

1 Rutherford's planetary model

The Ancient Greeks believed that you could divide something in half and then in half again, and so on until you ended up with something so small that it couldn't be divided any further. They called these **building blocks**¹ of matter **atoms**, which means **indivisible**² in Greek. However, experiments carried out in the 19th and 20th centuries showed that atoms are made up of smaller particles. The atom is made up of three differently charged particles: the electron, the neutron and the proton.

	Electron	Proton	Neutron
Electrical charge	Negative (-) $-1.602 \cdot 10^{-19}$ C	Positive (+) $1.602 \cdot 10^{-19}$ C	Neutral 0 C
Mass	$9.109 \cdot 10^{-31}$ kg	$1.673 \cdot 10^{-27}$ kg	$1.675 \cdot 10^{-27}$ kg
Discovered by	J. J. Thomson, 1897	E. Goldstein, 1886	J. Chadwick, 1932

Over the years, different scientists have developed **models**³ to try to understand how these particles are arranged within the atom.

- **Dalton's atomic model** (1808) **postulated**⁴ that all matter was made of indivisible atoms.
- **Thomson's atomic model** (1904) postulated that all atoms look like a plum pudding: with negatively-charged electrons embedded in a positively-charged 'dough' or pudding.
- **Rutherford's atomic model** (1911), called the 'planetary' model, postulated that electrons (the planets) travel in orbits around a tiny positive nucleus (the sun). According to this model:
 - atoms are mostly empty. Their mass is concentrated in a central nucleus containing the protons and neutrons.
 - the nucleus is positively-charged.
 - the negatively-charged electrons are far away from this nucleus, and they revolve around it very quickly.

Rutherford's model explains quite a few observable facts, for example, how matter is charged and the formation of ions.

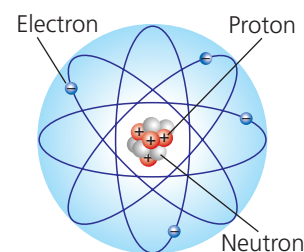


¹building block: basic unit which is necessary to make something else.

²indivisible: unable to be separated into parts.

³model: a simplified representation of an observed fact, used to explain what we see in experiments.

⁴postulate: assume or claim as true.



Rutherford's atomic model

CLIL activities

- 1 In your notebook, draw a diagram for each of the three atomic models. Work in pairs and discuss the differences and similarities between them.
- 2 Look at the table and answer the questions.
 - a. What's the charge of 5 electrons?
 - b. How many electrons are there in these?
 - i) -1 C
 - ii) -2 C
 - c. How many protons are there in:
 - i) $+1$ C
 - ii) $+5$ C
- 3 Listen and answer the questions.
 - a. Are atoms electrically neutral, positive or negative?
 - b. How do atoms become charged?
 - c. Explain ionisation.
 - d. What's an anion?
 - e. What's a cation?
- 4 With a classmate, write a paragraph summarising the explanation from activity 3. Read it out to the class.

4 Atomic spectra: a problem for Rutherford's model

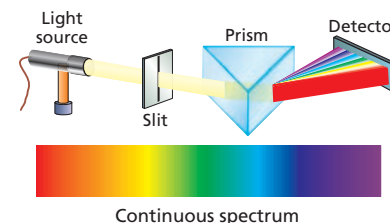
4.1. What are atomic spectra?

¹constituent: component.

²disperse: separate.

³fingerprint: lines on fingers that are unique for each human being.

When light goes through a glass prism, it's separated into its **constituent**¹ colours (the colours of the rainbow) and we obtain a continuous or full spectrum.



A **spectroscope** is a device that **dispersed**² light into its constituent colours and produces spectra.

Atomic spectra are discontinuous. Each element has its own unique atomic spectra, different from the atomic spectra of any other element. This is a bit like its **fingerprint**³.

- **Emission spectra:** when the atoms of an element are excited (due to receiving energy in the form of heat or an electrical current), they emit certain wavelengths of light, which correspond to different colours.
- **Absorption spectra:** when light goes through an element, the atoms absorb certain wavelengths of light, which correspond to certain colours. These will appear as black lines in the atomic spectra.

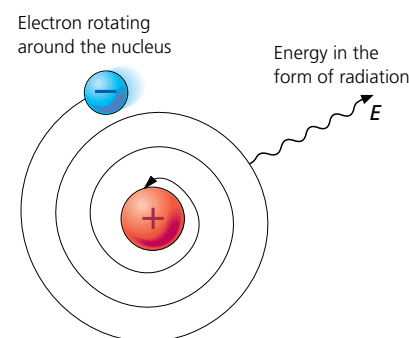
The black lines of the absorption spectrum coincide with the coloured lines in the emission spectrum.



