WEIRD CHEMIST

Where Chemistry Gets Interesting

DPP-10 Solution Sheet

p-Block Elements — Group 17: The Halogens

NEET/JEE Revision Series | 50 Questions | Complete Solutions

"Group 17 has ONE star and one anomaly. The STAR is F_2 being the strongest oxidant — explained by 3 factors: low bond dissociation enthalpy, high hydration enthalpy of F^-, and high electrode potential. The ANOMALY is also fluorine — it breaks every group trend. Master the fluorine anomaly and this entire chapter becomes a single connected story."

QUICK ANSWER KEY

Q	Ans	Q	Ans	Q	Ans	Q	Ans	Q	Ans
1	(2)	11	(3)	21	(2)	31	(4)	41	(4)
2	(4)	12	(1)	22	(4)	32	(2)	42	(1)
3	(3)	13	(2)	23	(3)	33	(3)	43	(4)
4	(1)	14	(3)	24	(4)	34	(1)	44	(3)
5	(2)	15	(2)	25	(2)	35	(2)	45	(1)
6	(2)	16	(3)	26	(1)	36	(2)	46	(3)
7	(4)	17	(1)	27	(4)	37	(4)	47	(4)
8	(3)	18	(2)	28	(4)	38	(3)	48	(2)
9	(1)	19	(4)	29	(1)	39	(1)	49	(3)
10	(4)	20	(2)	30	(3)	40	(3)	50	(2)

TOPIC 1 — Introduction

(a) Elements, Nature & Occurrence

SKILL: Analytical Ability

Q.1 Why do halogens show far greater similarity amongst themselves than elements of most other groups?

[A] Conceptual Approach

Sab halogens ka ek hi fundamental drive hai: apni ns^2np^5 configuration ko complete karke noble gas banna (ns^2np^6). Sirf ek electron chahiye. Woh electron ka "need" puri chemistry control karta hai. Baaki differences sirf size aur energy ke hain — direction same hai.

[S] Step-by-Step Solution

All halogens share the **same valence configuration: ns^2np^5** — just one electron short of the next noble gas.

This means every halogen has the same fundamental chemical drive: gain one electron to complete the octet.

Consequences:

- All are strong oxidising agents (electron gainers)
- All form -1 as their most stable OS
- All react similarly with metals, hydrogen, and each other
- Differences are only in *degree* (how strongly they oxidise), not in *kind*

Options (1) and (4) are wrong because not all halogens are gases (Br_2 = liquid, I_2 = solid). Option (3) is historically false.

[Ans] Final Answer

Option (2): All share ns^2np^5 — one electron short of noble gas; same drive, only size/energy differs $\Rightarrow$ strong regular gradation

[!] Common Mistake

Option (1) says "all halogens are gases." Bromine is a liquid and iodine is a solid at room temperature. Physical state is not the reason for their similarity. The reason is always the *electronic configuration*: same drive, same reactions, just different magnitudes.

SKILL: Conceptual Clarity

Q.2 "Halogen" — correct meaning and connection to chemistry?

[A] Conceptual Approach

Greek etymology: "halo" = salt, "genes" = born/produced. So halogen = salt producer. Perfect fit — all halogens react with metals to form ionic salts: NaCl, KBr, AgI etc. NaCl (table salt) is the most famous

example.

[S] Step-by-Step Solution

Etymology: Greek *halo* (salt) + *genes* (born/produced).
∴ Halogen = "salt producer."

Why this name fits perfectly:
All halogens readily combine with metals to form salts:

- $F_2 + 2Na \rightarrow 2NaF$ (sodium fluoride)
- $Cl_2 + 2Na \rightarrow 2NaCl$ (common table salt!)
- $Br_2 + 2K \rightarrow 2KBr$ (potassium bromide)

Option (1) "fire maker" = incorrect; option (2) "colour bearer" = not the etymology; option (3) "salt eater" = wrong direction.

