Unit 6

Periodic Table

This unit online, with its interactives:
science-playground.com/s/10a6

Drop a piece of sodium the size of a grain of rice into a bowl of water, and it fizzes, melts into a shining ball and races about on the surface, sometimes bursting into flame. Fill a glass tube with neon instead, and nothing you can do will make it react: heat it, spark it, mix it with fluorine, the most reactive element there is, and it stays neon. Yet a sodium atom has only one proton more than a neon atom.

That one proton brings one more electron, and the electron has to start a new shell, far from the nucleus and loosely held. Neon's electrons fill every place they have. In this unit you will see how the periodic table is built from the way electrons fill their shells and subshells, how that sorts the elements into s, p, d and f blocks, how atomic size, valency, ionization potential, electron affinity and electronegativity change across the table, with the real measured numbers, and how all of it decides the way groups 1, 2, 15, 16 and 17 react.

6.1 Periodic table

Modern periodic law

Mendeleev arranged the elements in 1869 in order of their atomic masses, and had to break his own order in a few places to keep similar elements together. In 1913 Henry Moseley showed with X-rays that each element has one more unit of positive charge on its nucleus than the one before it (the story is in Grade 9 Optional Science). The order the elements really follow is the number of protons, the **atomic number**. This is the **modern periodic law**:

Modern periodic law: the physical and chemical properties of the elements are periodic functions of their atomic numbers.

"Periodic" means repeating at regular intervals, like the days of the week. Go along the elements in order of atomic number, and similar properties come back again and again: a soft, very reactive metal at 3, 11, 19, 37 and 55; an unreactive gas at 2, 10, 18, 36, 54 and 86.

Why should atomic number, and not mass, decide the properties? Because an atom's chemistry is done by its electrons, and a neutral atom has as many electrons as protons. The mass of an atom also depends on its neutrons, which take no part in reactions. That is why every isotope of an element shares one place in the table, and why three pairs that were out of order by mass fall into place by atomic number:

Pair	Atomic masses	Atomic numbers	Order in the table
argon and potassium	Ar 39.95, K 39.10	Ar 18, K 19	Ar before K
cobalt and nickel	Co 58.93, Ni 58.69	Co 27, Ni 28	Co before Ni
tellurium and iodine	Te 127.60, I 126.90	Te 52, I 53	Te before I

In each pair the first element is heavier, because it has more neutrons, but has one proton fewer, so it comes first.

Modern periodic table

The table built on the modern periodic law is the **modern periodic table**, or the **long form** of the periodic table. Its main features:

1. The elements are in order of increasing atomic number.
2. There are **7 horizontal rows, the periods**, and **18 vertical columns, the groups**, numbered 1 to 18 as IUPAC recommends (the older numbering, IA to VIIA, IB to VIIIB and 0, is in Grade 10 Science).
3. Elements in one group have the same number of electrons in their outermost shell, so they have similar properties.
4. Elements in one period have the same number of shells: the **period number is the number of the outermost shell**. Sodium, 2, 8, 1, has three shells and is in period 3.
5. Fourteen elements after lanthanum and fourteen after actinium are printed as two rows below the table, with lanthanum and actinium themselves, so that the table isn't too wide to print.

The lengths of the periods, 2, 8, 8, 18, 18, 32 and 32, look odd until you see which subshells each period fills. A period starts when a new shell begins, with its s subshell, and ends when that shell's p subshell is full:

Period	Subshells filled, in order	Number of elements
1	1s	2
2	2s, 2p	2 + 6 = 8
3	3s, 3p	2 + 6 = 8
4	4s, 3d, 4p	2 + 10 + 6 = 18
5	5s, 4d, 5p	2 + 10 + 6 = 18
6	6s, 4f, 5d, 6p	2 + 14 + 10 + 6 = 32
7	7s, 5f, 6d, 7p	2 + 14 + 10 + 6 = 32

So the shape of the table is the shape of the electron subshells: you will see exactly how in 6.2. All 118 elements are now known and named, and period 7 has been complete since oganesson, element 118, was named in 2016.

Merits of the modern periodic table (over Mendeleev's):

- It is based on the atomic number, a basic property of the atom, not on the atomic mass.
- The places of argon and potassium, cobalt and nickel, and tellurium and iodine are no longer anomalies.
- Isotopes need no separate places: they have the same atomic number.
- It explains why properties repeat: they repeat because the outer electronic configuration repeats.
- It keeps metals, non-metals and metalloids in separate regions, and puts the transition metals in a block of their own.

Its limitations:

- Hydrogen has no proper place: its $1s^1$ suits group 1, but it is a non-metal that also gains one electron like the halogens.
- Helium ($1s^2$) is an s-block element by its electrons but is placed in group 18, with the gases it behaves like.
- The lanthanides and actinides are printed outside the main table.

To look up any element, with its story, open our Periodic Table.

Check 1

Why does argon (atomic mass 39.95) come before potassium (atomic mass 39.10) in the modern periodic table?

a argon was discovered first
b argon is a gas and potassium a solid
c argon has 18 protons and potassium 19
d argon has fewer neutrons than potassium

6.2 Electronic configuration

Orbitals, subshells and shells

You met subshells in Grade 10 Science: each shell n has n subshells, s, p, d and f, and each subshell has a fixed number of **orbitals**, regions where an electron is likely to be found. Each orbital holds at most two electrons.

Subshell	Orbitals	Most electrons
s	1	2
p	3	6
d	5	10
f	7	14

An orbital is often drawn as a box, and each electron in it as an arrow, ↑ or ↓, for the two ways an electron can **spin**. (Electrons behave like tiny magnets that point one of two ways; Going Further shows the experiment that found it.) The **electronic configuration** of an atom says how its electrons are arranged in these subshells. Three rules decide it.

Three rules for filling orbitals

1. The Aufbau principle. Electrons fill the subshells in order of increasing energy: a subshell starts to fill only when the subshells of lower energy are full. Up to atomic number 36 the order is

$$1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p$$

The surprise in it is that 4s comes before 3d. A quick way to get the order is the $(n + l)$ rule: give s, p, d, f the numbers $l = 0$, 1, 2, 3, add the shell number n, and the subshell with the smaller sum fills first; for equal sums, the one with the smaller n. For 4s, $4 + 0 = 4$; for 3d, $3 + 2 = 5$, so 4s fills first. (The diagonal diagram for the same order is in Grade 10 Science.)

2. Pauli's exclusion principle. An orbital can hold at most two electrons, and if it holds two, they must spin in opposite directions, ↑↓. Wolfgang Pauli stated it in 1925 in a more general form: no two electrons in an atom can be in exactly the same state.

3. Hund's rule. When electrons go into a subshell with more than one orbital (p, d or f), they go into separate orbitals first, one in each, all spinning the same way. Only when every orbital of the subshell has one electron do they start to pair up. Electrons repel each other, and in separate orbitals they keep farther apart. The rule was found by the German physicist Friedrich Hund in the 1920s.

