LOGIC BLOOM · CLASS 11 CHEMISTRY

Classification of Elements & Periodicity

Class 11 · **Chapter 3** · Reading the Periodic Table

From Doebereiner's triads to the modern law, then every trend the table hides: atomic and ionic radii, ionisation and electron gain enthalpy, electronegativity, and the way oxides swing from basic to acidic. Built around a full-colour periodic table, with the exceptions examiners love flagged in red.

NCERT Chapter 3 | Full Periodic Table | Every Trend | Anomalies Flagged | 1-Page Fact Sheet

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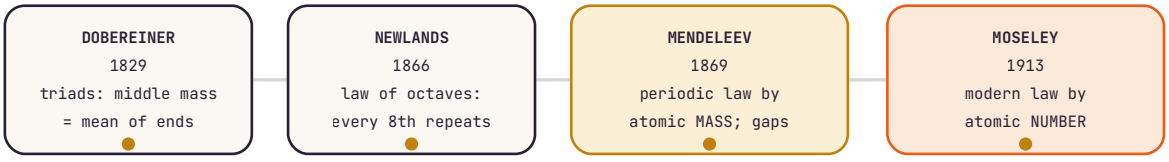
Contents

1 The Search for Order

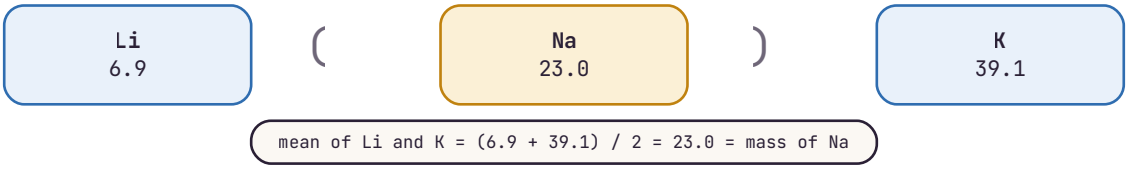
By the 1860s chemists knew over sixty elements with no system tying them together. The story of the periodic table is the story of spotting a repeating pattern, first in atomic mass, finally in atomic number.



Four steps to the modern table. Each thinker saw the pattern a little more clearly than the last.

Doebereiner's triads (1829)

Johann Doebereiner grouped elements in threes where the middle element's atomic mass was roughly the **average** of the other two, and its properties sat in between.



A triad: the atomic mass of sodium is almost exactly the mean of lithium and potassium. Neat, but only a handful of triads exist.

Triad	Atomic masses	Mean of ends
Li, Na, K	6.9, 23.0, 39.1	23.0
Ca, Sr, Ba	40, 88, 137	88.5
Cl, Br, I	35.5, 80, 127	81.2

Newlands' law of octaves (1866)

John Newlands arranged elements by rising mass and found every **eighth** element repeated properties, like the eighth note of a musical scale. It worked up to calcium, then broke down, and was ridiculed at the time.

Mendeleev's periodic table (1869)

Dmitri Mendeleev's leap was to arrange by atomic mass but **prioritise properties**: he left gaps for undiscovered elements and even predicted their properties. When gallium and germanium were later found matching his "eka-aluminium" and "eka-silicon", the table's power was undeniable.

Mendeleev predicted eka-silicon in 1871; germanium, discovered in 1886, matched almost exactly. Prediction, not just description.

★ MENDELEEV'S PERIODIC LAW

The properties of elements are a **periodic function of their atomic masses**. His table's triumphs were the gaps and predictions; its flaws were misplaced pairs (like Ar and K) and no fixed spot for isotopes or hydrogen.

The modern periodic law (Moseley, 1913)