[Ans] Final Answer

Option (4): "Salt producer" — from Greek *halo* (salt) + *genes* (born); halogens react with metals to form salts

[!] Common Mistake

Option (2) says "colour bearer" because halogens are coloured. While halogens ARE all coloured in their elemental state, that is NOT the etymology of the name. "Chlorine" itself comes from Greek "chloros" (greenish-yellow) — but "halogen" specifically means salt producer.

SKILL: Application Skills

Q.3 Sample A = fluorspar (CaF_2); B = carnallite ($KCl{\cdot}MgCl_2{\cdot}6H_2O$); C = dried seaweed (0.5% iodine). Match each to its halogen.

[A] Conceptual Approach

Direct formula reading: CaF_2 contains fluorine (F). $KCl{\cdot}MgCl_2$ contains chlorine (Cl). Seaweed contains iodine (I) — simple and direct.

[S] Step-by-Step Solution

Sample A: Fluorspar = CaF_2 (calcium **fluoride**) $\Rightarrow$ **Fluorine**.

Sample B: Carnallite = $KCl{\cdot}MgCl_2{\cdot}6H_2O$ (potassium **chloride** + magnesium **chloride**) $\Rightarrow$ **Chlorine**.

Sample C: Dried seaweed (contains up to 0.5% iodine) $\Rightarrow$ **Iodine**.

Historically: I_2 was first discovered from seaweed by Courtois (1811). KIO_3 (potassium iodate) from Chile saltpetre is another source.

[Ans] Final Answer

Option (3): A $\rightarrow$ F; B $\rightarrow$ Cl; C $\rightarrow$ I

[!] Common Mistake

Option (1) reverses A and B (CaF_2 = Cl, carnallite = F). Always read the chemical formula directly: CaF_2 has "F" in the name (fluorite/fluorspar), and KCl has "Cl" for chloride. Don't guess — read the formula.

SKILL: Diagram-Based Visual Literacy

Q.4 Table of natural sources for each halogen. Which statement is correctly supported?

[A] Conceptual Approach

Table pe dhyan do. F = mineral ores (CaF_2, Na_3AlF_6, fluoroapatite). Cl = sea water (NaCl 2.5%) + carnallite deposits. I = seaweeds + Chile saltpetre. Br = sea water (bromides). All three parts of option (1) match the table directly.

[S] Step-by-Step Solution

From the table, checking each option:

Option (1): "F occurs mainly in mineral ores" (CaF_2, Na_3AlF_6, fluoroapatite); "Cl found in sea water and dry deposits" ; "I's chief sources are seaweeds and Chile saltpetre" . **All correct.**

Option (2): "F most abundant in sea water" (F is in mineral ores, not primarily sea water; Cl dominates sea water at 2.5%).
Option (3): "I comparable to Cl in sea water" (I is trace, Cl is 2.5%).
Option (4): "Cl and Br in seaweeds" (Cl and Br are in sea water; seaweeds are I sources).

[Ans] Final Answer

Option (1): F in mineral ores; Cl in sea water and dry deposits; I in seaweeds and Chile saltpetre

[!] Common Mistake

Option (4) says Cl and Br are mainly in seaweeds. Wrong! Seaweeds are specifically the source of **iodine** (historically). Chlorine and bromine come from sea water (as dissolved NaCl and bromide salts), not seaweeds.

SKILL: Time Management and Strategic Thinking

Q.5 Statements: I. At = radioactive. II. Halogens show LESS similarity. III. F and Cl abundant; Br and I smaller quantities. IV. Sea water = NaCl at 2.5% by mass. True ones?

[A] Conceptual Approach

I = TRUE. II = FALSE (halogens show MORE similarity, not less). III = TRUE. IV = TRUE. So I, III, IV are true.

[S] Step-by-Step Solution

Statement I: TRUE — Astatine (At) is radioactive; the most stable isotope At-210 has a half-life of only 8.1 hours.

Statement II: **FALSE** — Halogens show *greater* similarity than elements of most other groups, NOT less. The question explicitly says this in the opening.

Statement III: TRUE — F (crustal: 0.065%) and Cl (sea water: 2.5%) are fairly abundant; Br and I are

present in smaller quantities.