4.2. Why's Rutherford's model no longer valid?

Rutherford's model cannot explain the discontinuous spectra. According to this model, the electrons spin around the nucleus at great speed.

However, any accelerated charged particle moving on a curved path like this would emit electromagnetic radiation. Thus, we should see a continuous spectrum. Furthermore, if the electron were constantly losing energy, it should get closer and closer to the nucleus, which isn't the case.



According to electromagnetic theory, any accelerated charged particle like an electron should emit continuous energy.

CLIL activities

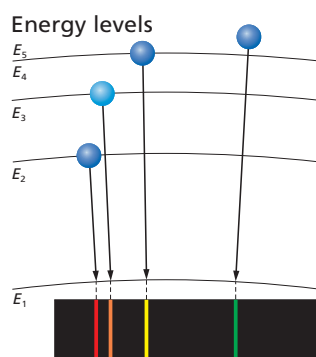
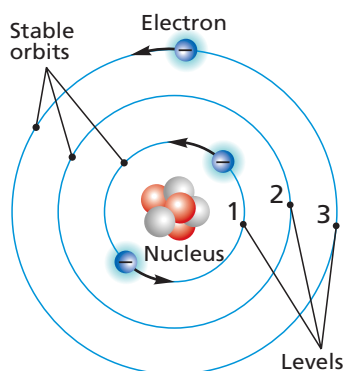
- 11 In your notebook, explain why atomic spectra are used to identify each element.
- 12 Why do the black lines of the absorption spectrum coincide with the coloured lines in the emission spectrum?

- 13 Listen to a presentation about different types of rainbows. Draw a picture of each rainbow.
 - a. lunar rainbow
 - b. twinned rainbows
 - c. fogbow rainbow

5 The atomic shell model

In 1913, Niels Bohr proposed a new model which explained the discontinuous spectra of the hydrogen atom.

- Electrons move in orbits at fixed distances from the nucleus, maintaining constant energy levels.
- Each orbit or **shell** has a different energy level identified by a number n ($n = 1, 2, 3\dots$). The further their orbit is from the nucleus, the more energy electrons will have.
- Electrons will emit energy in the form of radiation (light) when moving from a higher energy level to a lower one, and absorb it when jumping from a lower energy level to a higher one. In this way, if an electron moves from level 5 to level 1, it will emit the difference in energy between the two energy levels: $E_{\text{emit}} = E_5 - E_1$. This explains the discontinuous atomic spectra observed. As the energy levels for each atom are different, so are the emitted energies and therefore, the lines in its emission spectrum.



Each energy level can only accommodate a maximum of $2n^2$ electrons, where $n = 1, 2, 3, 4\dots$. There are also subshells, which can hold a maximum number of electrons, shown in brackets: **s** (2), **p** (6), **d** (10) and **f** (14).

The electron configuration of an element shows how electrons are distributed in an atom. The electrons occupy the different available shells, from lower to higher energy, taking into account the number of electrons that fit in each of the shells. The electrons in the outermost shell of an atom are called **valence electrons**².

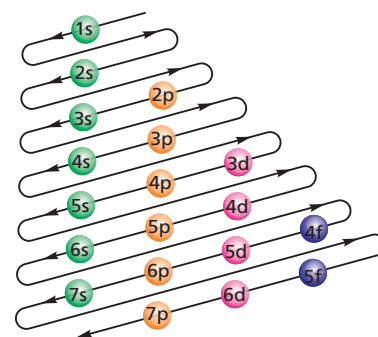


¹**shell**: orbit or level.

²**valence electrons**: electron in the outermost shell of an atom.



n	Sublevel	Maximum number of electrons per sublevel and per level ($2n^2$)	
1	s	2	
2	s	2	8
	p	6	
3	s	2	18
	p	6	
	d	10	
4	s	2	32
	p	6	
	d	10	
	f	14	



Möller's diagram shows in which order the electrons fill the different subshells.