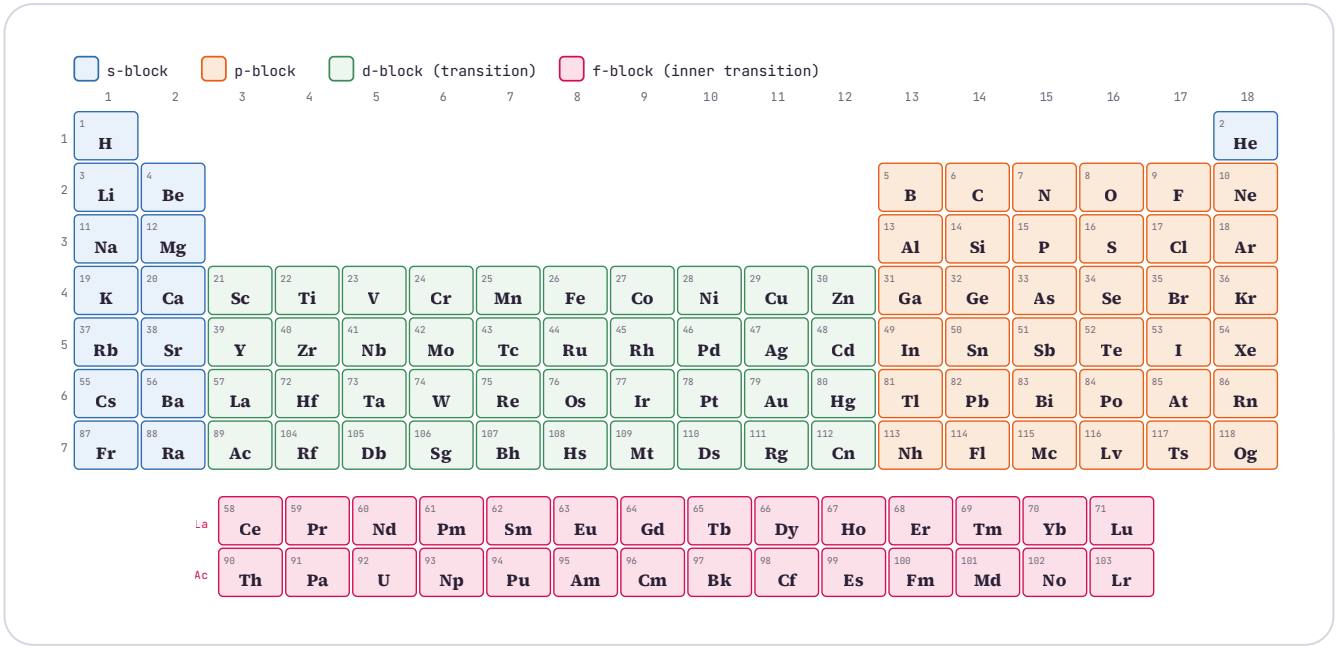
Henry Moseley showed from X-ray studies that **atomic number** (nuclear charge), not mass, is the true basis of order. This fixed the anomalous pairs at a stroke.

★ MODERN PERIODIC LAW

The properties of elements are a **periodic function of their atomic numbers (Z)**. Atomic number is fundamental because it fixes the number of electrons, and chemistry is the chemistry of electrons.

2 Building the Modern Table

The modern table has **7 horizontal periods** and **18 vertical groups**. Elements in a group share the same outer-shell configuration, so they behave alike, that is the whole point of the arrangement.



The full modern periodic table, coloured by block. Groups run 1 to 18; the f-block (lanthanides and actinides) is pulled out below to keep the table compact.

A **period number equals the principal quantum number (n)** of the outermost shell being filled. The number of elements in each period is set by the orbitals available at that energy level.

Period	n	Orbitals filled	Elements
1	1	1s	2
2	2	2s 2p	8
3	3	3s 3p	8
4	4	4s 3d 4p	18
5	5	5s 4d 5p	18
6	6	6s 4f 5d 6p	32

WHY 2, 8, 8, 18, 32

Each period fills a new set of orbitals. Period 4 adds the 3d set (10 elements), so it jumps to 18; period 6 adds the 4f set (14 elements), so it jumps to 32. The pattern is just orbital counting.

Naming elements beyond 100

Before an element is officially named, IUPAC uses a systematic name built from numerical roots for its atomic number, followed by "ium".

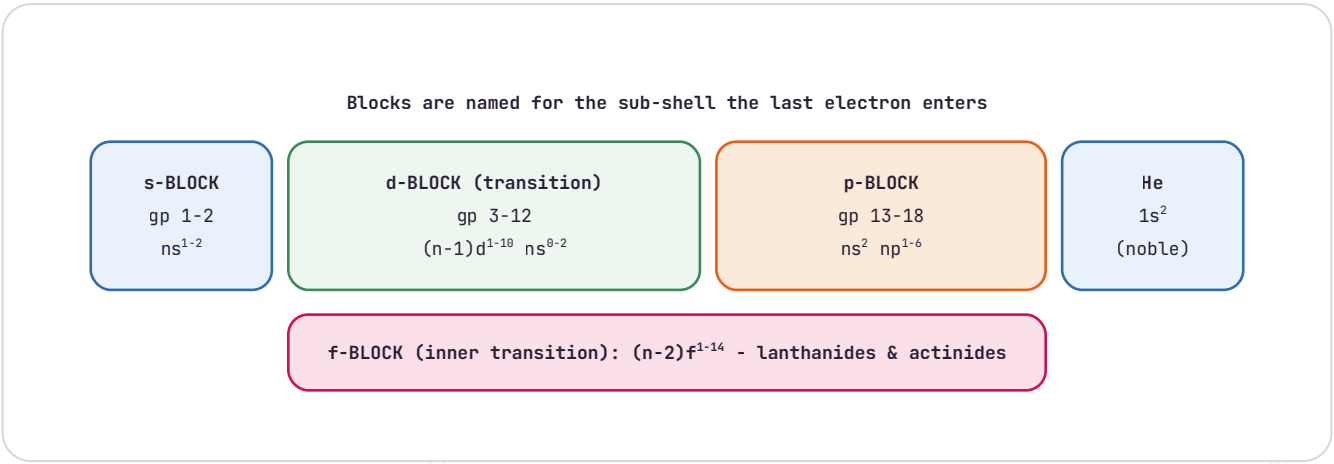
0	1	2	3	4	5	6	7	8	9
nil	un	bi	tri	quad	pent	hex	sept	oct	enn

