

ESAT

Chemistry

Atomic Structure and the Periodic Table

Free lesson notes

Covers	C1.1 to C1.6 and C2.1 to C2.5
Format	Lesson, worked examples, self-check
Price	Free

About this lesson

Almost all of chemistry follows from one fact: every atom is a tiny positive nucleus surrounded by electrons arranged in shells, and the way the outer shell is filled decides how that element behaves. The Periodic Table is simply a map of those arrangements, which is why it lets you predict properties of elements you have never met. In ESAT Chemistry this topic gives you quick marks (counting particles, reading a mass spectrum, writing an electron configuration, placing an element in the Table) and it also supplies the ideas you need for bonding, reactivity and quantitative chemistry.

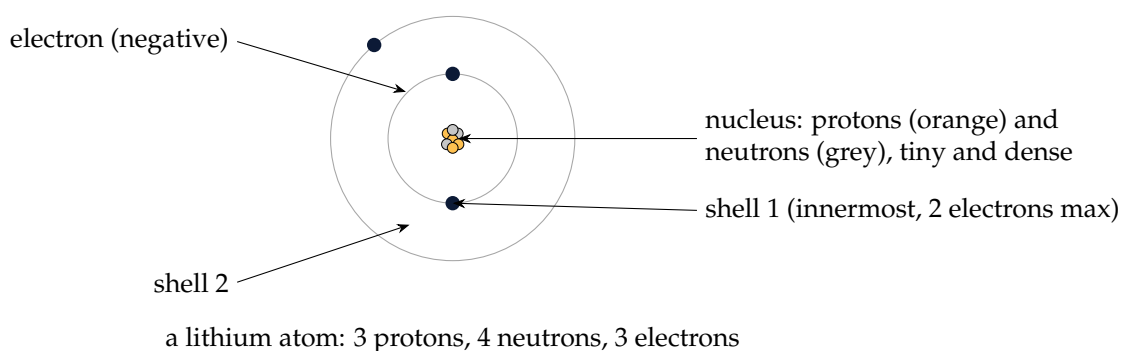
This lesson starts inside the atom and defines each particle. It then explains atomic number, mass number and isotopes, shows how a mass spectrometer detects isotopes and how to turn the data into a relative atomic mass by a weighted average, and builds the electron configurations of the first twenty elements one electron at a time. After that we explain how the Periodic Table is organised and *why*: the group tells you the outer electrons and the period tells you the number of shells. We finish with the trends in Group 1, Group 17 and Group 18, each explained from the electrons. Relative atomic masses are given in the exam, so there is no need to memorise them. Specification codes: C1.1–C1.6 (atomic structure) and C2.1–C2.5 (the Periodic Table).

How to use it. Work every example on paper before reading the answer, and draw the shell diagram whenever an electron configuration is involved. Finish with the self-check and the practice set.

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1 Inside the atom

An atom has a nucleus and shells. Every atom has a **nucleus** at its centre, containing **protons** and **neutrons** (together called *nucleons*). Around the nucleus are **electrons** that move in regions called **shells** (or energy levels). The nucleus is extremely small compared with the whole atom, about 10^{-15} m across against about 10^{-10} m for the atom, so almost all of the atom is empty space; yet almost all of the *mass* is in the nucleus.



The three particles. We describe each particle by a *relative charge* (the proton's charge is taken as +1) and a *relative mass* (the proton's mass is taken as 1), because the actual values are inconveniently tiny.

particle	relative charge	relative mass	where
proton	+1	1	nucleus
neutron	0	1	nucleus
electron	-1	$\frac{1}{1836}$ (negligible)	shells

Atoms are neutral. The charge of an electron is exactly equal and opposite to that of a proton, and an atom contains equal numbers of each, so the charges cancel: number of protons = number of electrons in a neutral atom. An electron is

about 1800 times lighter than a proton or neutron, so the mass of an atom is, to a very good approximation, the mass of its protons and neutrons.

Worked example: a particle audit

An atom has 6 protons, 6 neutrons and 6 electrons. What is its total charge and its total relative mass?
Charge: $6(+1) + 6(0) + 6(-1) = 0$, so it is neutral. Mass: $6 \times 1 + 6 \times 1 + 6 \times \frac{1}{1836} \approx 12$ (the electrons add almost nothing). It is carbon-12.

Common mistakes

- Saying the atom's mass is made of protons and electrons. Neutrons matter; electrons hardly do.
- Placing electrons inside the nucleus, or saying neutrons have a negative charge.
- Thinking that the shells are solid rings. They are regions where electrons are found (energy levels).

