JEE MAIN · PAPER 1 · CHEMISTRY

09 Classification of Elements & Periodicity

Formula & concept sheet

How the periodic table is organised and why properties repeat: periods, groups and blocks from electron configuration, then atomic and ionic radius, ionisation enthalpy, electron gain enthalpy, electronegativity and valence, with every exception JEE asks, the diagonal relationship and the nature of oxides. Each idea is explained first, then the tables, shortcuts and traps.

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READ ALONGSIDE NCERT XI · Classification of Elements and Periodicity in Properties

HOW TO USE THIS SHEET

1. **Skim** — Read the Key ideas, then the formula cards.
2. **Memorise** — Learn the formulas and check every Trap box.
3. **Test** — Do the worked examples, then this chapter's PYQs.
4. **Revise** — Use the last-minute checklist before the exam.

Formula memorise · **Key idea** why it works · **Trap** marks lost here · **PYQ** how it is asked

CHAPTER 09

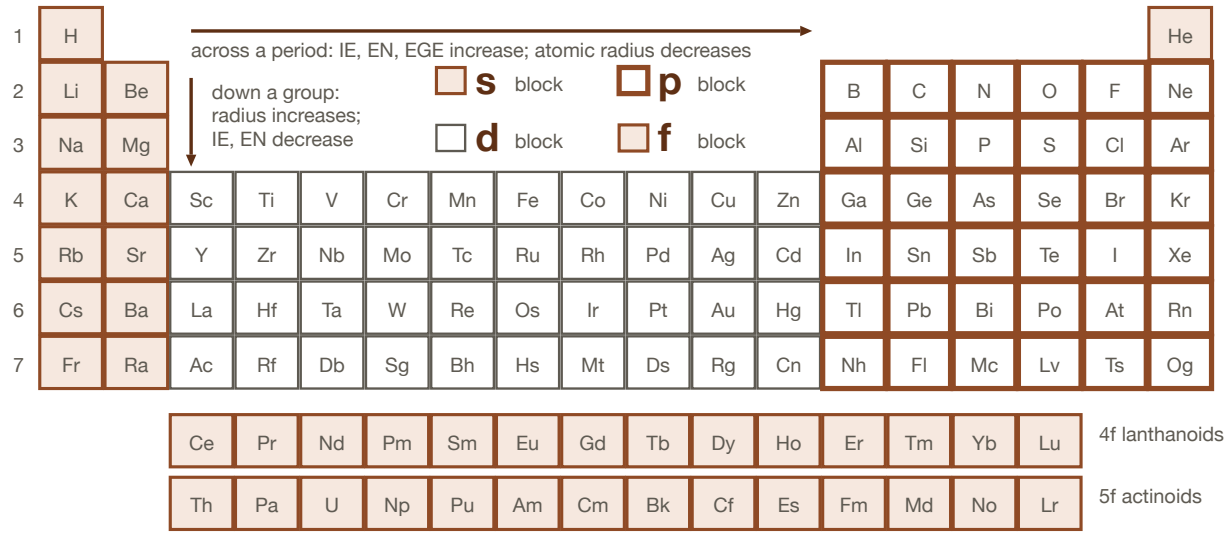
Classification of Elements & Periodicity

NCERT XI · Classification of Elements and Periodicity in Properties

Covers: modern periodic law, blocks · location from electron configuration · IUPAC names of $Z > 100$ · atomic and ionic radii, isoelectronic species · ionisation enthalpy and its exceptions · electron gain enthalpy · electronegativity · valence and oxidation state, inert-pair effect · anomalous behaviour of second-period elements, diagonal relationship · acidic and basic oxides.