WORKED EXAMPLE

Element 120: 1-2-0 gives un-bi-nil, so the name is **Unbinilium**, symbol Ubn. Element 104 is un-nil-quad, **Unnilquadium** (Unq), before it was named Rutherfordium.

3 Blocks and Electronic Configuration

The table splits into four **blocks**, each named for the sub-shell that receives the last electron. The block instantly tells you an element's outer configuration.



Four blocks. Reading the block off the table gives you the valence-shell configuration without memorising it.

Block	Groups	Valence config	Character
s	1, 2 (and He)	ns^{1-2}	reactive metals
p	13 to 18	ns^2np^{1-6}	metals to non-metals
d	3 to 12	$(n-1)d^{1-10}ns^{0-2}$	transition metals
f	La, Ac series	$(n-2)f^{1-14}$	inner transition

★ THE SHORTCUT

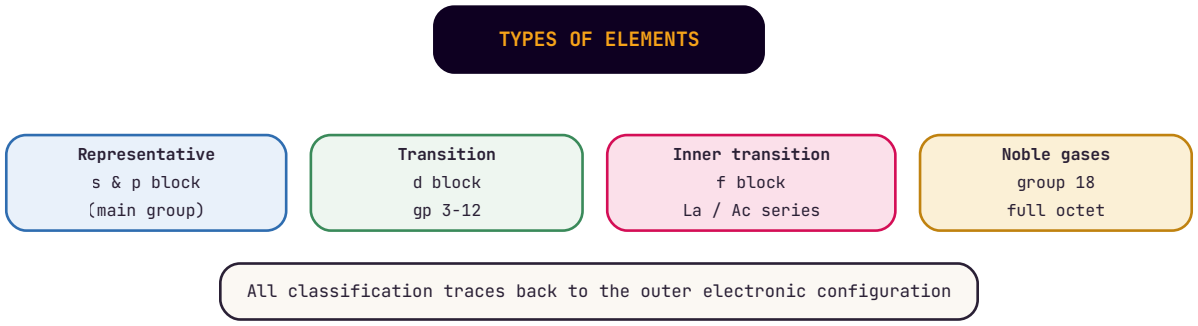
Group number tells you valence electrons: for s-block it is the group number; for p-block it is (group number minus 10). So group 15 has 5 valence electrons (ns^2np^3).

⚠ WATCH HELIUM

Helium is $1s^2$, an s-block configuration, yet it sits in group 18 with the noble gases because chemically it behaves as one (full shell). Position follows behaviour, not just the block.

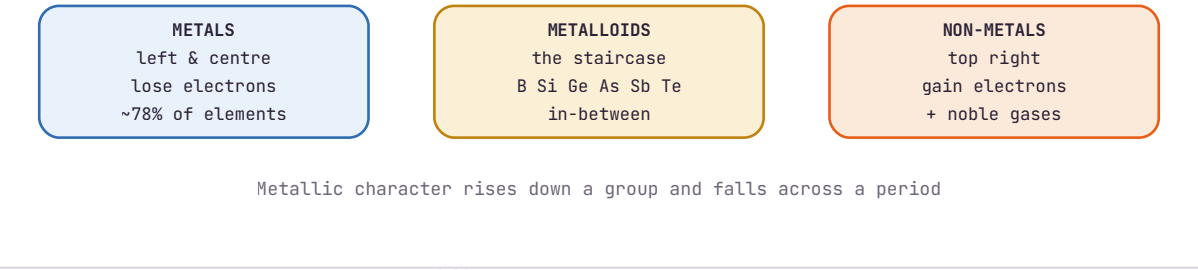
4 Types of Elements

Beyond blocks, elements fall into four behavioural families, all decided by the outer configuration.



The four families. Representative elements have partially filled outer shells; noble gases have complete ones.

Across any period, **metallic character falls** and non-metallic character rises. The dividing "staircase" runs through the metalloids, elements with in-between properties.



