three of these atoms in the substance per atom of the other elements that don't have subscripts. So, in $CaCO_3$, there are three oxygen atoms for every one calcium atom and carbon atom.

The relative formula mass (M_r) is equal to the relative atomic masses (A_r) of all the atoms in the formula - in this case, one Ca atom, one C atom and three O atoms.

$$M_r \text{ of } CaCO_3 = 40 + 12 + (3 \times 16)$$
$$= 40 + 12 + 48$$
$$= 100$$

If there is more than one atom in a group attached to another atom a bracket is used, for example:

$$Ca(OH)_2$$

1 **Calculate the relative formula mass of magnesium sulfate $MgSO_4$.**

2 **Calculate the relative formula mass of**

a $Ca(NO_3)_2$

b $Al_2(SO_4)_3$.

HIGHER TIER ONLY

Relative formula masses and molar masses

The unit for chemical amounts is the **mole**. The mass of a mole of any substance in grams, its **molar mass**, is numerically equal to its relative formula mass. If we know the mass of a substance that reacts, then we can find the number of moles that react by dividing the mass by the molar mass.

$$\text{amount of substance (mol)} = \frac{\text{mass of substance (g)}}{\text{molar mass (g/mol)}}$$

The number of particles in 1 mol of a substance is equal to **Avogadro's constant** (6.02 x 10^{23} mol^{-1}). The ratio of moles in a reaction is the same as the ratio of particles, so we can use the numbers of moles reacting to write chemical equations. For example:

2 mol NaOH : 1 mol H_2SO_4

$2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$

3 **How many moles is 132 g of CO_2? (A_r C = 12, Ar O = 16)**

4 **Calculate the number of atoms that two moles of neon contains.**

Working with moles

The key to working with moles is the chemical equation. The balanced symbol equation indicates the number of moles that are involved in any reaction. For example:

$$ZnCO_3 \rightarrow ZnO + CO_2$$

1 mole of $ZnCO_3$ produces 1 mole of ZnO + 1 mole of CO_2.

From the number of moles and the molar masses, we can work out the masses of reactants and products.

Molar mass of $ZnCO_3$ is 125 g/mol, molar mass of ZnO is 81 g/mol, and molar mass of CO_2 is 44.

5 **How many grams of ZnO would be produced by 6.25 g $ZnCO_3$?**

6 **Calculate the mass of CO_2 produced when 660 g of propane, C_3H_8, is combusted. (A_r C= 12, A_r H = 1, A_r O = 16)**

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

MATHS SKILLS

5.2k Change the subject of an equation

Learning objectives:

- to use an equation to demonstrate conservation of mass
- to rearrange the subject of an equation
- to carry out a multi-step calculation

KEY WORDS

multiply
subject

Using the relationship between mass of substance (g), molar mass (g/mol) and amount of substance (mol), we can form equations for chemical reactions. Because mass is conserved, we can use these equations to work out unknown masses. Sometimes this requires multi-step calculations, in which we need to rearrange equations.

KEY INFORMATION

You will need to know how to change the subject of an equation by using all four operations: add, subtract, multiply and divide.

Conservation of mass

If we use a total mass of reactants in a chemical reaction the products made will be of the same total mass. This is known as the conservation of mass. In this reaction:

aluminium + sulfur → aluminium sulfide

$2Al + 3S \rightarrow Al_2S_3$

Given the relative atomic masses Ar: Al = 27 and S = 32, then 54 g of aluminium can be reacted with 96 g of sulfur to form 150 g of aluminium sulfide.

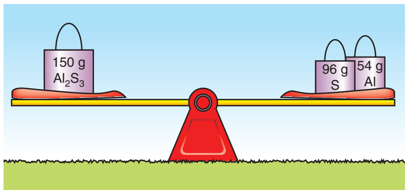


Figure 5.2.19

Changing the subject of an equation

Sometimes we use the principle of conservation of mass to find out the amount of reactant when given the amount of product(s). This involves changing the subject of the equation. In this reaction:

magnesium oxide	+ sulfuric acid	→	magnesium sulfate	+ water
MgO	+ H_2SO_4	→	$MgSO_4$	+ H_2O
40 g of MgO		will make	120 g of $MgSO_4$	+ 18 g of H_2O

Why do we get the bigger mass of product? It is because the mass of sulfuric acid has to be added in.