1 Structure of the periodic table

- **Modern periodic law**: the properties of elements are periodic functions of their **atomic numbers** (Mendeleev used atomic masses). Properties repeat because the outer electron configuration repeats.
- The table has 7 **periods** (rows) and 18 **groups** (columns). The period number is the highest principal quantum number n of the occupied shells; periods hold $2, 8, 8, 18, 18, 32, 32$ elements (filling $1s$; $2s2p$; $3s3p$; $4s3d4p$; $5s4d5p$; $6s4f5d6p$; $7s5f6d7p$).
- Elements of one group have the same valence configuration and similar chemistry. The **f-block** (lanthanoids, actinoids) is written separately so the table stays compact and similar elements stay together.



periodic table by block: IE, EN and EGE rise across a period and fall down a group, radius does the opposite

LOCATING AN ELEMENT FROM ITS CONFIGURATION

BLOCK	LAST ELECTRON ENTERS	GROUP NUMBER	EXAMPLE
s	ns	number of ns electrons (1 or 2)	$[\mathrm{Ne}]3s^1$: period 3, group 1
p	np	$10 + (ns + np)$ electrons	$3s^2 3p^5$: period 3, group 17
d	$(n-1)d$	$ns + (n-1)d$ electrons (3 to 12)	$3d^5 4s^1$ (Cr): group 6, period 4
f	$(n-2)f$	placed in group 3, period 6 or 7	$4f$: lanthanoids; $5f$: actinoids

IUPAC NAMES FOR $Z > 100$

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

READING THE ROOTS

- Join the roots of the digits and add "ium": $Z = 119$ is un-un-enn-ium, **ununennium**, symbol Uue.
- The symbol is the first letters of the roots: $Z = 111$ is un-un-un-ium, **unununium**, Uuu; $111 - 100$ puts it in group 11 (Rg today).
- Names run from $Z = 101$: $Z = 104$ is unnilquadium in old texts, now rutherfordium.

KEY IDEA

Properties repeat because the **valence-shell configuration** repeats, not because the atomic number does. The noble gases (ns^2np^6) close each period; every trend in this chapter follows from the nuclear charge felt by the outer electron and how far it is from the nucleus.

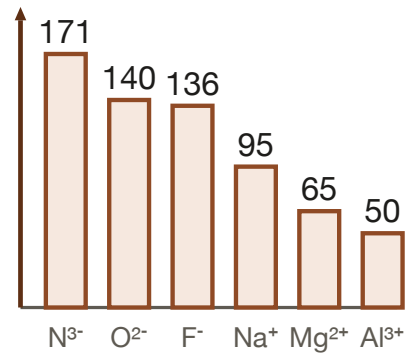
QUICK CHECK

Period, group and block of the element with configuration $[\mathrm{Ar}]3d^{10}4s^2 4p^3$?

Answer: Highest $n = 4$: period 4; $10 + (2 + 3) = 15$; p-block: arsenic.

2 Atomic and ionic radii

- The atomic radius is half the distance between nuclei: **covalent** radius for bonded atoms, **metallic** radius in a metal, **van der Waals** radius for non-bonded atoms (always larger; the "radius" of a noble gas is of this kind).
- Radius depends on the number of shells and on the **effective nuclear charge** $Z_{eff} = Z - \sigma$, where σ is the screening by inner electrons. Across a period the shell stays the same while Z_{eff} rises (by about 0.65 per element: each added proton adds 1, the extra same-shell electron screens only about 0.35, Slater), so atoms shrink. Down a group a new shell is added and the radius grows.



same 10 electrons, more nuclear charge, smaller ion (radius in pm)

TRENDS AND EXCEPTIONS IN SIZE

FACT	DETAIL
across a period	decreases: Li 152, Be 111, B 88, C 77, N 74, O 66, F 64 pm
down a group	increases: Li 152, Na 186, K 231, Rb 244, Cs 262 pm (NCERT values)
Ga smaller than Al	Ga 135 against Al 143 pm: the 10 poorly shielding $3d$ electrons raise Z_{eff}
4d, 5d same size	Zr $\approx$ Hf, Nb $\approx$ Ta: the lanthanoid contraction offsets the extra shell
cation	smaller than its atom (fewer electrons, same nuclear charge); the higher the charge, the smaller: $\mathrm{Fe^{3+} < Fe^{2+} < Fe}$
anion	larger than its atom (more repulsion): $\mathrm{Cl^- > Cl}$