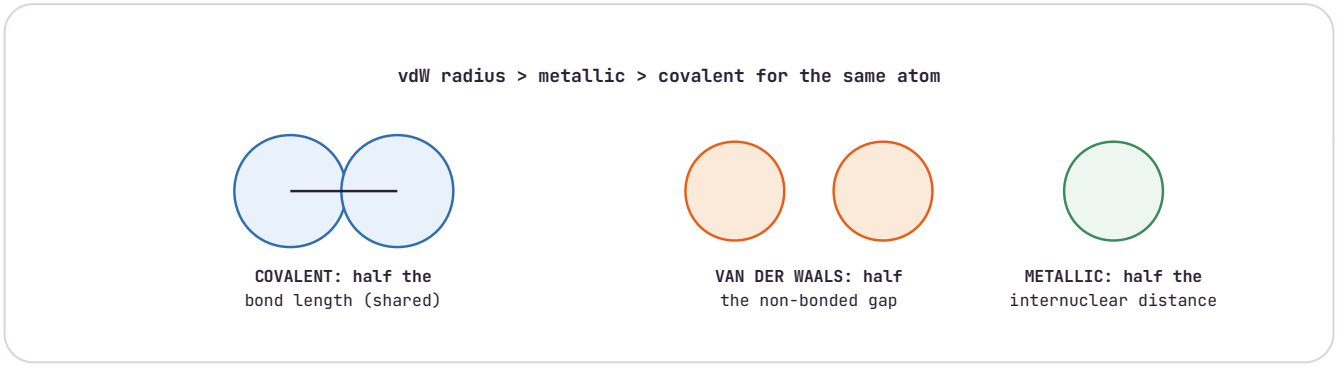
Metals dominate the left and centre, non-metals the top right, with a thin staircase of metalloids between them.

METALS VS NON-METALS

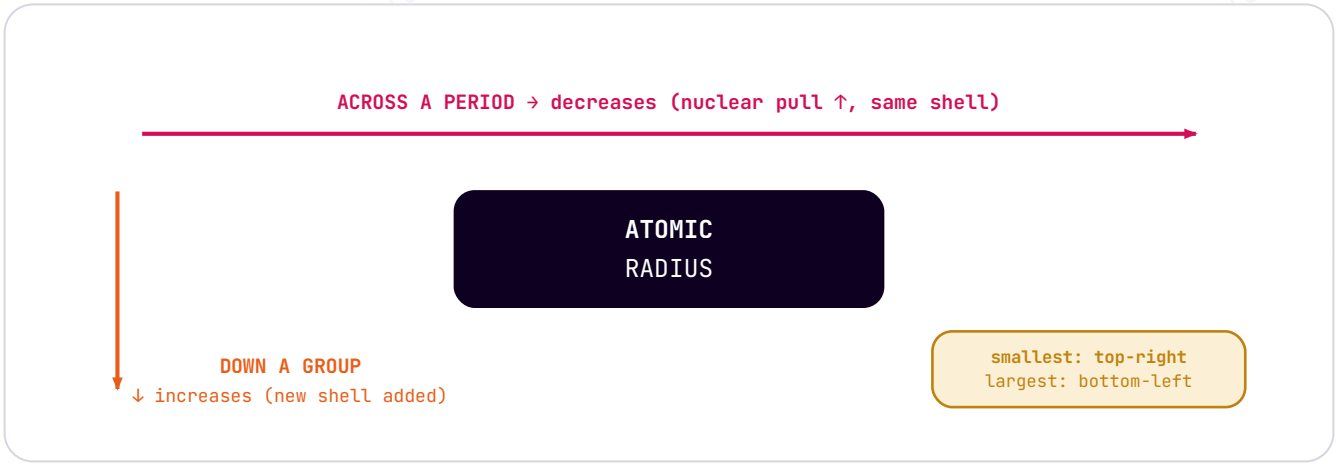
Metals are large, have low ionisation enthalpy and lose electrons to form cations and basic oxides. Non-metals are small, have high ionisation enthalpy and gain electrons to form anions and acidic oxides. Metalloids (B, Si, Ge, As, Sb, Te) straddle the line.

5 Atomic and Ionic Radii

Atoms have no hard boundary, so "radius" is defined by context. Three common measures come from how the atom is bonded to its neighbour.



Three operational radii. For the same atom, van der Waals radius is largest because non-bonded atoms sit further apart than bonded ones.



