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

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

[2]**indivisible:** unable to be separated into parts.

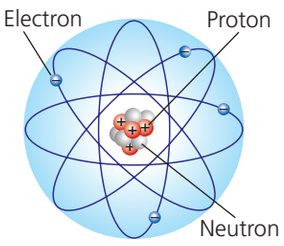
[3]**model:** a simplified representation of an observed fact, used to explain what we see in experiments.

[4]**postulate:** assume or claim as true.

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.



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.**

2 How can we identify the atoms of an element?

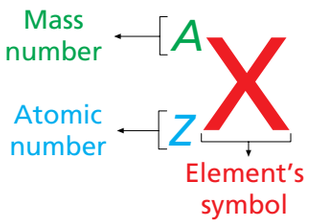
[1]**isotopes:** different forms of the same element that contain the same number of protons, but different number of neutrons in their nuclei.

All atoms in any given element are the same because they all have the same number of protons. The number of protons determines which kind of element an atom is. For example, an oxygen atom has eight protons. If the number of protons changes, for example to seven, it's no longer an oxygen atom, it's a nitrogen atom.

The atomic number (*Z*) is the number of protons in an atom. It's a whole number and is indicated as subscript in front of the element's symbol. For example: $_{11}Na$, $_{8}O$.

The mass number (*A*) is the total number of protons and neutrons in an atom. It's a whole number and is indicated as superscript in front of the element's symbol. For example: ^{23}Na, ^{16}O.

Each atom is identified by its atomic and mass numbers, which give information about the number of particles that are in its nucleus.



Number of protons = Z

Number of neutrons = $A - Z$

Number of electrons (for a neutral atom) = number of protons = Z

All atoms of a given element have the same atomic number (the same number of protons). For example, all hydrogen atoms have 1 proton (Z=1).

As atoms of different elements have different numbers of protons (different atomic numbers), we use Z to identify each element.

However, not all atoms of a given element necessarily have the same mass number; they can have the same number of protons, but a different number of neutrons. We call them **isotopes**[1]. For example, all hydrogen atoms have one proton (Z = 1), but hydrogen has three naturally-occuring isotopes:

- with 0 neutrons (A = 1).
- with one neutron (A = 2).
- with two neutrons (A = 3).

CLIL activities

5 **For the two carbon isotopes $^{12}_{6}C$ and $^{14}_{6}C$, write in your notebook:**

a. the mass number and the atomic number.

b. the number of protons, neutrons and electrons.

6 **Listen to a physicist giving a talk about hydrogen and answer the questions.**

a. Name the three different isotopes of hydrogen.

b. Explain why their mass numbers are 1, 2 and 3, respectively.

c. Which is the most common isotope?

7 **Work out the number of protons, neutrons and electrons for these elements. Then compare your answers with a classmate.**

a. $^{35}_{17}Cl^{-}$

b. $^{16}_{8}O^{2-}$

c. $^{31}_{70}Ga^{3+}$

d. $^{14}_{28}Si^{4+}$

e. $^{7}_{3}Li^{1+}$

f. $^{24}_{12}Mg^{2+}$

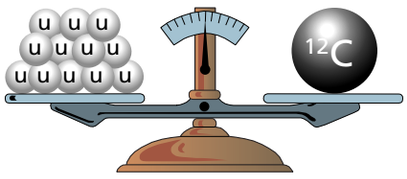
g. $^{32}_{16}S^{2-}$

3 How can we calculate the mass of an atom? The atomic mass (A_r)

The atomic mass of an element is the mass of one of its atoms. It's expressed in unified atomic mass units (u).

The **unified atomic mass unit** (u) is $\frac{1}{12}$ of the mass of a carbon atom ^{12}C.

The **relative atomic mass (A_r)** or simply the **atomic mass** of an element is the mass of one atom of that element compared to u, that is, to $\frac{1}{12}$ of the mass of ^{12}C, and it's **dimensionless**[1]. For example, the relative atomic mass of oxygen is 16. This means that its mass is 16 times the mass of $\frac{1}{12}$ of the mass of ^{12}C.