Statement IV: TRUE — Sea water contains approximately 2.5% NaCl by mass.

True: I, III and IV.

[Ans] Final Answer

Option (2): I, III and IV only

[!] Common Mistake

Statement II is the trap. Many students read it carelessly and accept it. The passage says halogens show *more* similarity — Statement II says *less*. Classic reversal trap. Always read "more" vs "less" very carefully in statement questions.

TOPIC 2 — Atomic Properties

(a–d) Electronic Config, Atomic Radii, Ionisation Enthalpy & Electron Gain Enthalpy

SKILL: Analytical Ability

Q.6 F's electron gain enthalpy (-333 kJ/mol) is LESS negative than Cl's (-349 kJ/mol) despite F being higher. Why?

[A] Conceptual Approach

Yeh wahi anomaly hai jo Group 15 mein N–N single bond weak tha aur Group 16 mein O ka EGE less negative tha. F ki 2p orbitals bahut compact hain. Already 7 electrons hain in 2s+2p. 8th electron add karo $\Rightarrow$ bahut crowded $\Rightarrow$ strong interelectronic repulsion $\Rightarrow$ less energy released.

[S] Step-by-Step Solution

Expected: EGE should become *more* negative going UP a group (smaller atom, better nuclear attraction for incoming electron).

Anomaly: F (-333) is *less negative* than Cl (-349), despite F being higher in the group.

Explanation: F's 2p orbitals are extremely **small and compact**. The orbital already contains 5 electrons (ns^2np^5). When an 8th electron tries to enter this congested space:

- The existing electron cloud is very dense and localized.
- Strong **interelectronic repulsion** in the compact 2p subshell reduces the net energy released.
- Cl has larger, more diffuse 3p orbitals $\Rightarrow$ less repulsion $\Rightarrow$ more energy released $\Rightarrow$ more negative EGE.

[Ans] Final Answer

Option (2): F's compact 2p orbitals are extremely electron-dense; incoming electron faces strong interelectronic repulsion $\Rightarrow$ less energy released $\Rightarrow$ EGE less negative than Cl

[!] Common Mistake

Option (3) says "the general rule states EGE always becomes less negative going UP." That would be the WRONG rule. EGE generally becomes more negative going up a group (smaller atom, better nuclear-electron attraction). Fluorine is the EXCEPTION, not the rule — and it's an exception for the specific reason of orbital congestion.

SKILL: Conceptual Clarity

Q.7 Why do halogens have the most negative EGE among all elements in their respective periods?

[A] Conceptual Approach

ns^2np^5 = just one electron short of noble gas config. Ek electron gain karo ⇒ full octet = maximum stability. This "drive" to gain one electron is the strongest of any element in their period. Group 18 is already complete — it has no drive. So Group 17 has the maximum EGE.

[S] Step-by-Step Solution

Configuration: Halogens = ns^2np^5 — one electron short of the noble gas configuration (ns^2np^6).

Why this gives maximum EGE:

- Gaining one electron completes the octet and achieves the highly stable noble gas configuration.
- The energy released is exceptionally large because the stability gain is maximal.
- No other element in the same period has such a strong drive to gain exactly one electron.

Group 18 (noble gases) already have complete octets — they have zero drive to gain an electron, so their EGE is essentially zero or slightly positive.

[Ans] Final Answer

Option (4): ns^2np^5 is one electron short of noble gas configuration — exceptionally strong drive to gain one electron and complete the octet

[!] Common Mistake

Option (3) says "halogens already have completely filled orbitals." Wrong! Halogens have 7 valence electrons, not 8. The p-subshell has only 5 electrons (np^5) — it needs one more to be filled (np^6). It is precisely because the orbital is NOT yet complete that halogens have such a strong drive to gain an electron.

SKILL: Application Skills

Q.8 IE values: F=1680, Cl=1256, Br=1142, I=1008 kJ/mol. Between which consecutive pair is the LARGEST drop?