Look at the 2p subshell of carbon, nitrogen and oxygen:

Element	2p electrons	Boxes for 2p	Unpaired electrons
carbon (6)	2	↑ · ↑ · empty	2
nitrogen (7)	3	↑ · ↑ · ↑	3
oxygen (8)	4	↑↓ · ↑ · ↑	2

Nitrogen's three 2p electrons sit one to an orbital, not two in one orbital and one in another. Oxygen's fourth has to pair up.

Configurations of the elements up to atomic number 30

A full configuration repeats the same inner electrons over and over, so chemists shorten it with the noble gas whose configuration is the same as the inner electrons. Sodium, $1s^2\ 2s^2\ 2p^6\ 3s^1$, is written **[Ne] $3s^1$**: the neon core and one more electron. The core electrons are held tightly and take no part in reactions; what follows the core is what makes the element behave as it does.

Z	Element	Configuration	Unpaired
1	H	$1s^1$	1
2	He	$1s^2$	0
3	Li	[He] $2s^1$	1
4	Be	[He] $2s^2$	0
5	B	[He] $2s^2\ 2p^1$	1
6	C	[He] $2s^2\ 2p^2$	2
7	N	[He] $2s^2\ 2p^3$	3
8	O	[He] $2s^2\ 2p^4$	2
9	F	[He] $2s^2\ 2p^5$	1
10	Ne	[He] $2s^2\ 2p^6$	0
11	Na	[Ne] $3s^1$	1
12	Mg	[Ne] $3s^2$	0
13	Al	[Ne] $3s^2\ 3p^1$	1
14	Si	[Ne] $3s^2\ 3p^2$	2
15	P	[Ne] $3s^2\ 3p^3$	3
16	S	[Ne] $3s^2\ 3p^4$	2
17	Cl	[Ne] $3s^2\ 3p^5$	1
18	Ar	[Ne] $3s^2\ 3p^6$	0
19	K	[Ar] $4s^1$	1
20	Ca	[Ar] $4s^2$	0
21	Sc	[Ar] $4s^2\ 3d^1$	1
22	Ti	[Ar] $4s^2\ 3d^2$	2
23	V	[Ar] $4s^2\ 3d^3$	3
24	Cr	[Ar] $4s^1\ 3d^5$	6
25	Mn	[Ar] $4s^2\ 3d^5$	5
26	Fe	[Ar] $4s^2\ 3d^6$	4
27	Co	[Ar] $4s^2\ 3d^7$	3

Z	Element	Configuration	Unpaired
28	Ni	[Ar] $4s^2\ 3d^8$	2
29	Cu	[Ar] $4s^1\ 3d^{10}$	1
30	Zn	[Ar] $4s^2\ 3d^{10}$	0

Two exceptions: chromium and copper

By the Aufbau order, chromium (24) should be [Ar] $4s^2\ 3d^4$ and copper (29) [Ar] $4s^2\ 3d^9$. Experiments on their spectra and magnetism show something else:

$$\text{Cr} : [\text{Ar}]\ 4s^1\ 3d^5 \qquad\qquad \text{Cu} : [\text{Ar}]\ 4s^1\ 3d^{10}$$

In each, one electron has moved from 4s into 3d. By this point in period 4 the energies of 4s and 3d are very close, and a **half-filled** (d^5) or **completely filled** (d^{10}) d subshell is especially stable: its electrons are spread evenly round the nucleus, and in d^5 all five spin the same way, which lowers their energy. The move costs a little energy and gains more. So chromium has six unpaired electrons, the most of any element up to 30.

> CAREFUL
> **Two ways of writing the order**
>
> Some books write the subshells in order of energy (iron: [Ar] $4s^2\ 3d^6$), as this book does, and others in order of shell number ([Ar] $3d^6\ 4s^2$). Both are correct and both get full marks; just don't change the number of electrons in each subshell.

There is one more surprise. When a transition metal forms an ion, it loses its **4s** electrons first, not the 3d electrons that went in last. Iron, [Ar] $4s^2\ 3d^6$, becomes Fe^{2+}, [Ar] $3d^6$, and then Fe^{3+}, [Ar] $3d^5$. Once 3d holds electrons, it drops below 4s in energy. Notice that Fe^{3+} has a half-filled 3d: it is the more stable of iron's two ions in air, one reason why rust is an iron(III) compound.

WORKED EXAMPLE

Iron by its electrons

Write the electronic configuration of iron (Z = 26), full and short, and find its number of unpaired electrons.

Answer

Fill 26 electrons in the Aufbau order: $1s^2$ (2), $2s^2$ (4), $2p^6$ (10), $3s^2$ (12), $3p^6$ (18), $4s^2$ (20), and the remaining 6 go into 3d.

Full: **$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$**. Short: **[Ar] $4s^2\ 3d^6$**.

The 3d subshell has 5 orbitals. By Hund's rule the first five electrons go in one each (↑ ↑ ↑ ↑ ↑) and the sixth pairs with the first (↑↓ ↑ ↑ ↑ ↑). All the inner subshells and 4s are full. So iron has **4 unpaired electrons**, which is why iron atoms are strong little magnets.

YOUR TURN

Three more

Write the short configurations of (i) manganese (25), (ii) copper (29) and (iii) the zinc ion, Zn^{2+}, and give the number of unpaired electrons in each.

Answer

i. Mn: [Ar] $4s^2\ 3d^5$, 5 unpaired (one in each 3d orbital).

ii. Cu: [Ar] $4s^1\ 3d^{10}$, 1 unpaired (the 4s electron), one of the two exceptions.

iii. Zn is [Ar] $4s^2\ 3d^{10}$; it loses its two 4s electrons, so Zn^{2+} is [Ar] $3d^{10}$, with 0 unpaired.

Finding an element's place from its configuration

The configuration tells you where an element sits, without looking at a table:

- **Period** = the number of the outermost shell (the highest n).
- **Block** = the subshell the last electron goes into, by the Aufbau order: s, p, d or f.
- **Group:** for the s block, the number of s electrons in the outermost shell (1 or 2); for the p block, 12 + the number of p electrons (so p^1 is group 13 and p^6 group 18); for the d block, the number of $(n-1)$d electrons plus the ns electrons.

WORKED EXAMPLE

Where do they belong?

Find the period, block and group of the elements with atomic numbers (i) 17, (ii) 20 and (iii) 24.

Answer

i. $Z = 17$: [Ne] $3s^2\ 3p^5$. Outermost shell 3, so **period 3**; last electron in 3p, so **p block**; $12 + 5 =$ **group 17**. It is chlorine, a halogen.

ii. $Z = 20$: [Ar] $4s^2$. **Period 4**, **s block**, two s electrons: **group 2**. It is calcium.

iii. $Z = 24$: [Ar] $4s^1\ 3d^5$. **Period 4**; the last electron by the Aufbau order goes into 3d, so **d block**; $5 + 1 =$ **group 6**. It is chromium.