We need 12 units of mass to balance one ^{12}C atom.

How can we calculate the relative atomic mass of an element?

1. **Monoisotopic elements:** If the element has only one type of atom, its relative atomic mass is the atomic mass of the atom (*A*). For example, sodium has only one naturally-occurring isotope, ^{23}Na, thus its relative atomic mass is 23.

2. **Elements with different isotopes:** the relative atomic mass of the element is the **weighted average**[2] of the relative atomic masses of the isotopes. Therefore, it can't be a whole number.

   $$A_r = m_1 \cdot \%\ \text{isotope}_1 + m_2 \cdot \%\ \text{isotope}_2 + \ldots$$

   where m_1 and m_2 are the relative atomic masses of the isotopes of element *A* and % isotope is the **abundance**[3] of that isotope compared to the other isotopes of that element.

   Example: there are three oxygen isotopes with the following abundances: ^{16}O (99.759%), ^{17}O (0.037%) and ^{18}O (0.204%). Therefore, the relative atomic mass of oxygen is:

   $$A_r = \frac{99.759 \cdot 16}{100} + \frac{0{,}037 \cdot 17}{100} + \frac{0{,}204 \cdot 18}{100} = 16.004$$

[1]**dimensionless:** has no physical dimensions or units.

[2]**weighted average:** average of all values calculated taking into account the importance (or abundance) of each of them.

[3]**abundance:** the ratio of something compared to something else, expressed as a percentage.

CLIL activities

8 **In your notebook, work out the following: an imaginary element *X* has only two isotopes; *A*, with an atomic mass of 52, and *B* with 53. If the abundance of *A* in nature is 42%, what's the relative atomic mass of this imaginary element?**

9 **The relative atomic mass of Rubidium is 85.47. It has two main isotopes: ^{85}Rb and ^{87}Rb. Calculate the abundance of each of these isotopes.**

10 **Listen to the radio interview. Copy and complete the sentences about carbon-14 dating.**

a. Carbon-14 dating is a way of...
b. When a living thing dies, carbon-14 starts to...
c. We can use this dating method for a maximum time period of...
d. After this time period, all the...

4 Atomic spectra: a problem for Rutherford's model

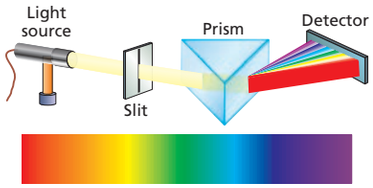
4.1. What are atomic spectra?

[1]**constituent:** component.

[2]**disperse:** separate.

[3]**fingerprint:** 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.



Continuous spectrum

A **spectroscope** is a device that **disperses**[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.

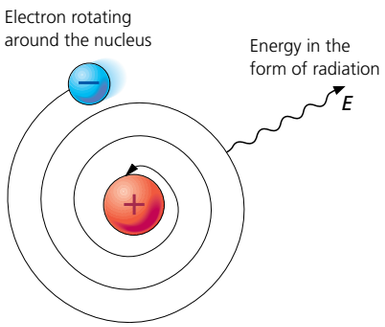

Absorption spectrum


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?

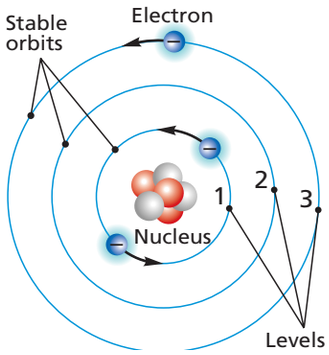
13 Listen to a presentation about different types of rainbows. Draw a picture of each rainbow.