CLIL activities

- In your notebook, copy and choose the correct words to complete the summary of Bohr's model.
 - Electrons orbit the nucleus at *fixed/variable* energy levels.
 - Electrons orbiting further away from the nucleus have *lower/higher* energy levels.
 - When electrons move to a lower energy level, they *absorb/emit* energy.
 - When electrons move to a higher energy level, they *absorb/emit* energy.
- Listen to the interview and answer the questions.
 - What are the lines found in the Sun's spectrum called?
 - Why is Joseph Fraunhofer important?
 - Which are the two most common elements found in all stars?
- Write the electron configuration of these elements. Compare your answers with a classmate. H ($Z=1$), Li ($Z=3$), O ($Z=8$), Fe ($Z=26$) and Ag ($Z=47$)

6 Classification of chemical elements

6.1. The periodic table

Mendeleev was one of the first scientists to organise chemical elements in a table in order of increasing atomic number. By doing so, he found a recurrent **pattern**¹ in their properties depending on their position. Today's periodic table has 118 chemical elements, arranged as follows.

- **Groups:** there are 18 columns (groups) from left to right. All the elements in a group have the same number of valence electrons, and thus they have similar chemical properties.
- **Periods:** there are seven rows (periods) from top to bottom. All elements in a row have the same number of electron shells and their properties change progressively from left to right.

6.2. Patterns found in periods and groups

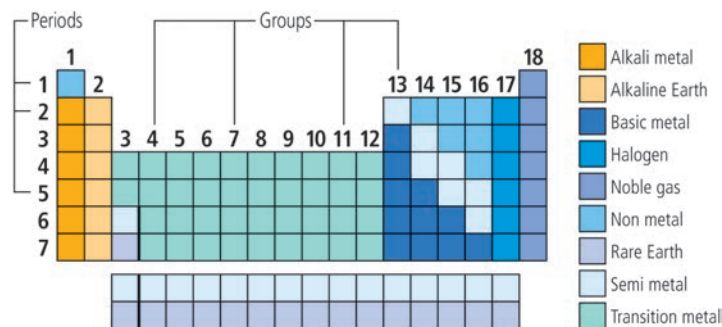
Patterns in groups: when we go down in a group, we find that for all the elements:

- the number of valence electrons is the same.
- the atomic mass increases.
- the **boiling point**² and **melting point**³ change slightly.
- **metallic properties**⁴ increase.

Some groups of elements (sometimes called families) have special names, as you can see in the table below, for example, Alkali metals (G1).

Patterns in periods: when we go from left to right in a period, we find that:

- the number of electrons in the outer shell increases by one from one group to the next.
- the atomic mass usually increases.
- metallic properties decrease.
- boiling and melting points increase up until the middle of the period and then decrease from there onwards.

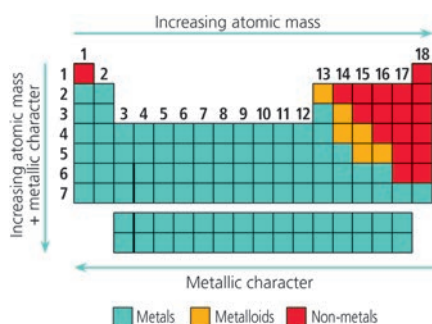


¹**pattern:** repeated characteristic.

²**boiling point:** temperature at which a substance starts to boil (turning from liquid to gas).

³**melting point:** temperature at which a substance starts to melt (turning from solid to liquid).

⁴**metallic properties:** physical properties associated with metals, such as conductivity.