[A] Conceptual Approach

Calculate each drop: F→Cl: 1680−1256=424. Cl→Br: 1256−1142=114. Br→I: 1142−1008=134. Largest = 424 (F to Cl).

[S] Step-by-Step Solution

Pair	Drop	Calculation
F → Cl	**424**	1680 − 1256
Cl → Br	114	1256 − 1142
Br → I	134	1142 − 1008

Largest drop = F to Cl (424 kJ mol^{-1})

Reason: F is a Period 2 element (no d-electrons, very compact). Cl is Period 3 (10 more electrons, including the 2p block). The jump in size and shielding from Period 2 to Period 3 is the largest change in the group, explaining the large IE drop.

[Ans] Final Answer

Option (3): F to Cl (drop = 424 kJ mol^{-1}) — by far the largest single drop

[!] Common Mistake

Guessing Br to I because I is the largest atom. The drop Br→I is only 134 kJ/mol. The biggest jump is always from Period 2 to Period 3 in any group (same pattern in Groups 13, 14, 15, 16 — N→P, O→S, F→Cl all show the biggest drop). Always calculate, don't guess.

SKILL: Diagram-Based Visual Literacy

Q.9 Table: Covalent radius (64, 99, 114, 133 pm) and ionic radius X^- (133, 184, 196, 220 pm) for F, Cl, Br, I. Which statement is correctly supported?

[A] Conceptual Approach

Check: For every halogen, ionic radius > covalent radius? F: 133 > 64 ; Cl: 184 > 99 ; Br: 196 > 114 ; I: 220 > 133 . Both increase from F to I. Option (1) checks out perfectly.

[S] Step-by-Step Solution

From the table, both sets of values increase from F to I:

Covalent: 64 < 99 < 114 < 133 pm (increases down group).
Ionic (X^-): 133 < 184 < 196 < 220 pm (increases down group).

For every halogen: ionic radius > covalent radius. This makes sense:
- When an electron is added ($X \rightarrow X^-$): extra electron-electron repulsion causes the electron cloud to expand.
- Nuclear charge stays the same but electrons increase by 1 ⇒ each electron is held less tightly ⇒ larger radius.

[Ans] Final Answer

Option (1): For every halogen, ionic radius (X^-) is substantially larger than covalent radius; both sets of radii increase regularly from F to I

[!] Common Mistake

Option (2) says "ionic radius is smaller than covalent radius since gaining an electron compresses the cloud." Gaining an electron EXPANDS the cloud (more repulsion between electrons, same nuclear attraction). The ionic radius is ALWAYS larger than the covalent radius for anions.

SKILL: Time Management and Strategic Thinking

Q.10 Statements: I. Halogens smallest radii in their period (max Z_{eff}). II. F IE=1680 > Cl IE=1256. III. Cl EGE more negative than F (F compact 2p). IV. EN of halogens INCREASES F to I. True ones?

[A] Conceptual Approach

I = TRUE. II = TRUE. III = TRUE (the F anomaly we just discussed). IV = FALSE (EN DECREASES going down a group — F = 4.0, Cl = 3.0, Br = 2.8, I = 2.5). So I, II, III are true.

[S] Step-by-Step Solution

Statement I: TRUE — In any period, halogens (Group 17) have the highest nuclear charge without a filled subshell ⇒ maximum Z_{eff} ⇒ smallest atomic radius.

Statement II: TRUE — IE: F (1680) > Cl (1256) kJ mol^{-1}; decreases down group.

Statement III: TRUE — Cl EGE (-349 kJ mol^{-1}) is more negative than F's (-333 kJ mol^{-1}) due to interelectronic repulsion in F's compact 2p orbitals.

Statement IV: **FALSE** — Electronegativity *decreases* down any group. F = 4.0 (highest of all elements); I = 2.5 (lowest in Group 17). EN does NOT increase F to I.

True: I, II and III.