2 Atomic number, mass number and isotopes

Two numbers describe every atom. The **atomic number** Z is the number of protons in the nucleus. It identifies the element: every atom with 6 protons is carbon, and no other element has 6. The **mass number** A is the total number of protons and neutrons (nucleons). So

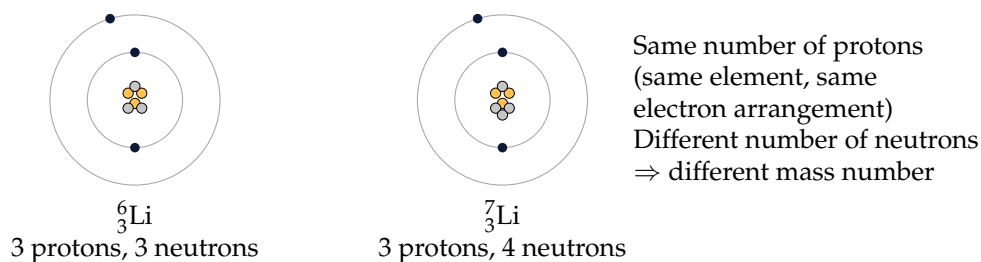
$$\text{number of neutrons} = A - Z, \quad \text{number of electrons (neutral atom)} = Z.$$

They are written as a nuclide symbol with the mass number at the top left and the atomic number at the bottom left: for example ${}^{23}_{11}\text{Na}$ has $Z = 11$ (11 protons, 11 electrons) and $A = 23$, so $23 - 11 = 12$ neutrons.

Worked example: reading a nuclide symbol

How many protons, neutrons and electrons are in an atom of ${}^{35}_{17}\text{Cl}$ and in a ${}^{37}_{17}\text{Cl}$ atom?
Both have $Z = 17$: 17 protons and 17 electrons. Neutrons: $35 - 17 = 18$ in the first atom and $37 - 17 = 20$ in the second.

Isotopes. Atoms of the *same element* always have the same number of protons, but they can have *different numbers of neutrons*. Such atoms are called **isotopes**: same atomic number, different mass number. Because the electrons are the same, isotopes of an element have the same chemical properties; they differ only in mass (so they differ in physical properties such as density).



Ions. An **ion** is an atom (or group of atoms) that has gained or lost electrons, so the numbers of protons and electrons are no longer equal. The nucleus does not change. Losing electrons gives a positive ion (a *cation*) and gaining electrons gives a negative ion (an *anion*). The charge is protons – electrons. For example, ${}^{27}_{13}\text{Al}^{3+}$ has 13 protons, $27 - 13 = 14$ neutrons, and $13 - 3 = 10$ electrons.

Worked example: ions

State the numbers of protons, neutrons and electrons in ${}^{16}\text{O}^{2-}$.
Protons = 8. Neutrons = $16 - 8 = 8$. The charge is $2-$, so there are 2 more electrons than protons: $8 + 2 = 10$ electrons.

Common mistakes

- Using the mass number as the number of neutrons. Neutrons = $A - Z$.
- Thinking isotopes have different numbers of protons. If the proton number changes, the element changes.

- Changing the number of protons to make an ion. Ions form by gaining or losing *electrons*.

3 The mass spectrometer and relative atomic mass

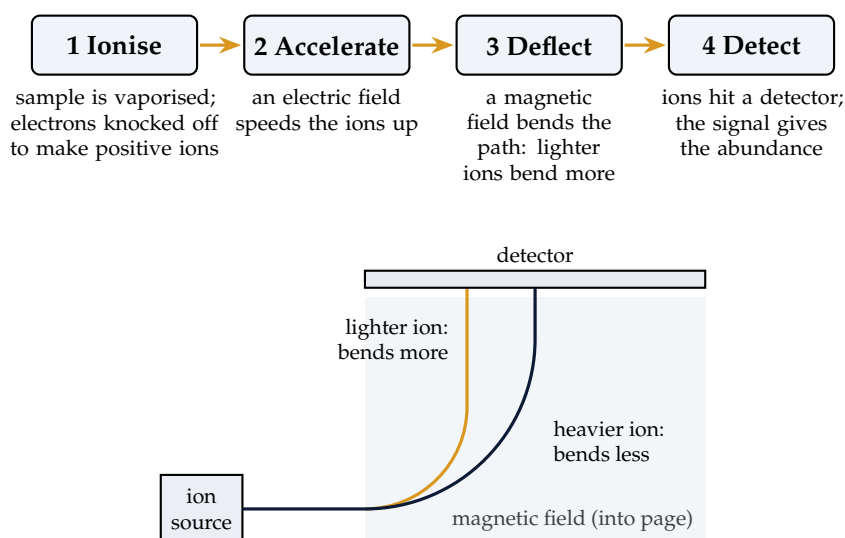
Why A_r is not a whole number. Most elements are a mixture of isotopes in fixed proportions. Chlorine, for example, is about three-quarters ^{35}Cl and one quarter ^{37}Cl in nature. The **relative atomic mass** A_r of an element is the *average* mass of its atoms (allowing for the proportions of its isotopes), measured relative to one-twelfth of the mass of a carbon-12 atom. So A_r is usually not a whole number: the A_r of chlorine is 35.5, which is not the mass of any single atom, but the average.