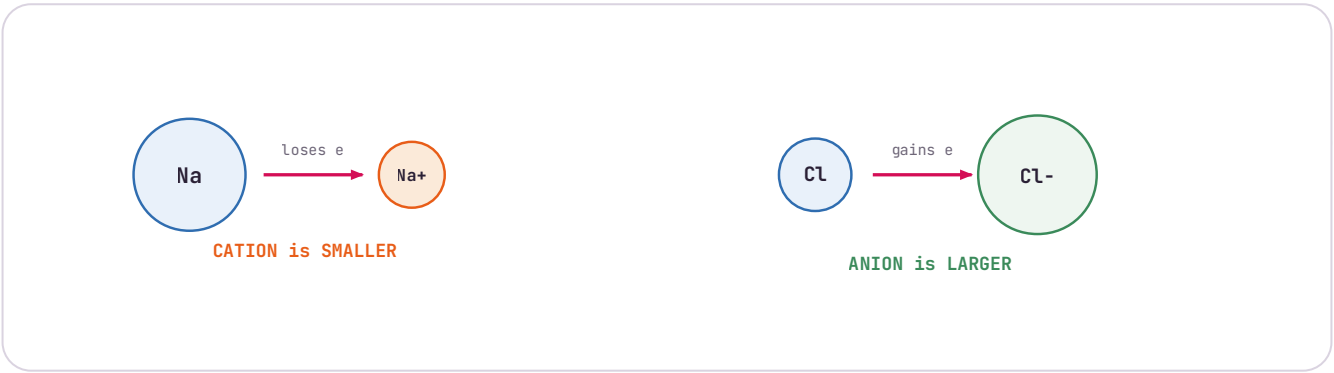
The master trend for size: atomic radius falls across a period as nuclear pull tightens the same shell, and grows down a group as new shells are added.

★ SIZE TREND

Across a period radius decreases: protons increase but electrons enter the same shell, so the stronger nuclear pull contracts the atom. **Down a group** radius increases: each period adds a whole new shell, outweighing the extra nuclear charge.

Ionic radii

Losing electrons shrinks an atom; gaining them swells it.



A cation is smaller than its parent atom (fewer electrons, often one shell lost); an anion is larger (added electron-electron repulsion).

For **isoelectronic species** (same number of electrons), size is decided purely by nuclear charge: more protons pull the same electron cloud in tighter.

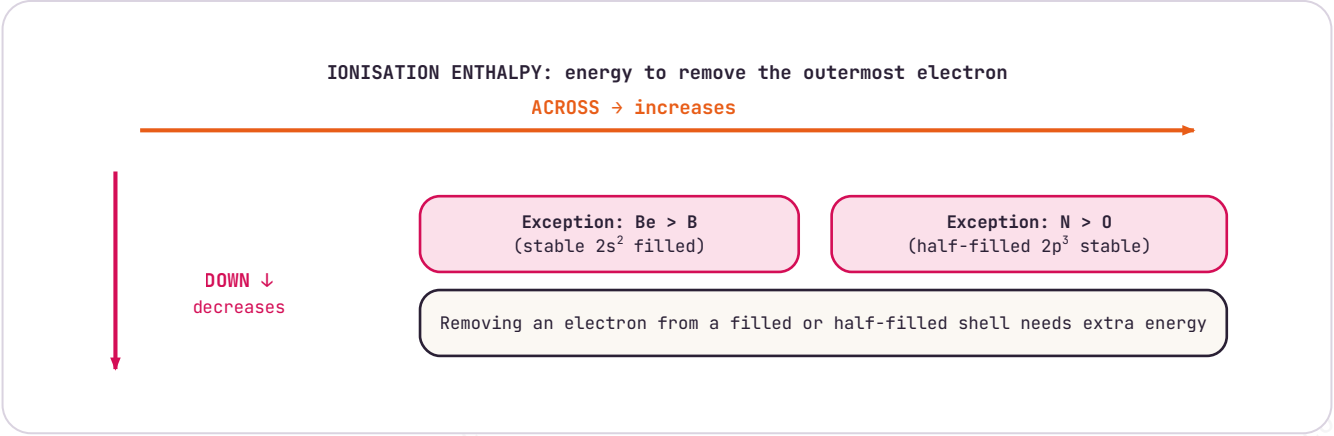
Species	N^{3-}	O^{2-}	F^{-}	Na^{+}	Mg^{2+}	Al^{3+}
Electrons	10	10	10	10	10	10
Protons	7	8	9	11	12	13
Radius (pm)	171	140	133	95	65	50

⚠ ISOELECTRONIC TRAP

All six species above have 10 electrons, yet their sizes differ hugely. More protons means a smaller ion, so Al^{3+} is the smallest and N^{3-} the largest. Never assume same electrons means same size.

6 Ionisation Enthalpy

Ionisation enthalpy is the energy needed to pull the outermost electron off an isolated gaseous atom. Successive removals get harder, so $IE_1 < IE_2 < IE_3$, because you are stripping a positive ion.



Ionisation enthalpy rises across a period and falls down a group, with two famous dips where a stable configuration resists.

★ IE TRENDS AND THEIR CAUSES

Across a period IE rises (smaller atom, higher nuclear charge hold electrons tighter). **Down a group** IE falls (bigger atom, more shielding, the outer electron is far and loosely held).