6.3 s-block, p-block, d-block and f-block elements

Description

Each block of the table is the set of elements whose last electron goes into one kind of subshell. A block's width is the number of electrons its subshell holds: the s block is 2 groups wide, the p block 6, the d block 10 and the f block 14.

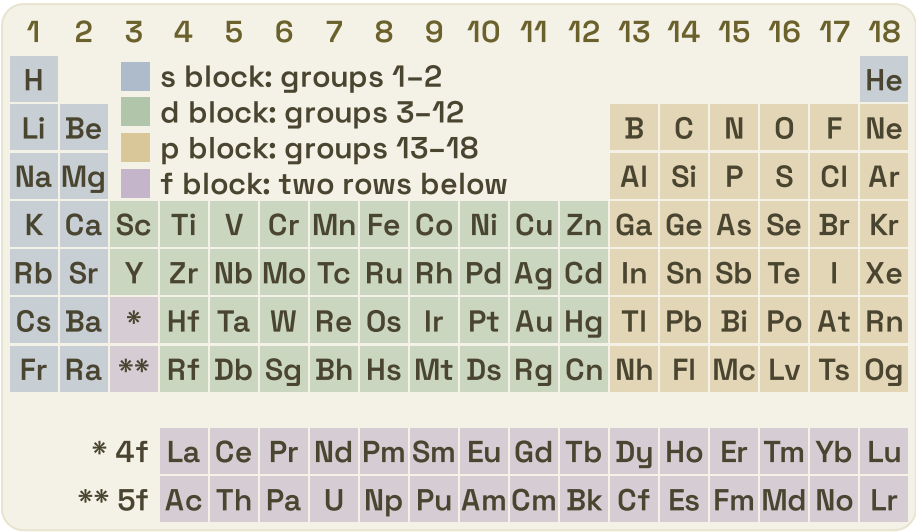


Figure 6.1 *The periodic table colored by block, with the group numbers above it. Helium ($1s^2$) belongs to the s block by its electrons, but is placed in group 18 because it behaves like the noble gases. The f block's two rows of 15 belong in group 3 of periods 6 and 7 (marked * and **), and are printed below the table.*

Block	Groups	Outer configuration	Elements
s	1, 2 (and He)	ns^1 or ns^2	alkali and alkaline earth metals, hydrogen, helium
p	13 to 18	$ns^2\ np^{1-6}$	metals, metalloids, non-metals and noble gases
d	3 to 12	$(n-1)d^{1-10}\ ns^{1-2}$	the transition metals
f	below the table	$(n-2)f^{1-14}\ (n-1)d^{0-1}\ ns^2$	lanthanides and actinides

s-block elements

The s block is groups 1 and 2, with hydrogen and helium. Group 1 (lithium to francium) are the **alkali metals**, ns^1; group 2 (beryllium to radium) the **alkaline earth metals**, ns^2.

Characteristics:

- All except hydrogen and helium are **metals**: soft (sodium and potassium can be cut with a knife), light, with low melting points that fall down group 1 (lithium 181 °C, sodium 98 °C, potassium 63 °C, cesium 28 °C, which melts in a warm hand).
- They have **low ionization potentials**, so they lose their one or two outer electrons easily, forming M^+ or M^{2+} ions. They are the most **electropositive** elements and strong reducing agents.
- They have a **fixed valency**: 1 for group 1, 2 for group 2.
- They are **very reactive**, so they are never found free in nature; group 1 metals are stored under oil.
- They form mostly **ionic** compounds, and their oxides and hydroxides are **basic** (NaOH and KOH are strong alkalis).
- Many give **colors to a flame**: lithium crimson, sodium yellow, potassium lilac, calcium orange-red, strontium red, barium pale green.
- Their compounds are usually white or colorless.

p-block elements

The p block is groups 13 to 18, from boron's group to the noble gases, with outer configurations from $ns^2\ np^1$ to $ns^2\ np^6$.

Characteristics:

- It is the only block with **all three kinds of element**: metals (aluminium, tin, lead, bismuth), metalloids (boron, silicon, germanium, arsenic, antimony, tellurium) and non-metals (carbon, nitrogen, oxygen, phosphorus, sulphur, the halogens). The noble gases end it.
- Most p-block elements have **high ionization potentials and electron affinities**: they tend to gain or share electrons, not lose them.
- Many show **variable valency**: phosphorus 3 and 5 (PCl_3, PCl_5), sulphur 2, 4 and 6 (H_2S, SO_2, SO_3).
- They form both **covalent** compounds (with each other) and **ionic** ones (with s-block metals).



Figure 6.2 *Flames colored by salts dissolved in methanol and burned on cotton wool. From the left: the chlorides of lithium, strontium, calcium, sodium and barium, then a boron compound, copper chloride, cesium chloride and potassium chloride. Each metal's electrons, excited by the heat, give out light of their own colors as they drop back to lower energy levels. Fireworks get their colors the same way.*
Photo: Hegelrast, CC BY-SA 4.0

- Oxides of the non-metals are **acidic** (CO_2, SO_2, P_4O_{10}); some of the metals' oxides are amphoteric (Al_2O_3).
- **Metallic character increases down each group**: in group 14, carbon is a non-metal, silicon and germanium are metalloids, tin and lead are metals.
- Group 17 (the **halogens**) are the most reactive non-metals; group 18 (the **noble gases**) have full outer shells and hardly react at all.

d-block elements

The d block is the ten groups 3 to 12, in the middle of the table, where a d subshell of the shell below the outermost one is being filled: $(n-1)d^{1-10}\ ns^{1-2}$. The first row, scandium to zinc, fills 3d. Most of these elements are called **transition elements** or transition metals, because they form a bridge between the very reactive s-block metals on the left and the p block on the right.



Figure 6.3 *Three halogens in sealed glass: chlorine, a yellow-green gas; bromine, a red-brown liquid giving off brown vapor; and iodine, gray-black crystals with a violet vapor. The color darkens and the melting point rises down group 17.*
Photo: W. Oelen, CC BY-SA 3.0

Characteristics:

- All are **metals**: mostly hard and strong, with high melting and boiling points (tungsten's, 3,422 °C, is the highest of any metal), good conductors of heat and electricity. Mercury, a liquid, and zinc, which melts at 420 °C, are the odd ones out, because their d subshells are full.
- Their atoms change **little in size** across each row (6.4), so one metal's atoms can take the place of another's, and they form **alloys** easily: stainless steel (iron, chromium, nickel), brass (copper, zinc).
- They show **variable valency**, because both the ns and some $(n-1)$d electrons can be lost: iron 2 and 3 ($FeCl_2$, $FeCl_3$), copper 1 and 2 (Cu_2O, CuO), manganese 2 to 7.
- Most of their compounds are **colored**: copper sulphate crystals are blue, iron(II) sulphate pale green, potassium permanganate deep purple, potassium dichromate orange. The color comes from electrons jumping between d orbitals of slightly different energy, absorbing part of the light. So an ion with an empty d subshell (Sc^{3+}, $3d^0$) or a full one (Zn^{2+}, $3d^{10}$) gives colorless or white compounds: zinc sulphate is white.
- Many are **paramagnetic**, because of their unpaired d electrons, and iron, cobalt and nickel are ferromagnetic (see Unit 14).
- Many are good **catalysts**: iron in making ammonia (the Haber process), vanadium(V) oxide in making sulphuric acid, nickel in hardening vegetable oils into vanaspati ghee, and platinum, palladium and rhodium in the catalytic converters of cars.
- They form **complex ions**, a metal ion with molecules or ions bound round it: the deep blue $[Cu(NH_3)_4]^{2+}$ that appears when ammonia is added to copper sulphate solution.

f-block elements

The f block is the two rows printed below the table, where an f subshell two shells inside the outermost is filled: $(n-2)f^{1-14}$. They are also called the **inner transition elements**.

- The **lanthanides**, lanthanum (57) to lutetium (71), fill 4f. They were once called the **rare earths**, but most are not rare: cerium is more common in the Earth's crust than copper.
- The **actinides**, actinium (89) to lawrencium (103), fill 5f.

Characteristics:

- All are **metals**, silvery and fairly reactive, and most show a valency of **3**.

- Within each series the properties are **very similar**, because the electrons being added go into a deep inner subshell that hardly affects the outside of the atom. Separating one lanthanide from another took chemists more than a century.
- Across the lanthanides the atoms and ions get steadily smaller (the **lanthanide contraction**).
- **All the actinides are radioactive.** Only thorium and uranium are found in nature in large amounts; the elements after uranium (the **transuranium elements**) are made in nuclear reactors and accelerators.
- Uses: neodymium in the strong magnets of phones, earbuds, hard disks, electric cars and wind turbines; europium in the red of older TV screens and of energy-saving lamps; cerium in lighter flints; uranium as nuclear fuel; a tiny piece of americium in many smoke detectors.



Figure 6.4 *Copper sulphate crystals, $CuSO_4 \cdot 5H_2O$. The blue comes from the copper(II) ion, [Ar] $3d^9$, whose one unpaired d electron can jump between d orbitals by absorbing red and orange light. The light left over looks blue.*

Photo: W. Oelen, CC BY-SA 3.0

CAREFUL

Is zinc a transition metal?

Strictly, IUPAC defines a transition element as one whose atom, or an ion it forms, has a **partly filled d subshell**. Zinc ($3d^{10}$ in the atom and in Zn^{2+}), cadmium and mercury keep their d subshells full, so by that rule they are d-block elements but not transition elements, and indeed they don't give colored compounds or show variable valency. **In the exam:** Grade 10 Science and most school books call all ten groups, 3 to 12, the transition metals, so for period 4 write Sc to Zn; you may add that zinc, with its full 3d, is not a typical one.



Figure 6.5 *Neodymium magnets on their steel brackets, taken from old computer hard disks. Made of neodymium, iron and boron ($Nd_2Fe_{14}B$), they are the strongest permanent magnets sold.*
Photo: Raimond Spekking, CC BY-SA 4.0

A STORY FROM SCIENCE

Moving the actinides

Until the 1940s, thorium, protactinium and uranium were placed in the main table, under hafnium, tantalum and tungsten, as transition metals. Working with the new elements neptunium and plutonium (he was one of plutonium's discoverers) during the Second World War, the American chemist **Glenn Seaborg** found that they behaved not like transition metals but like the lanthanides. In 1944 he proposed that the elements from actinium on form a second f series, and moved them out to a row of their own below the lanthanides. Colleagues warned him it would ruin his reputation. Instead it told chemists exactly where to look for the next elements, americium and curium, which he and his team soon made. He shared the Nobel Prize in Chemistry in 1951, and in 1997 element 106 was named seaborgium, the first element named after a living person.



Figure 6.6 *Glenn Seaborg (1912–1999) at Berkeley in 1950, beside an ion-exchange column, the kind of apparatus his team used to separate the actinides from one another, with a periodic table chart behind him.*
Photo: Unknown photographer, Public domain

Check 2

An element has the outer configuration $3d^5\ 4s^2$ ([Ar] $4s^2\ 3d^5$). It belongs to:

a the s block, group 2
b the p block, group 17
c the d block, group 7
d the f block

Check 3

Which property belongs to the transition metals and not to the alkali metals?

a they lose electrons to form positive ions
b they form colored compounds and show variable valency
c they conduct electricity
d they react with chlorine

6.4 Variation of periodic properties of elements

Atomic size, valency, ionization potential, electron affinity and electronegativity are called **periodic properties**: they change regularly along a period and down a group. Almost all of the changes come from two things:

- **Nuclear charge.** The nucleus, with Z protons, pulls the electrons in.
- **Shielding.** The inner electrons sit between the nucleus and the outer electrons and push the outer electrons away, canceling part of the nucleus's pull. Electrons in the same shell shield each other only a little.

A rough count of the **pull felt** by an outer electron is the number of protons minus the number of inner electrons. Across a period, each step adds a proton but no inner electron, so the pull felt rises. Down a group, each step adds a whole new shell, so the pull felt stays about the same, but the outer electrons are farther out.

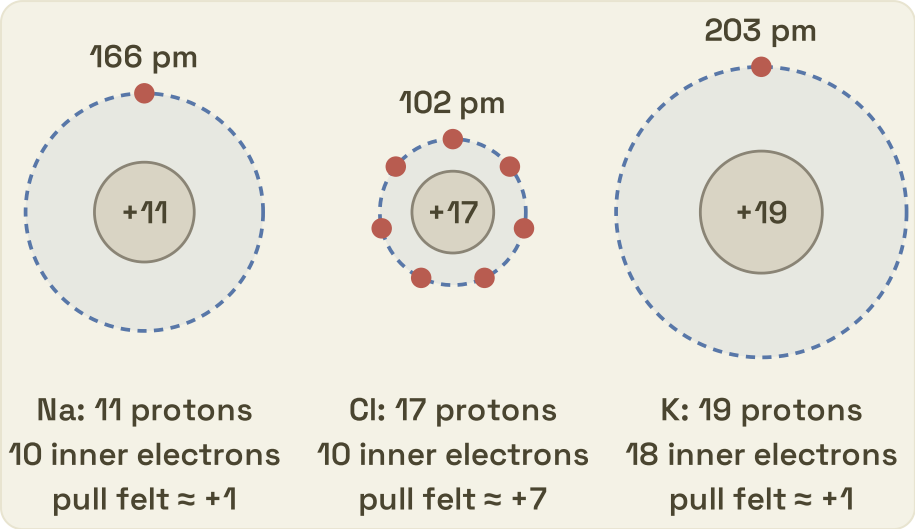


Figure 6.7 *Sodium, chlorine and potassium, with their outer shells drawn to scale. Chlorine's outer electrons feel a pull of about +7 (17 protons minus 10 inner electrons) and are drawn in close; sodium's one outer electron feels only about +1. Potassium's outer electron also feels about +1, but sits a shell farther out. (The rough count leaves out the small shielding by electrons in the same shell; the cores are not to scale.)*

Atomic size

An atom has no sharp edge, so its size is measured from the distances between nuclei. The **covalent radius** is half the distance between the nuclei of two identical atoms joined by a single covalent bond: two chlorine atoms in Cl_2 are 199 pm apart, so chlorine's covalent radius is about 100 pm (102 pm, averaged over thousands of its compounds). (For metals the **metallic radius**, half the distance between neighboring atoms in the metal, is used too; 1 pm = 10^{-12} m.) The values below are covalent radii, all measured the same way.