How an average is found. Suppose 100 atoms of chlorine contain 75 atoms of mass 35 and 25 atoms of mass 37. The total mass is $75 \times 35 + 25 \times 37 = 2625 + 925 = 3550$, so the average per atom is $3550/100 = 35.5$. This is a *weighted* average:

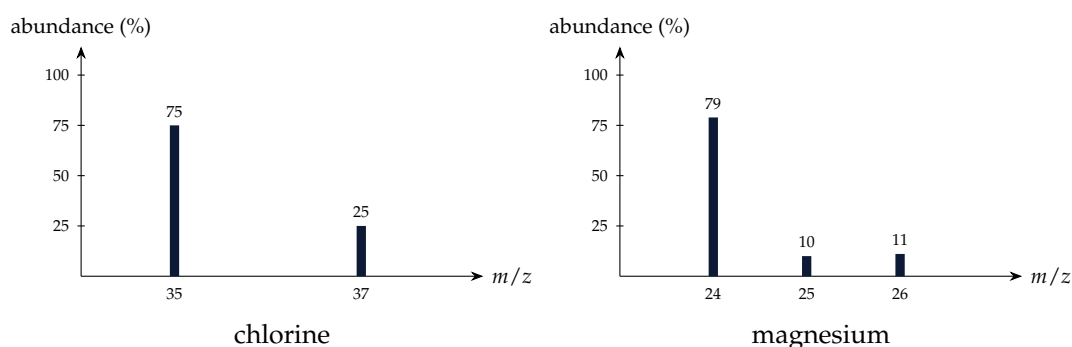
$$A_r = \frac{\sum(\text{isotope mass} \times \text{abundance})}{\sum(\text{abundances})}$$

If the abundances are percentages the denominator is 100.

The mass spectrometer. A mass spectrometer measures the mass of the atoms (or molecules) in a sample, and how much of each there is. The sample is vaporised and then **ionised**: electrons are knocked off to give positive ions. The ions are **accelerated** by an electric field, then passed through a **magnetic field**, which bends their path. A charged particle is bent less when it is heavier, so only ions of one mass reach the detector for a given field, and by changing the field the machine scans through all the masses. The **detector** measures how many ions arrive.



Reading a mass spectrum. The result is a bar chart. The horizontal axis is the *mass-to-charge ratio* m/z ; for ions with a single positive charge this is simply the mass number of the isotope. The height of each bar is the relative abundance (how much of that isotope there is). Each bar therefore tells you "there is an isotope of this mass, and this much of it". The number of bars is the number of isotopes.



Worked example: A_r of magnesium from a spectrum

Using the spectrum of magnesium (abundances 79% of mass 24, 10% of mass 25 and 11% of mass 26), find A_r to one decimal place.

$$A_r = \frac{24 \times 79 + 25 \times 10 + 26 \times 11}{100} = \frac{1896 + 250 + 286}{100} = \frac{2432}{100} = 24.3$$

Sanity check: the answer must lie between 24 and 26, and be nearer 24 because that isotope is far commoner.

Worked example: working backwards

Boron has two isotopes, ^{10}B and ^{11}B , and $A_r = 10.8$. Find the percentage of each.

Let the fraction of ^{11}B be f . Then $10(1 - f) + 11f = 10.8$, so $10 + f = 10.8$ and $f = 0.8$: 80% ^{11}B and 20% ^{10}B . (Shortcut: 10.8 is 0.8 of the way from 10 to 11.)