This is because:

mass of MgO + mass of H_2SO_4 = mass of $MgSO_4$ + mass of H_2O

What mass of sulfuric acid will be used?

40g + mass of H_2SO_4 = 120g + 18g

The **subject** of the equation needs to be the mass of sulfuric acid. In order to isolate the mass of the sulfuric acid, we need to subtract the mass of the MgO. In order to keep the equation balanced, we must also subtract 40 g from the right-hand side.

As the equation is balanced we need to do the *same to both sides* to move the 40 g.

Therefore we need to *subtract 40 g from both sides.*

$$(40g - 40g) + \text{mass of } H_2SO_4 = 20g + 18g - 40g$$
$$\text{mass of } H_2SO_4 = 98g$$

1 **In the reaction**

magnesium + oxygen → magnesium oxide

24 g of magnesium are reacted with 16 g of oxygen. Write a balanced equation and calculate how much magnesium oxide is produced in g.

2 **Use the equation: $MgCO_3 + H_2SO_4 \rightarrow MgSO_4 + H_2O + CO_2$ to find out how much $MgCO_3$ was needed to react with 49 g of H_2SO_4 to make 91 g of the total product. Write out your working to show how you change the subject of the equation.**

Calculating masses

Epsom bath salts are magnesium sulfate, which we can make in two ways:

$MgO + H_2SO_4 \rightarrow MgSO_4 + H_2O$ *or* $MgCO_3 + H_2SO_4 \rightarrow MgSO_4 + H_2O + CO_2$

We can use the relative atomic mass (A_r) to calculate the relative formula mass (M_r) of any of the substances in this equation and so find a ratio of how much reactant gives a quantity of product.

For example, A_r values: Mg = 24, O = 16, S = 32, H = 1 g
so: M_r: MgO = 40 M_r: $MgSO_4$ = 120 g

We can therefore say that 40 g of MgO will give 120 g of $MgSO_4$. So the ratio is expressed as:

$$\frac{\text{mass of MgO}}{\text{mass of MgSO}_4} = \frac{40\text{ g}}{120\text{ g}} = \frac{1}{3}$$

We want to find out how much $MgSO_4$ we can make using 150 g of MgO. To do this we need to change the subject of the ratio equation, keeping it balanced at every step.

$$\frac{150\text{ g}}{\text{mass of MgSO}_4} = \frac{1}{3}$$

Step 1: substitute the known value in the equation

$$\frac{150\text{ g} \times \text{mass of MgSO}_4}{\text{mass of MgSO}_4} = \frac{1 \times \text{mass of MgSO}_4}{3}$$

Step 2: multiply both sides by the mass of $MgSO_4$

$$\frac{150\text{ g} \times \cancel{\text{mass of MgSO}_4}}{\cancel{\text{mass of MgSO}_4}} = \frac{1 \times \text{mass of MgSO}_4}{3}$$

Step 3: cancel the mass of $MgSO_4$ on the left-hand side

$$3 \times 150\text{ g} = 3 \times \frac{\text{mass of MgSO}_4}{3}$$

Step 4: multiply both sides by 3

$$450\text{ g} = \frac{\cancel{3}}{\cancel{3}} \times \text{mass of MgSO}_4$$

Step 5: cancel the fraction on the right-hand side and work out the value of the left-hand side

$$450\text{ g} = \text{mass of MgSO}_4$$

Step 6: is this a reasonable answer? Yes.

DID YOU KNOW?

You can do steps 1–6 all at the same time by cross-multiplying the fractions at the start.

3 **What mass of $MgSO_4$ is produced by 62.5 g of MgO?**

4 **What mass of $MgSO_4$ is produced by 25.2 g of $MgCO_3$?**

Check your progress

You should be able to:

identify the elements in a compound using their symbols.	recognise compounds from their formula.	write equations for simple reactions including the physical states.
state the law of the conservation of mass.	explain how to balance equations in terms of numbers of atoms on both sides of the equation.	explain the meaning of subscripts within a formula and multipliers before a formula in a balanced equation.
explain any observed changes in mass in a chemical reaction.	identify the mass changes using a balanced symbol equation.	use ratio when writing and balancing equations.
be able to calculate a relative formula mass from the sum of the relative atomic masses.	calculate the sum of the relative formula masses of reactants and products.	show how the relative formula masses of reactants are equal to the relative formula masses of products.
explain that when there is a mass change in a reaction it may be because a gas is being given off.	explain why there appears to be a mass change when metal carbonates are heated or metals are heated in oxygen.	explain observed changes in mass in non-enclosed systems and explain the changes in terms of the particle model.
calculate the masses of substances in a balanced symbol equation.	calculate the masses of reactants and products from balanced symbol equations.	calculate the mass of a given reactant or product.
balance an equation given the masses of reactants and products.	change the subject of a mathematical equation.	perform a multi-step calculation.