Across a period, atomic size decreases. The number of shells stays the same, but the pull felt by the outer shell grows, drawing it in:

Period 2	Li	Be	B	C	N	O	F
radius (pm)	128	96	84	76	71	66	57

Period 3	Na	Mg	Al	Si	P	S	Cl
radius (pm)	166	141	121	111	107	105	102

Down a group, atomic size increases, because each step adds a new shell, farther from the nucleus, while the pull felt hardly changes:

Group 1	Li	Na	K	Rb	Cs
radius (pm)	128	166	203	220	244

Group 17	F	Cl	Br	I
radius (pm)	57	102	120	139

Across a transition series, the size changes much less. From scandium (170 pm) to zinc (122 pm), the atoms shrink at first and then most stay between 120 and 140 pm. Each new electron goes into the inner 3d subshell, where it shields the outer 4s electrons well from the extra proton, so the pull on the outside barely grows.

Ions. A positive ion is smaller than its atom, and a negative ion larger. Sodium (radius 166 pm) loses its whole outer shell to become Na^+, with an ionic radius of 102 pm. Chlorine (102 pm) gains an electron into its outer shell: the 18 electrons of Cl^- repel each other more, and the same 17 protons hold them less tightly, so Cl^- swells to 181 pm.

CAREFUL

The noble gases' size

Some books show the noble gases as the largest atoms of their periods. That compares two different measurements: a noble gas hardly forms bonds, so its size is usually given as a **van der Waals radius**, half the distance between two atoms that touch without bonding, which is always much larger than a covalent radius (argon 188 pm, chlorine's covalent radius 102 pm). Sizes calculated the same way for every free atom (a different scale from the covalent radii in the tables above) continue the trend to the end of the period: neon 38 pm against fluorine's 42, argon 71 against chlorine's 79. **In the exam:** follow Grade 10 Science, which says atomic size decreases across a period from group 1 to group 18. If a question says a noble gas is larger, the reason expected is that its radius is a van der Waals radius, not a covalent one.

Valency

The **valency** of an element is its combining capacity: the number of electrons an atom loses, gains or shares when it forms a compound. It depends on the number of valence electrons.

Down a group, valency stays the same, because the number of valence electrons is the same: all of group 1 has valency 1, all of group 2 valency 2, and the halogens valency 1 with hydrogen and metals.

Across a period, valency changes. With hydrogen it rises from 1 to 4 and falls back to 1; with oxygen it rises from 1 to 7, because an element can use all its valence electrons to combine with oxygen:

Period 3	Na	Mg	Al	Si	P	S	Cl
valence electrons	1	2	3	4	5	6	7
hydride	NaH	MgH_2	AlH_3	SiH_4	PH_3	H_2S	HCl
valency with hydrogen	1	2	3	4	3	2	1
highest oxide	Na_2O	MgO	Al_2O_3	SiO_2	P_4O_{10}	SO_3	Cl_2O_7
highest valency with oxygen	1	2	3	4	5	6	7

The noble gases have valency 0. The s-block elements keep one valency; the p-block non-metals and the d-block metals often have several (6.3).

Ionization potential

The **ionization potential** (ionization energy) of an element is the minimum energy needed to remove the most loosely held electron from one isolated atom of it, in the gas state:

$$\mathrm{X(g)} \rightarrow \mathrm{X^{+}(g)} + e^{-}$$

It is measured in **electronvolts per atom** (eV), or in **kilojoules per mole** of atoms (kJ/mol), the energy to remove one electron from each of 6.022×10^{23} atoms. The two are linked by

$$1\ \text{eV per atom} = 96.5\ \text{kJ/mol}$$

Hydrogen's is 13.6 eV, or 1,312 kJ/mol: the energy that lifts its electron from Bohr's first level to zero, outside the atom (as in Grade 9 Optional Science). A low ionization potential means an atom gives up an electron easily, as metals do.

It depends on:

- **Nuclear charge**: the more protons pulling, the higher.
- **Atomic size**: the farther the outer electron from the nucleus, the lower.
- **Shielding**: the more inner electrons in between, the lower.
- **Stable arrangements**: a full or half-filled subshell is hard to break into, so its ionization potential is a little higher than expected.

Across a period, ionization potential increases (the pull felt grows and the atom shrinks), from the alkali metal, lowest, to the noble gas, highest:

Period 3	Na	Mg	Al	Si	P	S	Cl	Ar
kJ/mol	496	738	578	787	1,012	1,000	1,251	1,521

Down a group, ionization potential decreases (the outer electron is farther out and better shielded):

Group 1	Li	Na	K	Rb	Cs
kJ/mol	520	496	419	403	376

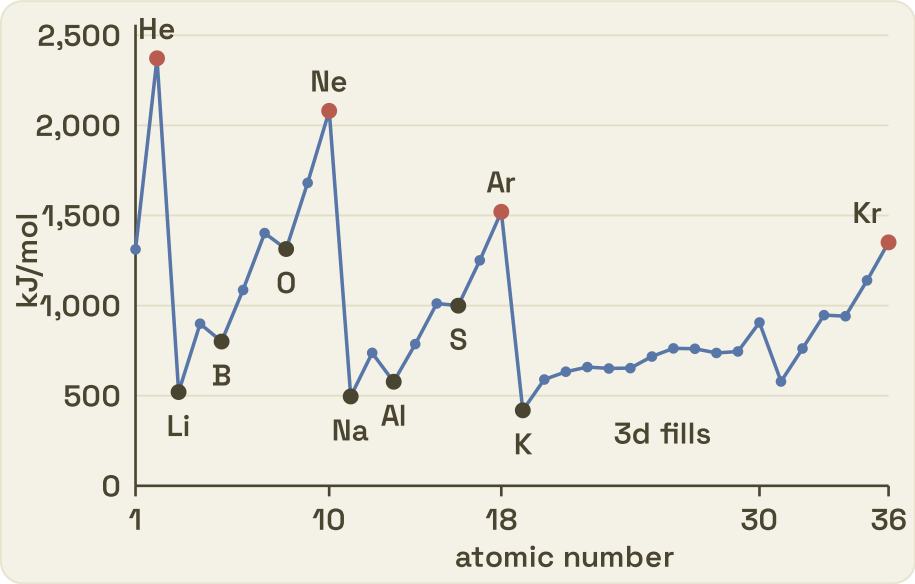


Figure 6.8 *First ionization potential of the elements 1 to 36, from measured values. Each period climbs to a peak at its noble gas (red) and falls to a trough at the next alkali metal. The small dips at boron, oxygen, aluminium and sulphur are real, and the 3d elements, scandium to copper (21 to 29), rise only gently; zinc, with full 3d and 4s, stands higher.*