CLIL activities

17 In your notebook, write which of these elements have the same number of valence electrons:

K, I, Li, Ag, Ca, Cu, Sr, Cl

18 Listen to the teacher talking about the history of the periodic table. What do these dates refer to?

1669 1886 1894 1945

19 For each pair, write which is more metallic.

- magnesium, strontium
- chlorine, bromine
- silicon, lead
- silver, gold

PERIODIC TABLE OF THE ELEMENTS

Representative elements

Representative elements

Group	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	
Period	I A	II A	III B	IV B	V B	VI B	VII B	VIII	IX	X	IB	II B	III A	IV A	V A	VIA	VII A	VIII A	
1	1 H Hydrogen 1.01																	2 He Helium 4.00	
2	3 Li Lithium 6.94	4 Be Beryllium 9.01											5 B Boron 10.81	6 C Carbon 12.01	7 N Nitrogen 14.01	8 O Oxygen 15.99	9 F Fluorine 19.00	10 Ne Neon 20.18	
3	11 Na Sodium 23.0	12 Mg Magnesium 24.31											13 Al Aluminium 26.98	14 Si Silicon 28.09	15 P Phosphorous 30.97	16 S Sulphur 32.06	17 Cl Chlorine 35.5	18 Ar Argon 39.95	
4	19 K Potassium 39.10	20 Ca Calcium 40.08	21 Sc Scandium 44.96	22 Ti Titanium 47.90	23 V Vanadium 50.94	24 Cr Chromium 51.99	25 Mn Manganese 54.94	26 Fe Iron 55.85	27 Co Cobalt 58.93	28 Ni Nickel 58.71	29 Cu Copper 63.54	30 Zn Zinc 65.37	31 Ga Gallium 69.72	32 Ge Germanium 72.59	33 As Arsenic 74.92	34 Se Selenium 78.97	35 Br Bromine 79.91	36 Kr Krypton 83.80	
5	37 Rb Rubidium 85.47	38 Sr Strontium 87.62	39 Y Yttrium 88.91	40 Zr Zirconium 91.22	41 Nb Niobium 92.91	42 Mo Molybdenum 95.95	43 Tc* Technetium (99)	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.91	46 Pd Palladium 106.4	47 Ag Silver 107.87	48 Cd Cadmium 112.41	49 In Indium 114.82	50 Sn Tin 118.69	51 Sb Antimony 121.75	52 Te Tellurium 127.60	53 I Iodine 126.90	54 Xe Xenon 131.30	
6	55 Cs Caesium 132.90	56 Ba Barium 137.34	Lanthanoids		72 Hf Hafnium 178.49	73 Ta Tantalum 180.95	74 W Tungsten 183.85	75 Re Rhenium 186.20	76 Os Osmium 190.20	77 Ir Iridium 192.22	78 Pt Platinum 195.09	79 Au Gold 196.97	80 Hg Mercury 200.59	81 Tl Thallium 204.37	82 Pb Lead 207.19	83 Bi Bismuth 208.98	84 Po Polonium (209)	85 At Astatine (210)	86 Rn Radon (222)
7	87 Fr Francium (223)	88 Ra Radium (226)	Actinoids		104 Rf* Rutherfordium	105 Db* Dubnium	106 Sg* Seaborgium	107 Bh* Bohrium	108 Hs* Hassium	109 Mt* Meitnerium	110 Ds* Darmstadtium	111 Rg* Roentgenium	112 Cn* Copernicium	113 Uut* Ununtrium	114 Fl* Flerovium	115 Uup* Ununpentium	116 Lv* Livermorium	117 Uus* Ununseptium	118 Uuo* Ununoctium

* Synthetic elements.

() The numbers in brackets represent the atomic mass of the most stable isotope of the element.

Atomic number	Atomic mass
Symbol	Name

 Metals	 Metalloids
 Non-metals	

6.3. The relationship between the electron configuration of an element and its position in the periodic table

'block': group in which each element is next to the others.

We can work out the electron configuration of an element from its position in the periodic table.

We can see that:

- all elements in groups 1 and 2 have their last electron in ns .
- all elements in group 3 to group 12 have their last electron in shell nd .
- all elements in groups 13 to 18 (except for He) have their last electron in shell np .

So the different **blocks**¹ in the periodic table depend on the number of electrons in their outer shell.

1st period

H, $Z = 1 \rightarrow 1s^1$. The only electron is in subshell $1s$.

He, $Z = 2 \rightarrow 1s^2$. Shell $1s$ is full.

2nd period

Li, $Z = 3 \rightarrow 1s^2 2s^1$. The last electron is in subshell $2s$. Continuing to the right, the electrons fill the succeeding subshells until we reach the last element of this period, Ne.

Ne, $Z = 10 \rightarrow 1s^2 2s^2 2p^6$. Shell 2 is full.

3rd period

Na, $Z = 11 \rightarrow 1s^2 2s^2 2p^6 3s^1$. The last electron is in subshell $3s$. Continuing to the right, the electrons start filling subshell $3p$, until the last element of this period, Ar.

Ar, $Z = 18 \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^6$. Shell 3 is full.

	1																		18	
1	1s	2																		
2	2s																			
3	3s		3	4	5	6	7	8	9	10	11	12		2p						
4	4s		3d											4p						
5	5s		4d											5p						
6	6s		5d											6p						
7	7s		6d											7p						
			4f																	
			5f																	

- S-block
- D-block
- P-block
- F-block

4th period

K, $Z = 19 \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$. The last electron is in subshell $4s$. If we continue to the right, after Ca ($\dots 4s^2$), the electrons start filling up subshell $3d$. This is followed by subshell $4p$.

Kr, $Z = 36 \rightarrow [Ar] 4s^2 3d^{10} 4p^6$. Shell 4 is full.

7th period

Fr, $Z = 87 \rightarrow Fr = [Rn]7s^1$. The last electron is in subshell $7s$. From Th ($Z = 90$) to Lr ($Z = 103$) are known as the **actinides** and they keep filling subshell $5f$.