Exam tip

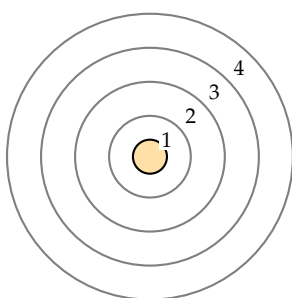
Always estimate first. A_r lies between the lightest and the heaviest isotope, and nearer the one with the greater abundance. If your answer does not, you have made an error.

Common mistakes

- Taking a simple mean of the isotope masses without weighting by abundance.
- Thinking a bar of height 75 means that 75 atoms exist in the sample. It is a relative abundance (a proportion).
- Thinking that the A_r value, for example 35.5, is the mass of some real atom. No chlorine atom has that mass.
- Believing that isotopes have different chemical properties. They have the same electron arrangement, so they react identically.

4 Electron configurations from hydrogen to calcium

Shells and capacity. Electrons occupy shells at increasing distances from the nucleus. The shell closest to the nucleus has the lowest energy, so electrons fill the shells from the inside outwards. Each shell can hold only a limited number of electrons: the first shell holds up to 2, the second up to 8, and for the elements up to calcium the third shell is filled to 8 before the fourth shell starts. (Later the third shell can hold up to 18, but you do not need that here.) The arrangement of electrons in the shells is the **electron configuration**, written as numbers separated by commas, innermost shell first.

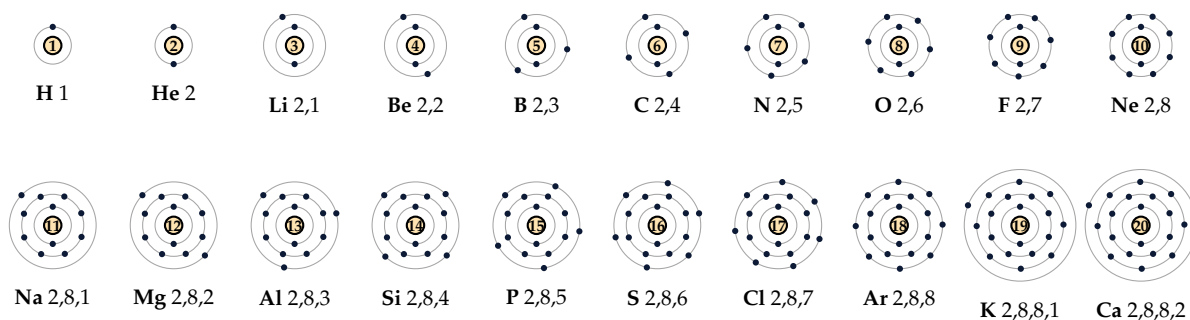


Rules for H to Ca

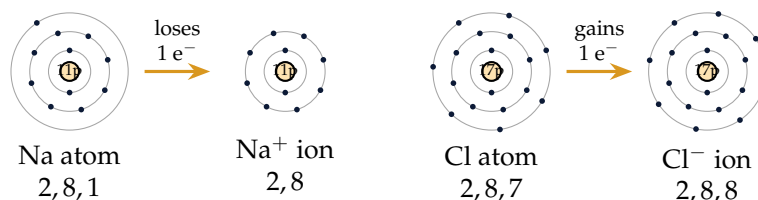
1. Fill the innermost shell first.
2. Shell 1 holds up to 2.
3. Shell 2 holds up to 8.
4. Shell 3 holds 8 (up to Ar);
potassium and calcium start shell 4
(2,8,8,1 and 2,8,8,2).

Total electrons = atomic number.

How to write a configuration. (1) The number of electrons equals the atomic number (for a neutral atom). (2) Put electrons into shell 1 until it holds 2, then shell 2 until 8, then shell 3 until 8, then shell 4. For example, sulfur has $Z = 16$: shell 1 takes 2 (14 left), shell 2 takes 8 (6 left), the remaining 6 go in shell 3: 2, 8, 6. Potassium has $Z = 19$: $2 + 8 + 8 = 18$, so the 19th electron starts shell 4: 2, 8, 8, 1.



Why the outer shell matters. The electrons in the outermost shell (the **outer electrons**, or valence electrons) are the ones involved when atoms react, because they are the farthest from the nucleus. Atoms are most stable when the outer shell is full, as in the noble gases helium (2), neon (2,8) and argon (2,8,8). Atoms of other elements react by gaining, losing or sharing electrons to reach this stable arrangement. An atom that *loses* electrons becomes a positive ion with a full outer shell (for example $\text{Na } 2,8,1 \rightarrow \text{Na}^+ 2,8$); an atom that *gains* electrons becomes a negative ion (for example $\text{Cl } 2,8,7 \rightarrow \text{Cl}^- 2,8,8$).