Look at the two small dips in each period. Magnesium (738) is higher than aluminium (578): magnesium's outer electrons fill 3s, while aluminium's third goes into 3p, which is higher in energy and easier to remove. Phosphorus (1,012) is higher than sulphur (1,000): phosphorus has a half-filled 3p (↑ ↑ ↑), and sulphur's fourth 3p electron has to share an orbital with another, which repels it. Both dips are evidence that subshells, and Hund's rule, are real. (The same graph up to xenon, and why the subshells' shapes cause the dips, is in Going Further.)

Successive ionization potentials

After one electron is gone, a second can be removed, then a third, each needing more energy than the last, because the electrons left are held by the same protons with fewer electrons to repel them. Watch for the **huge jump**:

	1st	2nd	3rd	4th
sodium, [Ne] $3s^1$	496	**4,562**	6,910	9,543
magnesium, [Ne] $3s^2$	738	1,451	**7,733**	10,543
aluminium, [Ne] $3s^2\,3p^1$	578	1,817	2,745	**11,577**

(all in kJ/mol). Sodium's second electron must come out of the neon core, a full shell much nearer the nucleus: the energy jumps more than nine times. For magnesium the jump comes after two electrons, for aluminium after three. The jump shows how many electrons an atom can lose in a reaction, and that is its valency: 1, 2 and 3.

WORKED EXAMPLE

From eV to kJ/mol

The first ionization potential of sodium is 5.14 eV. Express it in kJ/mol, and find the energy needed to turn 2.3 g of sodium vapor into Na^+ ions. (Na = 23)

Answer

1 eV per atom is 1.602×10^{-19} J per atom, and a mole is 6.022×10^{23} atoms:

$$5.14 \times 1.602 \times 10^{-19} \times 6.022 \times 10^{23} = 4.96 \times 10^{5}\ \text{J/mol} = 496\ \text{kJ/mol}$$

(The same as $5.14 \times 96.5 = 496$ kJ/mol.)

2.3 g of sodium is $\frac{2.3}{23} = 0.1$ mol, so the energy is $0.1 \times 496 =$ **49.6 kJ**.

Electron affinity

The **electron affinity** of an element is the energy released when an electron is added to one isolated atom of it, in the gas state, forming a negative ion:

$$\mathrm{X(g)} + e^- \rightarrow \mathrm{X^-(g)} + \text{energy}$$

The more energy released, the more strongly the atom attracts an extra electron. Chlorine releases 349 kJ/mol; sodium only 53 kJ/mol.

Across a period, electron affinity generally increases: the atoms get smaller and their pull felt grows, so an incoming electron is held more strongly. The halogens, one electron short of a noble gas, release the most energy of all.

Down a group, it generally decreases: the added electron goes into a shell farther from the nucleus.

Group 17	F	Cl	Br	I
kJ/mol	328	349	325	295

Group 16	O	S	Se
kJ/mol	141	200	195

Exceptions. The first element of a group is so small that its outer shell is crowded: fluorine's 2p electrons repel the new electron more than chlorine's roomier 3p electrons do. So **chlorine has the highest electron affinity of all the elements**, higher than fluorine, and sulphur's is higher than oxygen's. And atoms whose next electron would have to start a new subshell or break a stable one, beryllium and magnesium (full s), nitrogen (half-filled p) and the noble gases (full shell), form no stable negative ion at all: adding an electron to them costs energy instead of releasing it.

> CAREFUL
> **The sign of electron affinity**
>
> Some books write electron affinity as a negative number, chlorine −349 kJ/mol, because they give the **energy change** of the process, and energy released is negative. Others give the energy released, +349 kJ/mol, as this book does. Both describe the same thing; in your answer, say "349 kJ/mol of energy is released", and you are right either way.

Electronegativity

When two atoms share a pair of electrons in a covalent bond, they don't always share it equally. **Electronegativity** is the tendency of an atom in a molecule to attract the shared pair of electrons towards itself. In hydrogen chloride, H–Cl, chlorine pulls the shared pair closer to itself, so the chlorine end is slightly negative and the hydrogen end slightly positive.

It has no unit. The American chemist **Linus Pauling** made the first scale in 1932, from the energies of bonds, giving fluorine, the most electronegative element, a value near 4. On the Pauling scale fluorine is 3.98, oxygen 3.44, chlorine 3.16, nitrogen 3.04, carbon 2.55, hydrogen 2.20, sodium 0.93 and cesium 0.79.

Across a period, electronegativity increases, and **down a group it decreases**, for the same reasons as electron affinity:

Period 2	Li	Be	B	C	N	O	F
Pauling	0.98	1.57	2.04	2.55	3.04	3.44	3.98

Group 17	F	Cl	Br	I
Pauling	3.98	3.16	2.96	2.66

So the most electronegative elements are at the top right of the table (leaving out the noble gases, which hardly form bonds) and the least at the bottom left.

> CAREFUL
> **Electron affinity and electronegativity**
>
> They sound alike and run the same way, but they are different. **Electron affinity** is an energy, measured for a lone atom gaining an electron. **Electronegativity** is a number on a scale, for an atom pulling on electrons it **shares** in a bond. That is why fluorine has the highest electronegativity but not the highest electron affinity. (Some lower-grade books describe electronegativity simply as the tendency to gain electrons; the trends are the same.)

6.5 Chemical reactivity of elements

How readily an element reacts depends on what it does in a reaction:

- A **metal** reacts by **losing** electrons. The lower its ionization potential, the more easily it loses them, and the more reactive it is.
- A **non-metal** reacts by **gaining** (or pulling on shared) electrons. The higher its electron affinity and electronegativity, the more reactive it is.

So the reactivity of metals increases down a group and decreases across a period, and the reactivity of non-metals does the opposite. This section looks at the five groups the curriculum names: 1 and 2 (metals), and 15, 16 and 17 (mostly non-metals). Many

of the single reactions of oxygen, sulphur and chlorine are in Grade 9 Optional Science; here the point is how members of a group compare.

Group 1: the alkali metals

All have one electron outside a noble gas core (ns^1), lost to form M^+. **Reactivity increases down the group**: Li < Na < K < Rb < Cs, as the ionization potential falls (520, 496, 419, 403, 376 kJ/mol).