Higher tier only

describe the measurement of amounts of substance in moles. relate concentration to mass and volume.	calculate the number of moles in a given mass. define the Avogadro constant.	calculate the mass of a given number of moles.

Worked example

1. a **What is the mass number of an atom?**

   The mass of the atom.

   b **Why is the relative atomic mass of chlorine 35.5?**

   Some chlorine atoms have different masses.

   c **What is the relative formula mass?**

   All of the elements in the formula added together.

2. a **Iron sulfate has the formula $FeSO_4$.**

   Relative atomic masses (A_r): Fe = 56, S = 32, O = 16.

   Calculate the relative formula mass (M_r) of iron sulfate ($FeSO_4$).

   Relative formula mass (M_r) of iron sulfate = 56 + 32 + 16 = 104

   b **Calculate the percentage of iron in iron sulfate ($FeSO_4$).**

   Fe is one of 3 elements in it so 33%.

3. **What mass of iron reacts with 10.0 g of sulfur?**

   $Fe(s) + S(s) \rightarrow FeS(s)$

   The equation shows that it is one mole to one mole, so 10.0 g iron.

This is incorrect. The mass number is the number of protons and the number of neutrons added together.

This is a good start, but the answer needs to show that you know that isotopes of the same element have different numbers of neutrons, so the relative atomic mass is an average of the atomic masses of that element.

It is more accurate to say that it is the relative atomic masses of each element of the compound added together.

This is a common mistake to make. The oxygen has to be counted 4 times, as the 4 is a multiplier. So it is 56 + 32 + (4 × 16) = 152.

The atomic mass of iron = 56. The total formula mass is 152. So the amount of iron in iron sulfate is 56/152 = 36.8%.

Remember that moles and grams are not the same thing. The amount of sulfur = 10.0 g/32 g/mol (where 32 = relative atomic mass) = 0.313 mol. Therefore the number of moles of iron = 0.313 mol (the same as the sulfur).

Therefore the mass of iron = 0.313 mol × 56 g/mol = 17.5 g.

End of chapter questions

Getting started

1. Calculate the relative formula mass of $NaNO_3$. (Relative atomic masses: Na = 23, N = 14, O = 16)

   Circle your answer. 53 113 85 — 1 Mark

2. Alex burns a strip of magnesium weighing 0.7 g in air. What will be the more likely mass after heating?

   A 0.7 g **B** 0.64 g **C** 0.76 g — 1 Mark

3. Sam heats 4.2 g of a substance. After heating the mass is 3.8 g. Suggest why. — 2 Marks

4. Explain why mass is conserved in this reaction. (Relative atomic masses: Na = 23, O = 16)

   $$2Na + O_2 \rightarrow 2Na_2O$$

   2 Marks

Going further

5. Calculate the relative formula mass of $Ca(OH)_2$. (Relative atomic masses: Ca = 40, O = 16, H = 1) — 1 Mark

6. Manisha heats 5 g of calcium carbonate.

   Calcium carbonate calcium oxide + carbon dioxide.

   When she measures the amount of calcium oxide, it is less than 3g. Explain what has happened to the mass of the substances involved. — 2 Marks

7. Each of a pair of isotopes have:

   A the same number of protons
   B the same number of neutrons **C** the same atomic mass — 1 Mark

More challenging

8. What is the amount, in moles, of Mg in a 10.0 g sample of the metal? — 2 Marks

9. What is the number of particles in one mole of a substance?

   A 6.02×10^{23} **B** 6.02×10^{4} **C** 6.02×10^{-9} — 1 Mark

Most demanding

10. Using the relative atomic masses on the periodic table at the back of this book, complete the table below and work out the molecular formula of the hydrocarbon C_nH_{2n+2}

Formula	Mass (g)	Molar mass (g/mol)	Amount (mol)
$MgCO_3$	42		
$CaCO_3$			0.2
$Ca(OH)_2$	148		
C_nH_{2n+2}	25		0.25

4 Marks