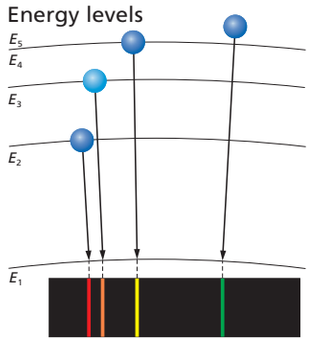
a. lunar rainbow
b. twinned rainbows
c. fogbow rainbow

5 The atomic shell model

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

1. Electrons move in orbits at fixed distances from the nucleus, maintaining constant energy levels.
2. Each orbit or **shell**[1] has a different energy level identified by a number *n* (*n* = 1, 2, 3...). The further their orbit is from the nucleus, the more energy electrons will have.
3. 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: $E_{emit} = E_5 - E_1$. 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.





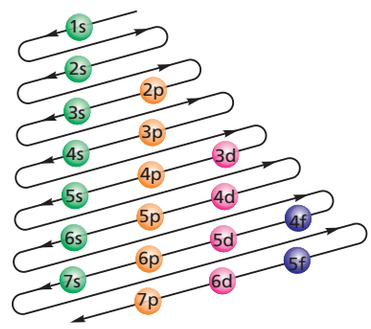
Each energy level can only accommodate a maximum of $2n^2$ electrons, where *n* = 1, 2, 3, 4... . There are also subshells, which can hold a maximum number of electrons, shown in brackets: **s** (2), **p** (6), **d** (10) and **f** (14).

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	



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

CLIL activities

14 **In your notebook, copy and choose the correct words to complete the summary of Bohr's model.**

a. Electrons orbit the nucleus at *fixed/variable* energy levels.

b. Electrons orbiting further away from the nucleus have *lower/higher* energy levels.

c. When electrons move to a lower energy level, they *absorb/emit* energy.

d. When electrons move to a higher energy level, they *absorb/emit* energy.

15 **Listen to the interview and answer the questions.**

a. What are the lines found in the Sun's spectrum called?

b. Why is Joseph Frauenhofer important?

c. Which are the two most common elements found in all stars?

16 **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.

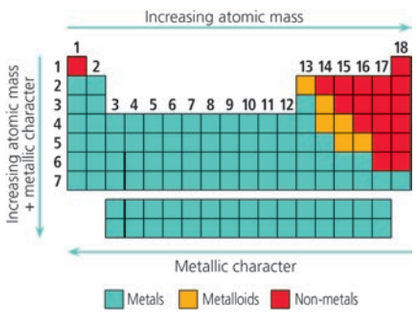
- **Groups:** there are 18 columns (groups) from left to right. All the elements in a group have the same number of valence electrons, and thus they have similar chemical properties.
- **Periods:** there are seven rows (periods) from top to bottom. All elements in a row have the same number of electron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

[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.



6.2. Patterns found in periods and groups

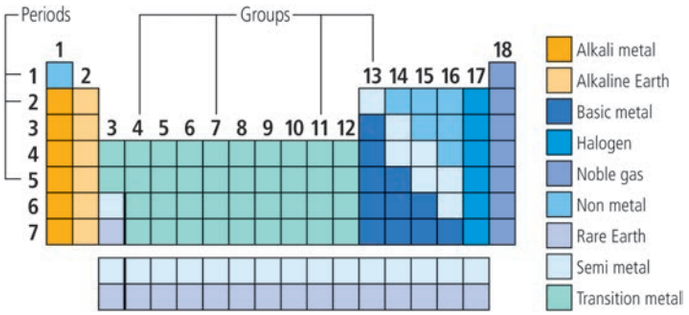
Patterns in groups: when we go down in a group, we find that for all the elements:

- 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.