With water, each gives a strong alkali and hydrogen:

$$2M + 2H_2O \rightarrow 2MOH + H_2 \qquad \text{for example} \qquad 2K + 2H_2O \rightarrow 2KOH + H_2$$

Lithium fizzes steadily; sodium melts into a ball and darts about; potassium catches fire at once with a lilac flame; rubidium and cesium explode.

With oxygen, they tarnish in air within seconds, which is why they are kept under oil: $4Li + O_2 \rightarrow 2Li_2O$, $4Na + O_2 \rightarrow 2Na_2O$. (Burned in plenty of air, sodium forms mainly the peroxide Na_2O_2, and potassium the superoxide KO_2.)

With chlorine, they burn to form white ionic chlorides: $2Na + Cl_2 \rightarrow 2NaCl$.

Group 2: the alkaline earth metals

All have two outer electrons (ns^2), lost to form M^{2+}. They are **less reactive than group 1** in the same period: two electrons must be removed, and the atoms are smaller, with a stronger pull. **Down the group reactivity increases**, Be < Mg < Ca < Sr < Ba, as the energy to remove the two electrons falls.

With water:

- Beryllium doesn't react, even with steam.
- Magnesium reacts very slowly with cold water and quickly when heated in steam: $Mg + H_2O \rightarrow MgO + H_2$.
- Calcium reacts steadily with cold water: $Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$.
- Strontium and barium react faster still.

With oxygen, magnesium burns with a dazzling white light, used in fireworks and old camera flashes: $2Mg + O_2 \rightarrow 2MgO$. With chlorine: $Ca + Cl_2 \rightarrow CaCl_2$.

Group 15: nitrogen to bismuth

All have five outer electrons ($ns^2\ np^3$), with a half-filled p subshell. They are three short of an octet, and show valency 3 (NH_3, PCl_3) and, except nitrogen, 5 (PCl_5).

Down the group, non-metallic character decreases and metallic character increases: nitrogen and phosphorus are non-metals, arsenic and antimony metalloids,

DID YOU KNOW?

Mountains of group 2

Group 2 metals are too reactive to be found free, but their compounds make up whole mountains. Limestone, calcium carbonate, is quarried in Udayapur, Palpa, Dang and many other districts of Nepal to make cement. Nepal also has magnesite, magnesium carbonate ($MgCO_3$): a huge deposit at Kharidhunga, near Lakuri Danda in Dolakha, a smaller one in Udayapur, and a little in Palpa and Baitadi. Nepal Orind Magnesite, set up in 1979, quarried it at Lakuri Danda and processed it in Sindhupalchok, but has been closed since 2001 (as of 2026, the government still speaks of restarting it). Heated, magnesite gives magnesium oxide, which melts only at about 2,800 °C, so it is made into the heat-proof bricks that line furnaces.

and bismuth a metal. Their oxides change the same way, from acidic (N_2O_5, P_4O_{10}) to basic (Bi_2O_3).

Their reactivity doesn't simply follow the trend, and the reasons are worth knowing:

- **Nitrogen is very unreactive.** Its two atoms are held by a triple bond, N≡N, one of the strongest bonds known (about 945 kJ/mol to break). That is why air, four-fifths nitrogen, lets things burn only in its oxygen. Nitrogen reacts only under harsh conditions: with oxygen in lightning, $N_2 + O_2 \rightarrow 2NO$, and with hydrogen at about 450 °C and 200 atm with an iron catalyst, $N_2 + 3H_2 \rightleftharpoons 2NH_3$ (the Haber process, which makes the ammonia for urea fertilizer).
- **White phosphorus is very reactive.** It is made of P_4 molecules whose bonds are bent tightly at 60°, so they break easily. It catches fire in air a little above room temperature, about 30 °C, and is stored under water: $P_4 + 5O_2 \rightarrow P_4O_{10}$. With chlorine it gives $2P + 3Cl_2 \rightarrow 2PCl_3$, then $PCl_3 + Cl_2 \rightarrow PCl_5$. Red phosphorus, with the atoms in long chains, is far safer: it is on the side of a matchbox.

Group 16: oxygen to polonium

All have six outer electrons ($ns^2\ np^4$), two short of an octet. They gain two electrons to form ions such as O^{2-} and S^{2-}, or share two (H_2O, H_2S): valency 2, and for sulphur also 4 and 6 (SO_2, SO_3).

Down the group reactivity decreases, O > S > Se > Te, and metallic character increases: oxygen and sulphur are non-metals, selenium mostly a non-metal, tellurium a metalloid and polonium a metal. Oxygen, small and the second most electronegative element, combines with almost every element, often burning: $2H_2 + O_2 \rightarrow$

$2H_2O$. Sulphur combines with hydrogen only when heated, $H_2 + S \rightarrow H_2S$, and with iron when the two are heated together, $Fe + S \rightarrow FeS$.



Figure 6.9 *A block of white phosphorus. It is kept under water so that air can't reach it.*
Photo: BXXXD, CC BY-SA 4.0

Group 17: the halogens

All have seven outer electrons ($ns^2\ np^5$), one short of an octet, which they complete by gaining one electron (X^-) or sharing one. They are the most reactive non-metals, and **reactivity decreases down the group**: F > Cl > Br > I.

Displacement. A more reactive halogen takes the place of a less reactive one in its salts:

$$Cl_2 + 2KBr \rightarrow 2KCl + Br_2 \qquad Cl_2 + 2KI \rightarrow 2KCl + I_2 \qquad Br_2 + 2KI \rightarrow 2KBr + I_2$$

Chlorine water turns potassium bromide solution orange (bromine) and potassium iodide solution brown (iodine). Iodine can't displace chlorine or bromine: $I_2 + KCl \rightarrow$ no reaction.

With hydrogen, all form hydrogen halides, ever less readily down the group:

Halogen	Reaction with hydrogen
fluorine	explodes even in the cold and dark: $H_2 + F_2 \rightarrow 2HF$
chlorine	explodes in sunlight, slow in the dark: $H_2 + Cl_2 \rightarrow 2HCl$
bromine	only when heated, or with a platinum catalyst
iodine	slowly and only partly, even when heated, and the reaction reverses: $H_2 + I_2 \rightleftharpoons 2HI$

With metals: chlorine takes iron to its higher valency, $2Fe + 3Cl_2 \rightarrow 2FeCl_3$, but iodine, a weaker oxidizer, only to the lower one, $Fe + I_2 \rightarrow FeI_2$.