Worked example: write the configuration

Write the electron configuration of (a) nitrogen ($Z = 7$), (b) magnesium ($Z = 12$), (c) calcium ($Z = 20$).
 (a) $2 + 5$: **2, 5**. (b) $2 + 8 = 10$, remaining 2: **2, 8, 2**. (c) $2 + 8 + 8 = 18$, remaining 2 go into the fourth shell: **2, 8, 8, 2**.
 (Check that each sums to Z : 7, 12, 20.)

Worked example: from a configuration to an element

An atom has the configuration 2, 8, 4. Which element is it?
 The number of electrons is $2 + 8 + 4 = 14$, and a neutral atom has as many protons as electrons, so $Z = 14$: silicon.
 (Note 2, 8, 4 is a neutral atom; the ion Ca^{2+} , by contrast, has the configuration 2, 8, 8 of argon but is calcium.)

Common mistakes

- Putting more than 8 electrons in the second shell, or starting the third shell before the second is full.
- For potassium, writing 2, 8, 9. The third shell holds only 8 here, so the 19th electron goes into the fourth shell.
- Forgetting that the configuration of an *ion* has a different number of electrons from the atom.

5 The Periodic Table: how it is organised

Ordering. The elements are arranged in order of increasing **atomic number**, not mass. (Mendeleev originally ordered them by mass but left gaps and swapped a few pairs so that elements with similar properties lined up; ordering by atomic number removes the oddities, because it is the number of protons, and so of electrons, that determines chemistry.)

Rows and columns. A horizontal row is a **period**; a vertical column is a **group**. Elements in the same group have similar chemical properties. The groups are numbered 1 to 18 across the table. Older books number the tall main groups 1, 2, then 3, 4, 5, 6, 7 and 0 for the groups that are now numbered 13, 14, 15, 16, 17 and 18. So the halogens are **Group 17** (also called Group 7) and the noble gases are **Group 18** (or Group 0). The block in the middle (Groups 3 to 12) holds the **transition metals**.

Gp	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	¹ H																	² He
2	³ Li	⁴ Be											⁵ B	⁶ C	⁷ N	⁸ O	⁹ F	¹⁰ Ne
3	¹¹ Na	¹² Mg											¹³ Al	¹⁴ Si	¹⁵ P	¹⁶ S	¹⁷ Cl	¹⁸ Ar
4	¹⁹ K	²⁰ Ca	²¹ Sc	²² Ti	²³ V	²⁴ Cr	²⁵ Mn	²⁶ Fe	²⁷ Co	²⁸ Ni	²⁹ Cu	³⁰ Zn	³¹ Ga	³² Ge	³³ As	³⁴ Se	³⁵ Br	³⁶ Kr

Period numbers on the left, group numbers along the top. Metals are on the left and in the middle; non-metals are on the right.

 alkali metals	 alkaline earth	 transition metals	 other metals
 semi-metals	 non-metals	 halogens	 noble gases

Metals and non-metals. Most elements are **metals**: they lie on the left and in the middle of the table. Metals are shiny, good conductors of heat and electricity, malleable and ductile, and they form positive ions when they react. **Non-metals** lie on the right (and hydrogen): they are poor conductors (graphite is an exception), their solids are brittle, they have lower melting points, and they form negative ions or share electrons. A few elements near the dividing line (such as silicon and germanium) have properties in between.

Named groups. Group 1: the **alkali metals** (lithium, sodium, potassium, ...). Group 2: the **alkaline earth metals** (beryllium, magnesium, calcium, ...). Groups 3–12: **transition metals**. Group 17: the **halogens** (fluorine, chlorine, bromine, iodine, astatine). Group 18: the **noble gases** (helium, neon, argon, krypton, ...), which are very unreactive.

Worked example: naming positions

Use the table to give the group and period of (a) fluorine, (b) calcium, (c) argon.

(a) Fluorine ($Z = 9$): period 2, Group 17. (b) Calcium ($Z = 20$): period 4, Group 2. (c) Argon ($Z = 18$): period 3, Group 18.