ISOELECTRONIC SPECIES

$$r \propto \frac{1}{Z}$$

same electrons, more protons pull harder: $\mathrm{N^{3-} > O^{2-} > F^- > Ne > Na^+ > Mg^{2+} > Al^{3+}}$

POLARISING POWER

$$\phi = \frac{\text{charge}}{\text{radius}}$$

high ϕ gives covalent character (Fajans) and less thermally stable oxo-salts: $\mathrm{Li^+}$, $\mathrm{Be^{2+}}$, $\mathrm{Al^{3+}}$

TRAP

Noble-gas radii in tables are van der Waals values, so they look larger than the halogen next to them. In a transition series the radius is almost constant in the middle (the d electrons screen each other poorly) and the whole d and f block shows **contraction** that reaches the third row.

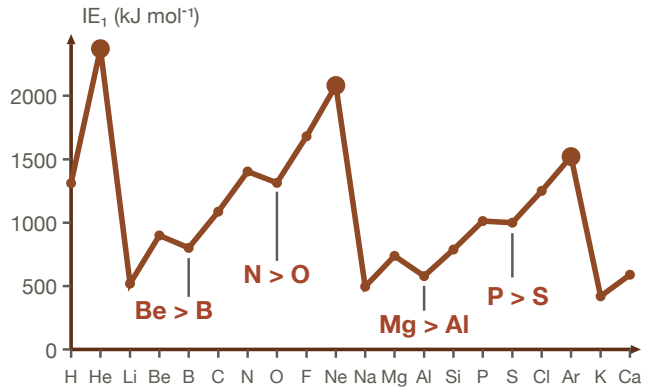
QUICK CHECK

Order Na^+, Mg^{2+}, F^-, O^{2-} by size.

Answer: All have 10 electrons: $O^{2-} > F^- > Na^+ > Mg^{2+}$ (fewest protons largest).

3 Ionisation enthalpy

IONISATION ENTHALPY The energy needed to remove the outermost electron from an isolated gaseous atom, $M(g) \longrightarrow M^+(g) + e^-$, always positive. Successive enthalpies rise: $IE_1 < IE_2 < IE_3$, because the same nuclear charge holds fewer electrons. 1 eV per atom $= 96.49$ kJ mol^{-1}.



IE rises across a period and drops at each new shell; the dips come from filled and half-filled subshells

GENERAL TREND AND ITS REASONS

- Increases across a period (higher Z_{eff}, smaller atom); noble gases are highest, He the highest of all.
- Decreases down a group (bigger atom, more shielding).
- A **filled** or **half-filled** subshell is extra stable, so removing from it costs more.
- s electrons are held more tightly than p of the same shell (penetration).

EXCEPTIONS AND ORDERS TO MEMORISE

ITEM	ORDER OR FACT	REASON
Be and B	$Be > B$ (900, 801)	B loses a $2p$ electron, easier than a filled $2s$
Mg and Al	$Mg > Al$ (738, 578)	same
N and O	$N > O$ (1402, 1314)	N has a half-filled $2p^3$; O has a paired electron to remove
P and S	$P > S$ (1012, 1000)	same
period 2	$Li < B < Be < C < O < N < F < Ne$	both dips
period 3	$Na < Al < Mg < Si < S < P < Cl < Ar$	both dips
group 13	$B > Tl > Ga > Al > In$	poor d shielding (Ga) and f shielding (Tl)
group 14	$C > Si > Ge > Pb > Sn$ ($1086, 786, 762, 716, 708$)	f electrons (Pb) shield badly
group 1, 2	Li > Na > K > Rb > Cs; Be > Mg > Ca > Sr > Ba	regular fall
d block	Cr (653) < Mn (717); IE_3: Cr 2987 < Mn 3248	Cr loses $4s^1$ to reach a stable $3d^5$ (low IE_1); Mn^{2+} is $3d^5$, so Mn has a large IE_3

GROUP FROM SUCCESSIVE IONISATION ENTHALPIES

1. List $IE_1, IE_2, \ldots$ and find the first ratio IE_{k+1}/IE_k that is far larger than the others (about 5 or more).
2. The jump after IE_k means k valence electrons: $k = 1, 2$ gives group $1, 2$; otherwise group $10 + k$ for a p-block element.
3. Check with $800, 2427, 3658, 25024$ kJ mol^{-1}: the jump is after the third, so group 13 (boron, $IE_1 = 801$).