⚠ THE TWO EXCEPTIONS YOU WILL BE TESTED ON

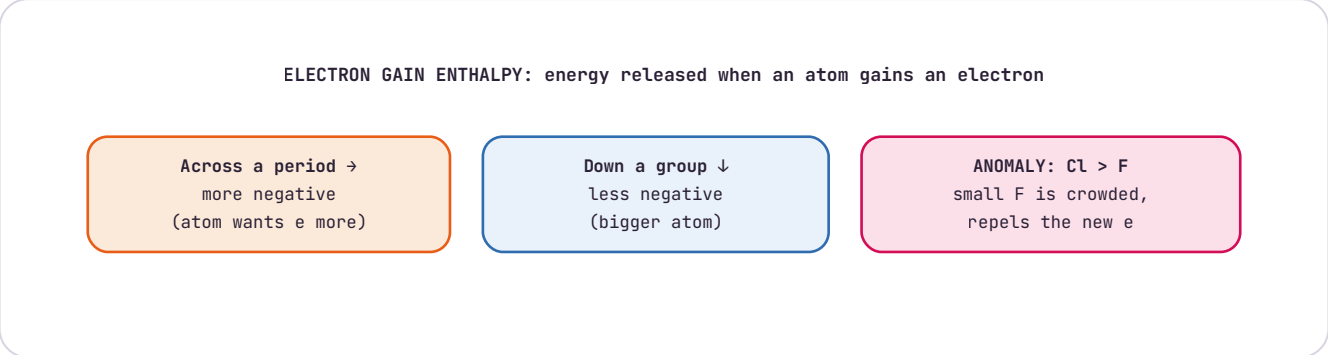
Be > B: boron's 2p electron is easier to remove than beryllium's stable filled $2s^2$. **N > O:** nitrogen's half-filled $2p^3$ is extra stable, so oxygen loses its electron more easily. Half-filled and fully-filled shells are stubborn.

WHAT CONTROLS IE

Four factors: nuclear charge (up raises IE), atomic size (up lowers IE), shielding by inner electrons (up lowers IE), and the stability of the resulting configuration (filled or half-filled shells resist).

7 Electron Gain Enthalpy

Electron gain enthalpy is the energy change when a gaseous atom accepts an electron. It is usually **negative** (energy released) because most atoms are happy to gain an electron toward a full shell.



Electron gain enthalpy grows more negative across a period, less negative down a group, with a famous anomaly at the top of group 17.

★ EGE TRENDS

Across a period it becomes more negative (smaller atoms with higher nuclear charge grab electrons eagerly). **Down a group** it becomes less negative (larger atoms hold the new electron weakly).

⚠ CHLORINE BEATS FLUORINE

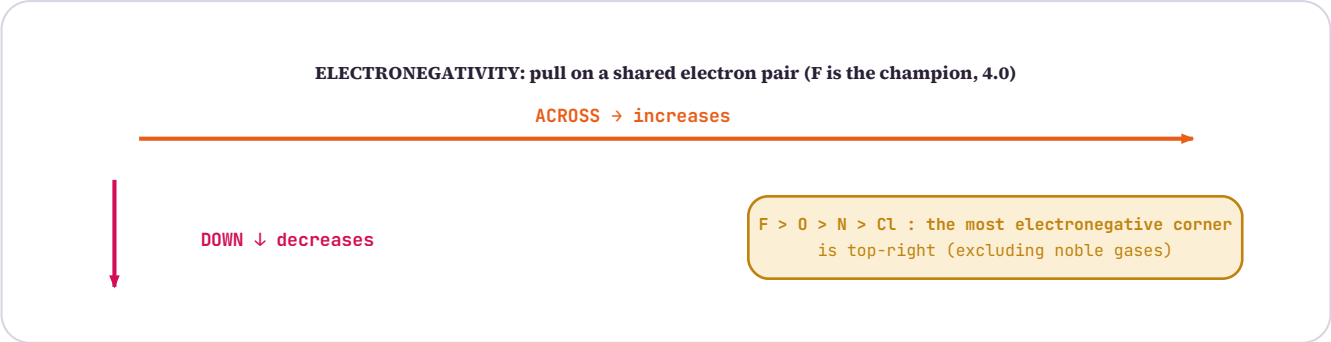
You would expect fluorine to have the most negative electron gain enthalpy, but **chlorine does**. Fluorine is so small that its 2p shell is crowded; the incoming electron feels strong repulsion. Chlorine's larger 3p shell has room, so it releases more energy. The same logic explains O < S.

SIGN CONVENTIONS

Noble gases have **positive** electron gain enthalpy (they resist adding an electron to a full shell). Adding a second electron is always endothermic because you force a negative electron onto a negative ion.

8 Electronegativity

Electronegativity is the tendency of an atom in a **bond** to pull the shared electron pair toward itself. On the Pauling scale fluorine is the champion at 4.0.



Electronegativity rises across a period and falls down a group, so the most electronegative elements cluster at the top right.

★ EN TREND

It increases across a period (smaller atoms pull harder) and decreases down a group (larger atoms pull less). The order F > O > N > Cl is worth memorising.