Common mistakes

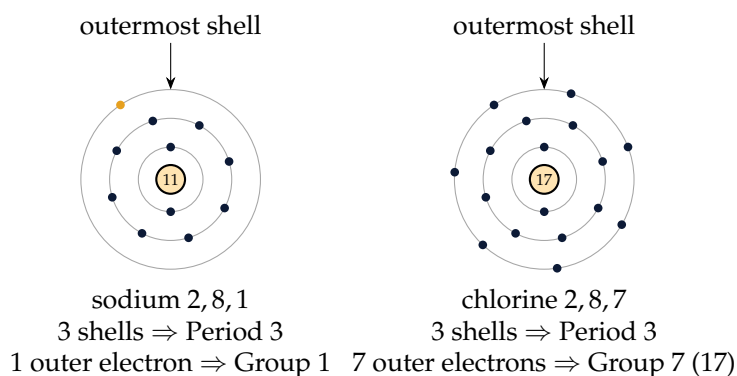
- Saying that the elements are ordered by relative atomic mass. They are ordered by atomic number (the occasional pair such as Ar and K is out of order by mass).
- Confusing periods (rows) with groups (columns).
- Believing hydrogen is an alkali metal because it is in Group 1. It is a non-metal.

6 Position and electron configuration

Two simple rules. The shape of the Periodic Table is a picture of how the shells fill, which gives a pair of rules that you should use every day:

- The **period number** equals the **number of occupied shells**.
- For the main groups, the **number of outer electrons** equals the group number in the old numbering (Groups 1 and 2 are the same in both, and Group 13 has 3 outer electrons, Group 14 has 4, ..., Group 17 has 7, Group 18 has 8, except helium which has 2).

In the figure, sodium (2, 8, 1) has three shells, so it is in period 3, and one outer electron, so Group 1; chlorine (2, 8, 7) is also in period 3 but has seven outer electrons, so it is in Group 17 (the "Group 7" of older books).



Using it both ways. From a configuration to a position: count the shells (period) and the outer electrons (group). From a position to a configuration: the period gives the number of shells; the group gives the outer electrons; fill the inner shells completely (2, then 8, ...). For example, an element in period 3, Group 15: three shells, five outer electrons, so 2, 8, 5, with $2 + 8 + 5 = 15$ electrons: phosphorus.

Why similar properties down a group? All elements in a group have the *same number of outer electrons*, and the outer electrons control how an atom reacts. Lithium (2, 1), sodium (2, 8, 1) and potassium (2, 8, 8, 1) each have a single outer electron to lose, so each forms a $1+$ ion and they react in similar ways. The *trend* in reactivity then depends on how far the outer electrons are from the nucleus (next section).

Worked example: period 3, three outer electrons

An element is in period 3 and has three electrons in its outer shell. Give its configuration, atomic number and group, and the ion it will form.

Three shells and three outer electrons: 2, 8, 3. The atomic number is $2 + 8 + 3 = 13$ (aluminium). It is in Group 13, and it can reach the neon arrangement 2, 8 by losing its three outer electrons, forming Al^{3+} .

Worked example: deducing a position

The configuration of an atom of X is 2, 8, 8, 2 and an ion of Y is Y^- with configuration 2, 8, 8. Place X and Y in the table.

X: four shells (period 4) and two outer electrons: Group 2 (calcium). Y^- has 18 electrons, so the neutral atom has 17 electrons: 2, 8, 7, period 3, Group 17 (chlorine).

Common mistakes

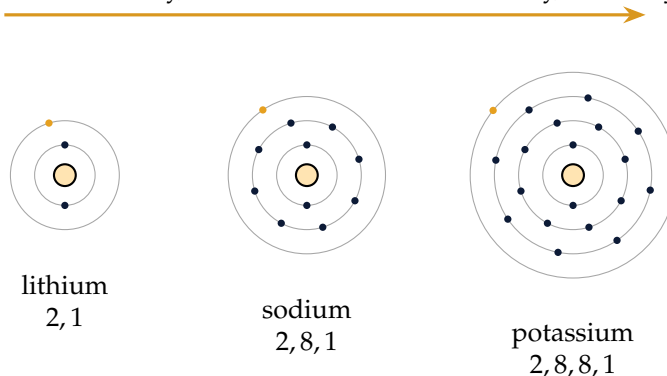
- Using the number of electrons to find the period. The period is the number of *shells*.
- Forgetting that Group 17 has 7 outer electrons, and Group 18 has 8 (not 18).
- Placing an element using the configuration of its *ion*.