TRAP

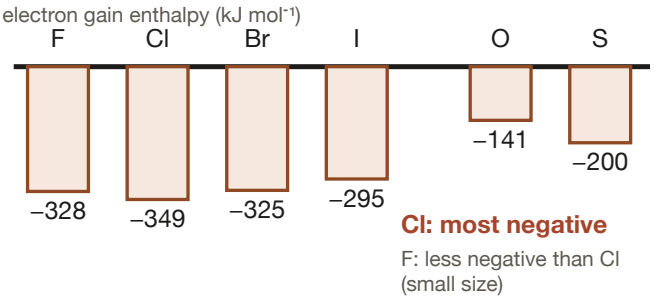
Ionisation enthalpy applies to isolated gaseous atoms; do not compare it with electrode potential: Li has the lowest E° (hydration energy is huge) but Cs has the lowest IE_1. IE_1 of Pb is higher than Sn but Ge is lower than Si; both are asked.

QUICK CHECK

Energy to ionise 0.1 mol of gaseous Na atoms ($IE_1 = 496$ kJ mol^{-1})? And Na to Na^+ in eV per atom?

Answer: $0.1 \times 496 = 49.6$ kJ; $\dfrac{496}{96.49} = 5.14$ eV per atom.

4 Electron gain enthalpy and electronegativity



second-period elements gain electrons less readily than the third-period ones in the same group (F less than Cl, O less than S)

ELECTRON GAIN ENTHALPY The enthalpy change when a gaseous atom takes up an electron, $X(g) + e^- \longrightarrow X^-(g)$. It is **negative** (energy released) for most atoms; "more negative" means more readily gained. Noble gases, N, Be, Mg have values near zero or positive because their orbitals are full or half-full. The second electron gain ($O^- \rightarrow O^{2-}$) is positive.

TREND AND EXCEPTIONS

FACT	DETAIL
across a period	becomes more negative (smaller atoms, higher Z_{eff})
down a group	less negative, except that the **second-period** elements are lower than the third: Cl $>$ F $>$ Br $>$ I ($-349, -328, -325, -295$ kJ mol^{-1})
why F is low	the very small $2p$ shell repels the incoming electron; same for O against S (-141 against -200)
near zero or positive	Be, Mg (ns^2), N ($2p^3$), noble gases
ionic character	more negative EGE of the anion-forming atom gives a more ionic compound with the same metal: with $-328, -349, -325, -295$ the ionic character order is ED $<$ EC $<$ EA $<$ EB

ELECTRONEGATIVITY (PAULING SCALE)

- The power of an atom in a bond to attract the shared pair. It is not measurable directly, and not the same as electron gain enthalpy.
- Rises across a period, falls down a group; F 4.0 $>$ O 3.5 $>$ N $\approx$ Cl 3.0 $>$ C 2.5 $\approx$ S; H 2.1; Na 0.9; Cs about 0.7.
- Rises with oxidation state and with the s character of the hybrid orbital ($sp > sp^2 > sp^3$): C in ethyne is more electronegative than in ethane.

SYLLABUS NOTE

Coaching extras (not NCERT): Mulliken $\chi = \dfrac{IE + EA}{2}$ (eV); Pauling $\chi_A - \chi_B = 0.208\sqrt{\Delta}$ with $\Delta = E_{A-B} - \sqrt{E_{A-A}E_{B-B}}$ in kcal mol^{-1}; Allred-Rochow $\chi = 0.744 + 0.359\, Z_{eff}/r^2$ (r in Å).