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.**

a. magesium, strontium
b. chlorine, bromine
c. silicon, lead
d. silver, gold

PERIODIC TABLE OF THE ELEMENTS

Atomic number | Atomic mass
Symbol
Name

Metals
Non-metals
Metalloids

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

Period	Group 1 (Representative elements)	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7	Group 8	Group 9	Group 10	Group 11	Group 12	Group 13 (Representative elements)	Group 14	Group 15	Group 16	Group 17	Group 18
	I A	II A	III B	IV B	V B	VI B	VII B	VIII			I B	II B	III A	IV A	V A	VI A	VII A	VIII A
1	1 1.01 **H** Hydrogen																	2 4.00 **He** Helium
2	3 6.94 **Li** Lithium	4 9.01 **Be** Beryllium											5 10.81 **B** Boron	6 12.0 **C** Carbon	7 14.01 **N** Nitrogen	8 15.99 **O** Oxygen	9 19.0 **F** Fluorine	10 20.18 **Ne** Neon
3	11 23.0 **Na** Sodium	12 24.31 **Mg** Magnesium											13 26.98 **Al** Aluminium	14 28.09 **Si** Silicon	15 30.97 **P** Phosphorous	16 32.06 **S** Sulphur	17 35.5 **Cl** Chlorine	18 39.95 **Ar** Argon
4	19 39.10 **K** Potassium	20 40.08 **Ca** Calcium	21 44.96 **Sc** Scandium	22 47.90 **Ti** Titanium	23 50.94 **V** Vanadium	24 51.99 **Cr** Chromium	25 54.94 **Mn** Manganese	26 55.85 **Fe** Iron	27 58.93 **Co** Cobalt	28 58.71 **Ni** Nickel	29 63.54 **Cu** Copper	30 65.37 **Zn** Zinc	31 69.72 **Ga** Gallium	32 72.59 **Ge** Germanium	33 74.92 **As** Arsenic	34 78.97 **Se** Selenium	35 79.91 **Br** Bromine	36 83.80 **Kr** Krypton
5	37 85.47 **Rb** Rubidium	38 87.62 **Sr** Strontium	39 88.91 **Y** Yttrium	40 91.22 **Zr** Zirconium	41 92,91 **Nb** Niobium	42 95.95 **Mo** Molybdenum	43 (99) **Tc*** Technetium	44 101.07 **Ru** Ruthenium	45 102.91 **Rh** Rhodium	46 106.4 **Pd** Palladium	47 107,87 **Ag** Silver	48 112.41 **Cd** Cadmium	49 114.82 **In** Indium	50 118.69 **Sn** Tin	51 121.75 **Sb** Antimony	52 127.60 **Te** Tellurium	53 126.90 **I** Iodine	54 131.30 **Xe** Xenon
6	55 132.90 **Cs** Caesium	56 137.34 **Ba** Barium	Lantha-noids	72 178.49 **Hf** Hafnium	73 180.95 **Ta** Tantalum	74 183.85 **W** Tungsten	75 186.20 **Re** Rhenium	76 190.20 **Os** Osmium	77 192.20 **Ir** Iridium	78 195.09 **Pt** Platinum	79 196.97 **Au** Gold	80 200.59 **Hg** Mercury	81 204.37 **Tl** Thallium	82 207.19 **Pb** Lead	83 208.98 **Bi** Bismuth	84 (209) **Po** Polonium	85 (210) **At** Astatine	86 (222) **Rn** Radon
7	87 (223) **Fr** Francium	88 (226) **Ra** Radium	Actinoids	104 (267) **Rf*** Rutherfordium	105 (268) **Db*** Dubnium	106 (271) **Sg*** Seaborgium	107 (270) **Bh*** Bohrium	108 (277) **Hs*** Hassium	109 (276) **Mt*** Meitnerium	110 (281) **Ds*** Darmstadium	111 (282) **Rg*** Roentgenium	112 (285) **Cn*** Copernicium	113 (285) **Uut*** Ununtrium	114 (289) **Fl*** Flerovium	115 (289) **Uup*** Ununpentium	116 (291) **Lv*** Livermorium	117 (294) **Uus*** Ununseptium	118 (294) **Uuo*** Ununoctium