7 Group trends: Group 1, Group 17 and Group 18

Group 1: the alkali metals. Lithium, sodium and potassium are soft metals that can be cut with a knife, with low melting points and low densities for metals. They have *one* outer electron, which they readily lose to form a $1+$ ion. They react vigorously with water to give hydrogen and an alkaline solution (the hydroxide), which is why they are called alkali metals. The trends going *down* the group are: melting point **decreases**, and reactivity **increases**.

Why does the reactivity increase down Group 1? Reactions of these metals involve losing the outer electron. Going down the group the outer electron is in a shell further from the nucleus, so it is attracted less strongly by the positive nucleus; it is also screened (shielded) from the nucleus by the extra full inner shells. So it is lost more easily, and the reaction is more vigorous.

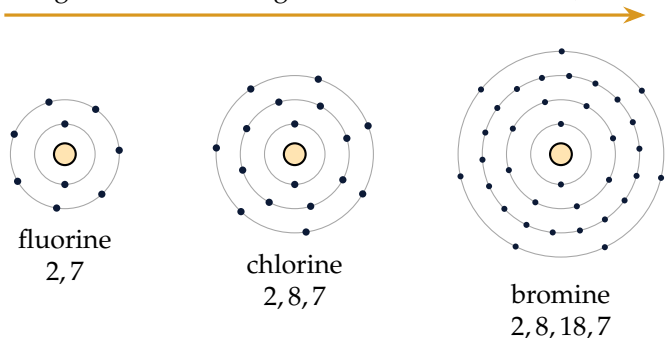
outer electron further away, more shielded: lost more easily, reactivity increases



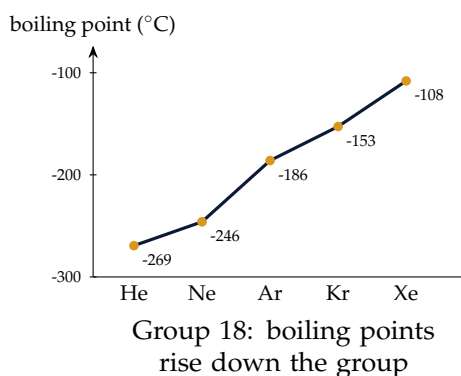
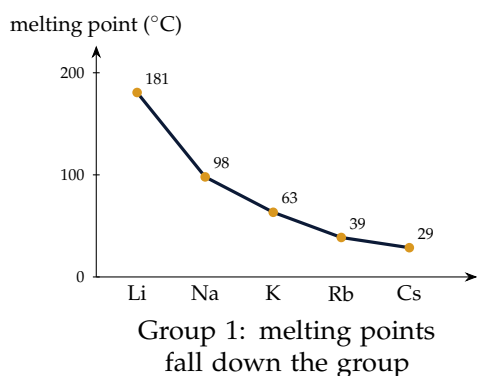
Group 17: the halogens. Fluorine, chlorine, bromine and iodine have *seven* outer electrons and form diatomic molecules (F_2 , Cl_2 , ...). They react by *gaining* one electron to form a 1– ion (the halide) with a full outer shell. Going down the group: melting and boiling points **increase** (fluorine and chlorine are gases at room temperature, bromine is a liquid, iodine is a solid), the colours get darker, and reactivity **decreases**.

Why does the reactivity decrease down Group 17? Now the atom must *capture* an electron into its outer shell. Going down the group the outer shell is further from the nucleus and more shielded, so an incoming electron is attracted less strongly and the atom is a less effective electron-grabber. Fluorine is the most reactive non-metal.

gap to the nucleus grows: an incoming electron is attracted less, reactivity decreases



Group 18: the noble gases. Helium, neon, argon, krypton and xenon all have a full outer shell, so they have no tendency to gain, lose or share electrons: they are chemically unreactive and exist as single atoms (monatomic). Their boiling points are very low and *increase* down the group.



Using trends to predict. A trend allows you to predict the properties of an element you have never met. Rubidium sits below potassium in Group 1, so we predict it is more reactive than potassium, softer and with a lower melting point. Astatine sits below iodine in Group 17: from the trend in melting points (fluorine –220, chlorine –101, bromine –7, iodine 114 °C) we predict it is a solid with a melting point above that of iodine, and that it is less reactive than iodine.

Worked example: explain the order of reactivity

Explain why potassium is more reactive than sodium.

Both have one outer electron and react by losing it. In potassium (2,8,8,1) the outer electron is in the fourth shell, further from the nucleus and more shielded by the extra inner shell than in sodium (2,8,1). It is attracted less strongly, so it is lost more easily.