KEY IDEA

F is the most electronegative element yet Cl has the more negative electron gain enthalpy: electronegativity is the pull inside a bond, electron gain enthalpy the energy of taking up a free electron, and the compact $2p$ shell of F repels it.

QUICK CHECK

Arrange the EGE (most negative first): F, Cl, Br, O, S.

Answer: $\mathrm{Cl} > \mathrm{F} > \mathrm{Br}$ and $\mathrm{S} > \mathrm{O}$: overall $\mathrm{Cl} > \mathrm{F} > \mathrm{Br} > \mathrm{S} > \mathrm{O}$ ($-349, -328, -325, -200, -141$).

5 Valence, anomalous behaviour and the nature of oxides

VALENCE AND OXIDATION STATE

IDEA	RULE OR FACT
valence	groups $1, 2, 13, 14$: the group electrons ($1, 2, 3, 4$); groups 15 to 17 usually $8-$ valence electrons ($3, 2, 1$)
fluorine	only -1 (the most electronegative element); oxygen shows -2, -1, $+2$ (in $\mathrm{OF_2}$) but never $+6$
second period	maximum covalency 4 (no d orbitals): $\mathrm{NCl_5}$, $\mathrm{OF_4}$, $\mathrm{BF_6^{3-}}$ do not exist
inert-pair effect	down groups 13 to 15 the oxidation state two below the group value becomes more stable: $\mathrm{Tl^{+}}$, $\mathrm{Pb^{2+}}$, $\mathrm{Bi^{3+}}$ ($\mathrm{Tl^{3+}}$, $\mathrm{Pb^{4+}}$ are oxidising)
lowest oxidation state	$\mathrm{A_2B}$ with B at -2 means B is in group 16 with 6 valence electrons

WHY THE FIRST ELEMENT OF A GROUP DIFFERS

- small size and high charge/radius ratio; high electronegativity.
- no d orbitals, so covalency stops at 4.
- readily forms $p\pi$-$p\pi$ multiple bonds ($\mathrm{C{=}C}$, $\mathrm{N{\equiv}N}$) which the heavier members do not.

DIAGONAL RELATIONSHIP

- Li and Mg: both burn to give normal oxide and a nitride; carbonates decompose on heating; hydrated chlorides.
- Be and Al: amphoteric oxides and hydroxides; form complexes; chlorides are covalent.
- B and Si: covalent hydrides and acidic oxides; similar acidic nature.

NATURE OF OXIDES

TYPE	EXAMPLES
basic	$\mathrm{Na_2O}$, MgO, CaO, BaO, $\mathrm{K_2O}$
amphoteric	BeO, $\mathrm{Al_2O_3}$, ZnO, SnO, PbO, $\mathrm{Ga_2O_3}$
acidic	$\mathrm{B_2O_3}$, $\mathrm{CO_2}$, $\mathrm{SiO_2}$, $\mathrm{N_2O_5}$, $\mathrm{P_4O_{10}}$, $\mathrm{SO_3}$, $\mathrm{Cl_2O_7}$
neutral	CO, NO, $\mathrm{N_2O}$, $\mathrm{H_2O}$

QUICK FACTS ON THE S-BLOCK AND HYDROGEN (KEPT SHORT; STANDARD PYQ ITEMS)

FACT	EQUATION OR ORDER
Mg burns in air	$2\,\mathrm{Mg} + \mathrm{O_2} \longrightarrow 2\,\mathrm{MgO}$ and $3\,\mathrm{Mg} + \mathrm{N_2} \longrightarrow \mathrm{Mg_3N_2}$
saline hydride with water	$\mathrm{NaH} + \mathrm{H_2O} \longrightarrow \mathrm{NaOH} + \mathrm{H_2}$
diborane	$3\,\mathrm{LiAlH_4} + 4\,\mathrm{BF_3} \longrightarrow 2\,\mathrm{B_2H_6} + 3\,\mathrm{LiF} + 3\,\mathrm{AlF_3}$
hydride classes	electron-deficient $\mathrm{B_2H_6}$; electron-precise $\mathrm{CH_4}$, $\mathrm{SiH_4}$; electron-rich $\mathrm{NH_3}$, $\mathrm{H_2O}$, HF
group 2 carbonates	thermal stability $\mathrm{BeCO_3} < \mathrm{MgCO_3} < \mathrm{CaCO_3} < \mathrm{SrCO_3} < \mathrm{BaCO_3}$ (a large cation polarises $\mathrm{CO_3^{2-}}$ less)
group 1 and 2 hydroxides	basic strength and solubility rise down the group; $\mathrm{Be(OH)_2}$ amphoteric