Across a period

Period 3 shows all the trends in one row. From left to right, the elements change from very reactive metals, through less reactive ones, to non-metals that get more reactive,

ending in an unreactive noble gas:

Element	Kind	With water	Highest oxide, its nature
Na	metal	violently, cold: $NaOH + H_2$	Na_2O, basic
Mg	metal	slowly with cold, fast with steam	MgO, basic
Al	metal	no reaction: a thin oxide layer protects it	Al_2O_3, amphoteric
Si	metalloid	no reaction	SiO_2, weakly acidic
P	non-metal	no reaction (white P is stored under water)	P_4O_{10}, acidic
S	non-metal	no reaction	SO_3, acidic
Cl	non-metal	dissolves and partly reacts: $Cl_2 + H_2O \rightleftharpoons HCl + HOCl$	Cl_2O_7, strongly acidic
Ar	noble gas	no reaction	none

The metals' reactivity falls from sodium to aluminium, as their ionization potentials rise and they must lose more electrons. The non-metals' pull on electrons rises from phosphorus to chlorine, as their electronegativity rises (2.19, 2.58, 3.16). The same is true of groups 1 and 2 against each other (sodium is more reactive than magnesium, potassium than calcium) and of groups 15, 16 and 17 (nitrogen < oxygen < fluorine as oxidizing agents).

CAREFUL

"Reactive" for a non-metal

White phosphorus catches fire in air more easily than sulphur, yet sulphur is called the more reactive non-metal. A non-metal's reactivity, in these trends, means its power to **take** electrons from other substances (its oxidizing power). When phosphorus burns, it is oxygen that takes electrons from phosphorus; phosphorus burns easily because its strained P_4 molecules fall apart easily, not because it pulls electrons strongly.

Check 4

What happens when chlorine water is added to potassium iodide solution?

a nothing: iodine is more reactive than chlorine

b the solution turns brown as iodine is set free

c chlorine gas bubbles out

d a white precipitate of potassium chloride forms

IN SHORT

- **Modern periodic law:** properties are periodic functions of atomic number. The table has 7 periods (period = number of shells) and 18 groups; the period lengths 2, 8, 8, 18, 18, 32, 32 come from the subshells each period fills.
- **Filling rules:** Aufbau ($1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p$), Pauli (two electrons per orbital, opposite spins), Hund (one in each orbital, same spin, before pairing). Exceptions: Cr [Ar] $4s^1\ 3d^5$, Cu [Ar] $4s^1\ 3d^{10}$.
- **Blocks:** s (groups 1, 2: reactive metals, fixed valency), p (13–18: metals, metalloids, non-metals, noble gases), d (3–12: hard metals, variable valency, colored compounds, catalysts), f (lanthanides and actinides; actinides radioactive).
- **Across a period:** size decreases; ionization potential, electron affinity and electronegativity increase; valency with oxygen rises from 1 to 7.
- **Down a group:** size increases; ionization potential, electron affinity and electronegativity decrease; valency stays the same.
- **Exceptions:** $Mg > Al$ and $P > S$ in ionization potential; $Cl > F$ and $S > O$ in electron affinity.
- **Reactivity:** metals react by losing electrons (groups 1, 2: more reactive down the group, less across); non-metals by gaining them (groups 16, 17: less reactive down the group, more across). Nitrogen is unreactive because of its triple bond.

Exercise

1. Choose the correct option

a. Which rule says that electrons enter the orbitals of a subshell singly before they pair up?

a Aufbau principle

b Pauli's exclusion principle

c Hund's rule

d the octet rule

b. The electronic configuration of chromium (Z = 24) is:

- **a** $[Ar]\ 4s^2\ 3d^4$
- **b** $[Ar]\ 4s^1\ 3d^5$
- **c** $[Ar]\ 4s^2\ 4p^4$
- **d** $[Ar]\ 3d^6$

c. Which element has the largest atomic radius?

- **a** Li
- **b** Na
- **c** Mg
- **d** K

d. The element with the highest electron affinity is:

- **a** fluorine
- **b** chlorine
- **c** oxygen
- **d** neon

e. Transition elements show variable valency because:

- **a** they are metals
- **b** they have a large size
- **c** electrons from both the ns and (n − 1)d subshells can take part in bonding
- **d** they have full outer shells

2. Give reasons

a. The properties of elements repeat at regular intervals of atomic number.

b. Potassium's 19th electron goes into 4s and not 3d.

c. Copper's configuration is $[Ar]\ 4s^1\ 3d^{10}$, not $[Ar]\ 4s^2\ 3d^9$.

d. Atomic size decreases from sodium to chlorine.

e. The ionization potential of magnesium is higher than that of aluminium, although aluminium is to its right in period 3.

f. Chlorine has a higher electron affinity than fluorine.

g. Nitrogen is much less reactive than its position in the periodic table suggests.

h. The compounds of zinc are white, while those of copper are usually colored.

3. Differentiate between

i. s-block and p-block elements

ii. d-block and f-block elements

iii. Ionization potential and electron affinity

iv. Electron affinity and electronegativity

4. Answer the following questions

a. State the modern periodic law. Write any three merits of the modern periodic table over Mendeleev's.

b. Explain why the periods of the periodic table contain 2, 8, 8, 18 and 18 elements.

c. State the Aufbau principle, Pauli's exclusion principle and Hund's rule. Show the orbital diagram of oxygen.

d. Write the electronic configurations of the elements with atomic numbers 11, 16, 21 and 29, and find the block, period and group of each.

e. Write four characteristics each of s-block and d-block elements.

f. What is ionization potential? How does it vary in a period and in a group? Explain with reasons.

g. What is electronegativity? How does it vary in the periodic table? Name the most and least electronegative elements.

h. How does the reactivity of the alkali metals change down the group? Explain, with the reaction of potassium with water.

i. How does the reactivity of the halogens change down group 17? Describe one experiment that shows it, with equations.

j. Compare the elements of group 15 from nitrogen to bismuth in metallic character and reactivity.

5. Problems

a. The atomic numbers of A, B, C and D are 9, 12, 17 and 20. (i) Which two are in the same group? (ii) Which is the most reactive non-metal? (iii) Which is the most reactive metal? (iv) Write the formula of the compound formed by B and C.

b. The successive ionization potentials of an element Y, in kJ/mol, are 578, 1,817, 2,745 and 11,577. (i) How many valence electrons does Y have? (ii) To which group does it belong? (iii) Write the formula of its oxide.

c. Calculate the energy released when 3.55 g of chlorine atoms, in the gas state, each gain an electron. The electron affinity of chlorine is 349 kJ/mol. (Cl = 35.5)

d. How many unpaired electrons are there in (i) phosphorus (15), (ii) iron (26), (iii) Fe^{3+} and (iv) chromium (24)? Which of them would be attracted most strongly by a magnet?

e. Arrange in order, giving a reason in each case: (i) Na, Mg, Al, Cl by atomic size (largest first); (ii) F, Cl, Br, I by electronegativity (highest first); (iii) Li, Na, K by first ionization potential (highest first); (iv) Mg^{2+}, Na^{+}, Na, Cl^{-} by size (smallest first).

f. **Think:** Yangchen says, “Fluorine has 9 protons and lithium only 3, so fluorine’s atom must be bigger, because it has more of everything.” Explain her mistake, using the numbers.

Found a mistake? Please tell us.