NOT THE SAME AS ELECTRON GAIN ENTHALPY

Electron gain enthalpy is a measurable energy for an **isolated** atom gaining an electron. Electronegativity is a relative, unitless property of an atom **already in a bond**. They trend the same way but describe different situations.

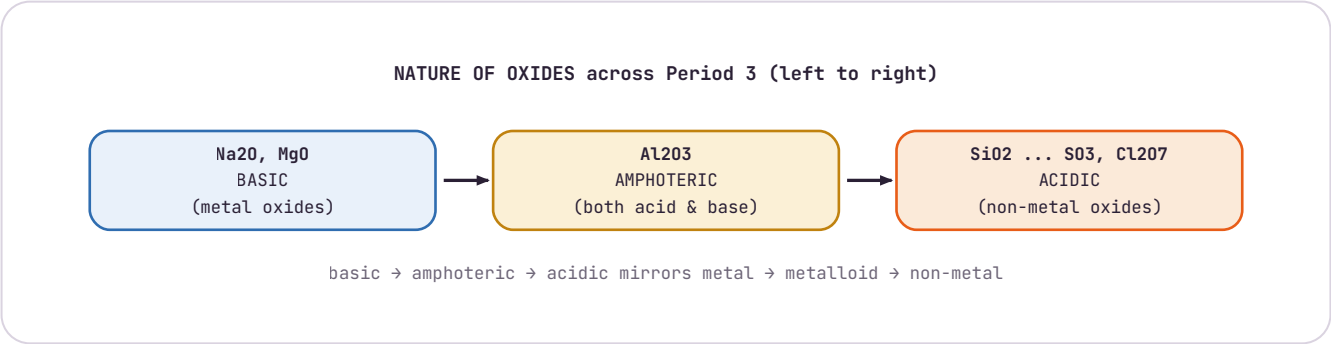
9 Periodicity in Chemical Properties

Physical trends drive chemical behaviour: valence, the acid-base nature of oxides, and reactivity all repeat periodically.

Valence

Valence toward hydrogen rises then falls across a period (1, 2, 3, 4, 3, 2, 1), while valence toward oxygen often rises to a maximum equal to the group number.

Nature of oxides



Across period 3 the oxides swing from basic (metals) through amphoteric (aluminium) to acidic (non-metals), mirroring the metal to non-metal shift.

★ OXIDE RULE

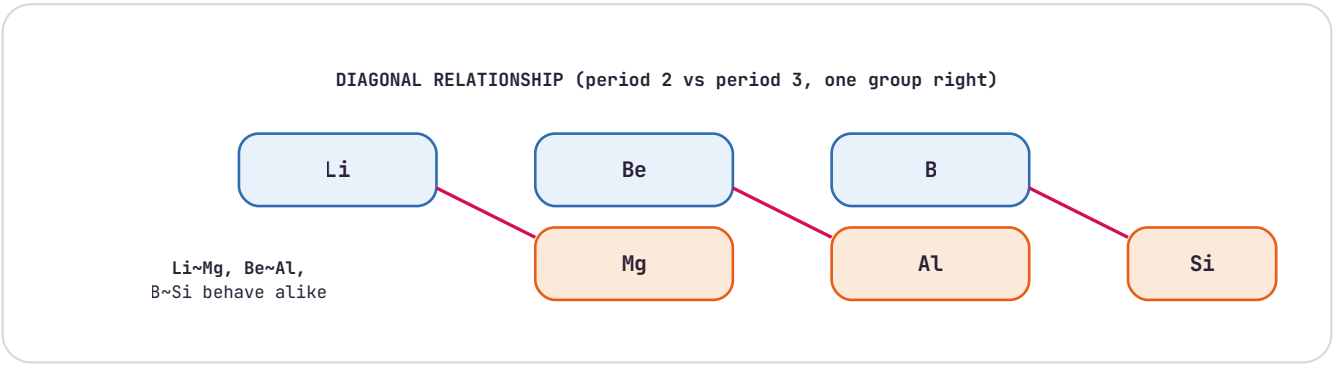
Metal oxides are **basic**, non-metal oxides are **acidic**, and the borderline oxide (Al_2O_3, also ZnO) is **amphoteric**, reacting with both acids and bases. This is a direct read-out of position in the period.

⚠ THE NEUTRAL-OXIDE TRAP

Not every non-metal oxide is acidic. **CO, NO and N_2O** are **neutral**: they form no salt with either acids or bases. Examiners slip these in among the acidic oxides, so keep the short neutral list handy.

Reactivity and anomalies

Metals get more reactive down a group (electrons lost more easily); non-metals get more reactive up a group (electrons gained more easily). The **second-period elements** (Li to F) are anomalous: they are unusually small and have no d-orbitals, so they differ from the rest of their groups.