Og ($Z = 118$) $\rightarrow [Rn] 5f^{14} 6d^{10} 7s^2 7p^6$. Shell 7 gets completed.

6th period

Cs, $Z = 55 \rightarrow Cs = [Xe]6s^1$. The last electron is in subshell $6s$. From Ce ($Z = 58$) to Lu ($Z = 71$) are known as the **lanthanides** and they keep filling subshell $4f$.

Rn, $Z = 86 \rightarrow [Xe] 4f^{14} 5d^{10} 6s^2 6p^6$. Shell 6 gets completed.

5th period

Rb, $Z = 37 \rightarrow Rb = [Kr] 5s^1$. The last electron is in subshell $5s$. Then, from Y to Cd, subshell $4d$ gets completed. From In to Xe, subshell $5p$ also gets completed.

Xe, $Z = 54 \rightarrow [Kr] 4d^{10} 5s^2 5p^6$. Shell 5 gets completed.

CLIL activities

20 In your notebook, write the electron configuration of: Kr ($Z=36$) and Xe ($Z=54$). Why does the configuration of Xe appear as $[Kr]4d^{10} 5s^2 5p^6$?

21 Name five elements that have the same number of electrons in their last shell as:

- a. Si ($Z = 14$) b. Cl ($Z = 17$)

22 Listen to a mother helping her son with his project on synthetic elements.

- Note down the questions the boy asks.
- Discuss with a classmate what the answers to his questions are.
- Listen again to check.

7 Types of chemical elements

We can classify elements into types based on their electron configuration.

7.1. The representative elements (main group)

These are elements whose valence electrons occupy only shells s (groups 1 and 2) and p (groups 13 to 18), that is: from ns^1 to ns^2np^6 .

They're the most abundant elements on Earth and in the Universe.

- the s-block → groups 1 and 2
- the p-block → groups 13 to 17
- the noble gases → group 18 are extremely stable elements.

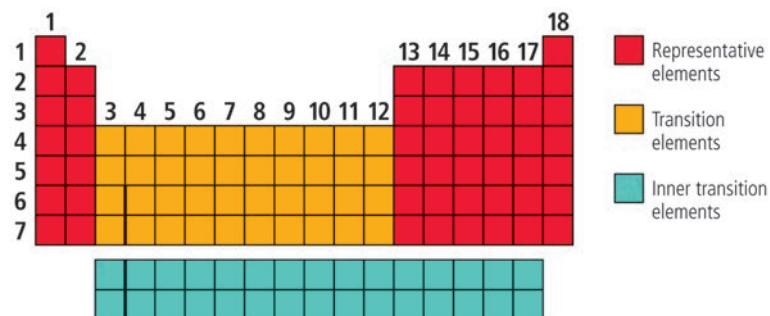
7.2. Metals, metalloids and non-metals

Metals conduct electricity because they have few electrons in their outer shell, which are easily separated from the atom. Non-metals have many electrons in their outer shell, which are strongly tied to the atom, so the electrons can't move freely. This is why non-metals are poor conductors. Metalloids are in between these two groups and are semi-conductors.

7.3. Transition elements

These are elements whose valence electron is in subshell d. They're in the middle of the periodic table (groups 3 to 12) and share the following properties:

- they're all metals.
- they're solid, except for mercury, which is liquid.
- they're very hard.
- their boiling and melting points are very high.
- they're good conductors of heat and electricity.
- they usually form coloured compounds.



7.4. Inner transition elements

The inner transition elements are elements divided into two blocks of 14 elements: the lanthanoids for subshell 4f and the actinoids for subshell 5f. Their properties are similar to those of the transition elements. All actinoids are **radioactive**¹ by nature. Those from $Z = 92$ to $Z = 103$ are also **synthetic**².

CLIL activities

23 Listen and choose the correct option.

- An element can become radioactive if the *nucleus/electrons* aren't stable.
- There are *two/three* types of particles that radioactive elements can emit.
- Half-life is the time it takes for the radioactivity to be *reduced by half/disappear*.

24 Nuclear power stations use the isotope ^{235}U as fuel. Use the Internet to research the advantages and disadvantages of this fuel and make a presentation including examples, such as Chernobyl or Fukushima.

25 Work in with a classmate. Choose an element. Try to guess your partner's element by asking questions about its characteristics, such as its boiling point, if it's a metal or not and its s-block.

¹**radioactive**: not stable, it decays into other elements.

²**synthetic**: not found naturally on Earth; artificial.