TRAP

CO, NO and N_2O are neutral, not acidic. Al_2O_3 is amphoteric, not basic. Down a group acidic oxides become less acidic and basic ones more basic. In group 13, Tl (not Pb) shows $+1$ chiefly; do not extend this to carbon.

QUICK CHECK

Which of Li, Na, Be, Mg, B, Al has the smallest radius and its oxide's nature?

Answer: Boron (88 pm); B_2O_3 is acidic.

6 Worked examples

GROUP FROM IE DATA

Successive IE values $800, 2427, 3658, 25024, 32824$ kJ mol^{-1}. Group?

ratios: $3.0,\ 1.5,\ 6.8,\ 1.3$: the jump follows IE_3, so 3 valence electrons, group 13 (B)

IONIC CHARACTER

E has $IE_1 = 374$ kJ mol^{-1}; it reacts with A, B, C, D (EGE $-328, -349, -325, -295$). Order of ionic character of EA, EB, EC, ED?

the more negative the EGE, the more ionic: $ED < EC < EA < EB$

IUPAC NAME

The group number of ununnunium (Z $= 111$)?

Z $= 111$: 11 (it is in group 11, now roentgenium, $[Rn]5f^{14}6d^{10}7s^1$)

STATEMENTS

True or false: (I) IE_1 of $Pb > Sn$; (II) IE_1 of $Ge > Si$.

(I) true ($716 > 708$). (II) false ($762 < 786$, Ge below Si)

CR AND MN

(I) IE_1 of Cr is lower than of Mn. (II) IE_2 and IE_3 of Cr are higher than of Mn.

(I) true ($653 < 717$). (II) false: IE_2 is higher ($1590 > 1509$) but IE_3 is lower ($2987 < 3248$)

ISOELECTRONIC

Which has the largest radius among K^+, Cl^-, Ca^{2+}, S^{2-} (all 18 electrons)?

fewest protons: S^{2-} ($Z = 16$), then $Cl^- > K^+ > Ca^{2+}$

7 Last-minute checklist

BEFORE YOU ANSWER

Same shell or new shell? Z_{eff} up or down? Any filled or half-filled subshell? Is the question about IE, EGE (sign!) or electronegativity?

- Group: s block $= ns$ electrons; p block $= 10 + (ns + np)$; d block $= ns + (n-1)d$. Roots: nil, un, bi, tri, quad, pent, hex, sept, oct, enn.
- Radius: down group up, across period down; Ga < Al; $N^{3-} > O^{2-} > F^- > Na^+ > Mg^{2+} > Al^{3+}$; cation smaller, anion larger than the atom.
- IE: Be > B, Mg > Al, N > O, P > S; B > Tl > Ga > Al > In; C > Si > Ge > Pb > Sn; $IE_1 < IE_2 < IE_3$; jump after IE_k: k valence electrons.
- EGE (kJ mol^{-1}): Cl -349 > F -328 > Br -325 > I -295; S -200 > O -141; Be, Mg, N, noble gases near zero; EN: F 4.0, O 3.5, N and Cl 3.0.
- Second period: covalency ≤ 4, anomalous; diagonal pairs Li-Mg, Be-Al, B-Si; inert pair: Tl^+, Pb^{2+}, Bi^{3+}; oxides: Al_2O_3, ZnO amphoteric, CO, NO neutral.

Now test it. Practise this chapter with past-paper and bank questions on Prep Vault. **Practice this chapter →**