57 138.91 **La** Lanthanum	58 140.12 **Ce** Cerium	59 140.91 **Pr** Praseodymium	60 144.24 **Nd** Neodymium	61 145 **Pm*** Promethium	62 150.35 **Sm** Samarium	63 151.96 **Eu** Europium	64 157.25 **Gd** Gadolinium	65 158.92 **Tb** Terbium	66 162.50 **Dy** Dyprosium	67 164.93 **Ho** Holmium	68 167.26 **Er** Erbium	69 144.24 **Tm** Thulium	70 173.04 **Yb** Ytterbium	71 174.97 **Lu** Lutetium
89 (227) **Ac** Actinium	90 232.04 **Th** Thorium	91 (231) **Pa** Protactinium	92 238.03 **U** Uranium	93 (237) **Np*** Neptunium	94 (244) **Pu** Plutonium	95 (243) **Am*** Americium	96 (247) **Cm*** Curium	97 (247) **Bk*** Berkelium	98 (251) **Cf*** Californium	99 (252) **Es*** Einsteinium	100 (257) **Fm*** Fermium	101 (258) **Md*** Mendelevium	102 (259) **No*** Nobelium	103 (260) **Lr*** Lawrencium

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

[1]**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 *n*s.
- all elements in group 3 to group 12 have their last electron in shell *n*d.
- all elements in groups 13 to 18 (except for He) have their last electron in shell *n*p.

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

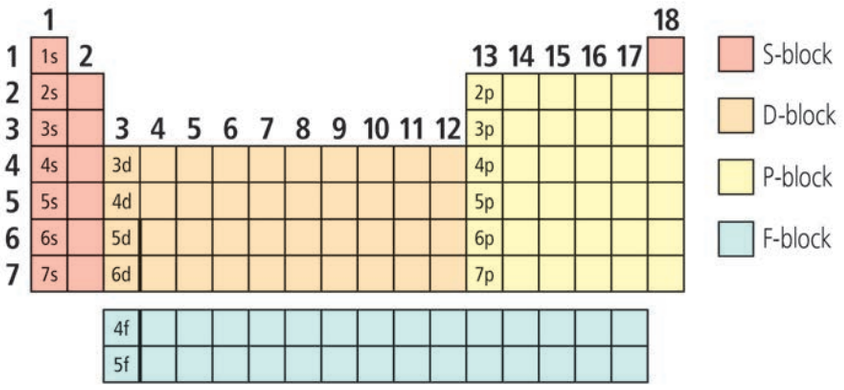
Li, $Z = 3 \rightarrow 1s^2\ 2s^1$. The last electron is in subshell 2s. Continuing to the right, the electrons fill the succeeding subshells until we reach the last element of this period, Ne.

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.



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 (...$4s^2$), the electrons start filling up subshell 3d. This is followed by subshell 4p.

Kr, Z 36 = → [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 ($Z = 118$) → [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] $5s^2$. 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]4d10 5s2 5p6?

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.

a. Note down the questions the boy asks.

b. Discuss with a classmate what the answers to his questions are.

c. 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.

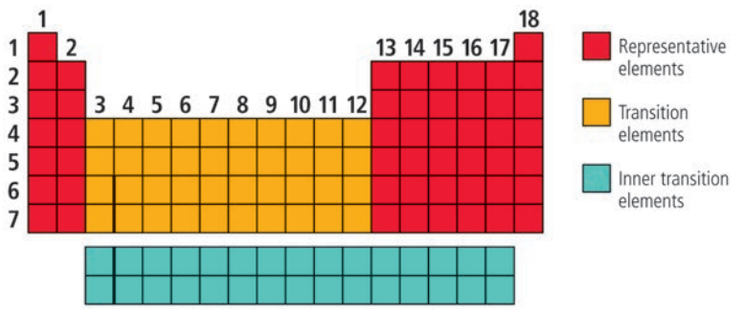
7.2. Metals, metalloids and non-metals

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].

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

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

CLIL activities

23 **Listen and choose the correct option.**

a. An element can become radioactive if the *nucleus/electrons* aren't stable.

b. There are *two/three* types of particles that radioactive elements can emit.

c. Half-life is the time it takes for the radioactivity to *be reduced by half/disappear.*

24 TASK STEAM **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.**