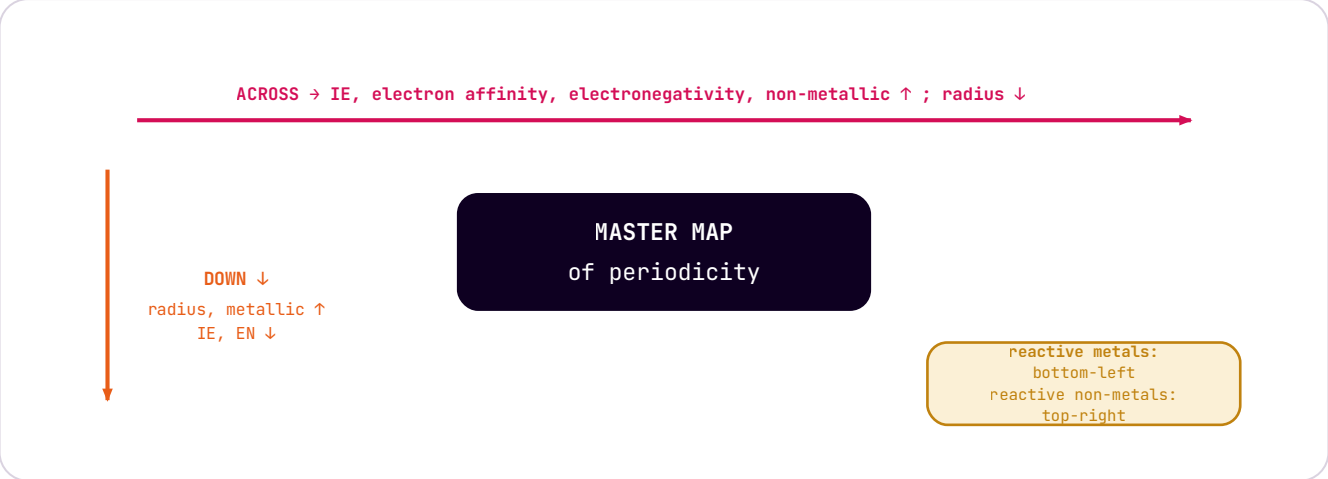
The diagonal relationship: an element resembles the one diagonally below-right because moving right and down have opposing, cancelling effects on size and charge.

⚠ DIAGONAL RELATIONSHIP

Lithium resembles magnesium, beryllium resembles aluminium, and boron resembles silicon, more than they resemble their own group-mates. It happens because going across (size down) and going down (size up) roughly cancel, leaving similar polarising power.

10 The Master Map

Every trend in this chapter reduces to two directions of travel. Fix this one picture and you can reconstruct the rest in the exam.



One map for the whole chapter. Read left-to-right for the properties that rise, top-to-bottom for the ones that grow with size.

Property	Across a period →	Down a group ↓	Driver
Atomic radius	decreases	increases	nuclear pull vs new shells
Ionisation enthalpy	increases	decreases	hold on outer electron
Electron gain enthalpy	more negative	less negative	eagerness to gain
Electronegativity	increases	decreases	pull in a bond
Metallic character	decreases	increases	ease of losing electrons
Non-metallic character	increases	decreases	ease of gaining electrons

THE ONE-LINE SUMMARY

Toward the **top-right**: small, high IE, high electronegativity, non-metallic. Toward the **bottom-left**: large, low IE, metallic and reactive. Almost every question is a variation on this.

★ Periodicity · One-Page Fact Sheet

Everything for revision day. Print this page alone.

THE TWO LAWS

Mendeleev: periodic function of atomic MASS.
Modern (Moseley): of atomic NUMBER (Z).

THE TABLE

7 periods, 18 groups.
Period no. = valence n.
2, 8, 8, 18, 18, 32 elements.

BLOCKS

s (gp 1-2), p (13-18),
d (3-12, transition),
f (inner transition).

ATOMIC RADIUS

Across → decreases.
Down ↓ increases.
vdW > metallic > covalent.

IONS

Cation < atom < anion.
Isoelectronic: more protons = smaller ion.

IONISATION

Across ↑, down ↓.
Exceptions: Be > B,
N > O (stable shells).

ELECTRON GAIN

Across more -ve, down less.
Anomaly: Cl > F
(F too small, crowded).

ELECTRONEGATIVITY

Across ↑, down ↓.
F = 4.0 is highest.
Order F > O > N > Cl.

OXIDES & DIAGONAL

Basic → amphoteric → acidic.
Al_2O_3 amphoteric; CO, NO, N_2O neutral.
Diagonal: Li~Mg, Be~Al.

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★ TREND NOT STICKING?

Play it. Then practice it.

If a periodic trend or exception here didn't land, open the chapter in the Logic Bloom app: play it as a game, then drill practice questions till it clicks.

SCAN → PLAY THE GAME · PRACTICE NOW