Worked example: predicting a missing value

Using the Group 1 melting points Li 181, Na 98, K 63, Rb 39 ($^{\circ}\text{C}$), is it reasonable to predict that caesium's melting point is about 29°C or about 90°C ?

The melting point falls down the group, and the decreases get smaller (83,35,24, ...), so the next value should be a bit lower than 39: about 29°C is reasonable. A value of 90 would reverse the trend.

Common mistakes

- Saying that reactivity of the halogens increases down the group (it *decreases*) while reactivity of the alkali metals increases.
- Explaining the trend by "more electrons" or "bigger atoms" alone. Say that the outer electron is further from the nucleus and more shielded, so it is attracted less.
- Saying that the noble gases are unreactive because they "don't want to react". They have a full outer shell, so there is no gain in losing, gaining or sharing electrons.

8 Metals, non-metals and trends across a period

Across a period. Moving left to right across period 3 from sodium to argon, the atomic number rises by one each step, and one more electron is added to the *same* (third) shell. The number of outer electrons rises from 1 to 8, so the properties change in a steady pattern: the elements change from metals (Na, Mg, Al), through semi-metals (Si), to non-metals (P, S, Cl) and finally to an unreactive noble gas (Ar). Metals lose their outer electrons to form positive ions with a charge equal to the number of outer electrons (Na^+ , Mg^{2+} , Al^{3+}); non-metals gain electrons until the shell is full (Cl needs 1 more: Cl^- ; S needs 2: S^{2-}), or share electrons.

Pattern of ion charges.

group	1	2	13	14	15	16	17
outer electrons	1	2	3	4	5	6	7
typical ion	1+	2+	3+	(shares)	3-	2-	1-

A useful way to see the pattern: a metal ion has the same configuration as the noble gas *before* it, and a non-metal ion has the same configuration as the noble gas *after* it. This is why Na^+ , Mg^{2+} , O^{2-} and F^- all have the neon configuration 2,8 and are called *isoelectronic*.

Worked example: isoelectronic ions

Which atom has the same electron configuration as S^{2-} , and how many protons does S have?

S has 16 electrons (2,8,6); S^{2-} has two more: 18 electrons, 2,8,8, which is the configuration of argon. S has 16 protons.

Exam tip

For any ESAT question on a species, make a small table with rows protons, neutrons and electrons, and fill it using Z , $A - Z$ and $Z - \text{charge}$. Then every other question (group, period, ion, isotope) follows from it.

Common mistakes

- Confusing a species' number of electrons with its group number when it is an ion.
- Saying that sulfur forms S^{6-} because it has six outer electrons. It *gains* two to make eight.
- Thinking the number of shells changes across a period. It stays the same; only the number of outer electrons (and protons) rises.

Check yourself

Try these without looking back, then compare with the answers on the next page.

1. How many protons, neutrons and electrons are in $^{31}_{15}\text{P}$?
2. Write the electron configurations of (a) oxygen ($Z = 8$) and (b) potassium ($Z = 19$).
3. An element is in period 2 and Group 15. Which element is it?
4. Lithium has isotopes of mass 6 and 7 in the ratio 1 : 9 (by number of atoms). Find A_r .
5. What does the number of bars in a mass spectrum of an element tell you?
6. Why do sodium and potassium react in similar ways?
7. Why is caesium more reactive than lithium?
8. Magnesium forms Mg^{2+} . How many protons and electrons does the ion have, and which atom has the same electron configuration?

Answers to the self-check

- 15 protons, $31 - 15 = 16$ neutrons, 15 electrons.
- (a) 2, 6. (b) 2, 8, 8, 1.
- Two shells and five outer electrons: 2, 5, so $Z = 7$: nitrogen.
- $(6 \times 1 + 7 \times 9) / 10 = 69 / 10 = 6.9$.
- The number of different isotopes (in the sample).
- Both have one outer electron, which they lose to form a $1+$ ion.
- Its outer electron is in a shell much further from the nucleus and is more shielded, so it is attracted less strongly and lost more easily.
- 12 protons and 10 electrons (2, 8, the configuration of neon).

Next step

Heurisco's *ESAT Chemistry: Atomic Structure and the Periodic Table Practice Set* has twenty original exam-style questions on this lesson, with shell diagrams, mass spectra and Periodic Table extracts, in order of difficulty, with a full worked solution for each.