[Ans] Final Answer

Option (4): I, II and III only

[!] Common Mistake

Statement IV is a classic reversal trap. Electronegativity ALWAYS decreases going down a group (atoms get larger, bonding electrons are farther from the nucleus, less tightly held). F = 4.0 is the HIGHEST EN of all 118 elements. Saying EN increases from F to I is like saying a small person can reach higher shelves than a tall person.

TOPIC 3 — Physical Properties

(a–f) Physical State, Boiling/Melting Point, Colour, Solubility & Bond Dissociation Enthalpy

SKILL: Analytical Ability

Q.11 F_2 has SMALLER bond dissociation enthalpy than Cl_2 (expected order Cl>Br>I holds, but F is anomalously low). Correct explanation?

[A] Conceptual Approach

Same root cause as N–N single bond being weak (Group 15) and O's EGE anomaly (Group 16). F atoms are tiny. In F_2, two F atoms are VERY close. Three lone pairs on each F are extremely close to the lone pairs on the other F. **Lone pair–lone pair repulsion** weakens the F–F bond.

[S] Step-by-Step Solution

Expected order: Since F is smallest, F–F should be strongest bond. But actual BDE: F_2 = 158, Cl_2 = 242 kJ mol^{-1}.

Reason: In F_2, two very small F atoms are extremely close. Each F has **three lone pairs**.

- The lone pairs on one F atom and the lone pairs on the other F atom are in very close proximity.
- This causes strong **lone pair–lone pair electron repulsion** across the bond.
- This repulsion weakens the F–F bond, giving it a smaller BDE than Cl–Cl.

From Cl onwards, atoms are larger, lone pairs are farther apart ⇒ less repulsion ⇒ expected trend holds: Cl > Br > I.

[Ans] Final Answer

Option (3): F_2 lone pairs are extremely close (small atom); strong lone pair–lone pair repulsion weakens the F–F bond ⇒ smaller BDE than Cl_2

[!] Common Mistake

Option (2) says "fluorine's high EN means it always forms weaker bonds with itself." This is a vague non-explanation. EN describes how strongly an atom attracts electrons in a bond with ANOTHER element — it doesn't directly explain why the F–F homoatomic bond is weak. The specific reason is lone pair repulsion at close range.

SKILL: Conceptual Clarity

Q.12 All halogens are coloured. Why, and which are the correct colours?

[A] Conceptual Approach

Halogens absorb visible light, exciting electrons to higher energy levels. The colour WE SEE is the complementary colour to what's absorbed. F_2 = pale yellow; Cl_2 = greenish-yellow; Br_2 = reddish-brown/orange; I_2 = purple/violet solid, violet vapour.

[S] Step-by-Step Solution

Cause of colour: Halogens absorb radiation in the **visible region of the electromagnetic spectrum**. This excites outer electrons to higher energy levels. The absorbed wavelength is "missing" from the transmitted/reflected light, so we see the complementary colour.

Actual colours:

- F_2: **pale yellow/yellow** gas
- Cl_2: **greenish-yellow** gas
- Br_2: **reddish-brown/orange** liquid
- I_2: **violet** vapour; greyish-black solid (shiny)

Option (2) says "emit radiation" — wrong. Halogens ABSORB radiation; the colour comes from absorption, not emission.

[Ans] Final Answer

Option (1): Halogens absorb visible radiation, exciting electrons; F_2 = yellow, Cl_2 = greenish-yellow

[!] Common Mistake

Remembering iodine's colour incorrectly. I_2 solid = greyish-black/lustrous purple; I_2 vapour = **violet**; I_2 in organic solvent (CCl_4) = violet. In water, I_2 gives brown colour (forms I_3^- complex). Don't confuse vapour colour with solution colour.

SKILL: Application Skills

Q.13 F_2, Cl_2, Br_2, I_2 added to water and to CCl_4. Correct observations?

[A] Conceptual Approach

F_2 and Cl_2 react with water (not just dissolve). Br_2 and I_2 are sparingly soluble in water but dissolve well in non-polar organic solvents like CCl_4. In CCl_4: Br_2 gives orange, I_2 gives violet.

[S] Step-by-Step Solution

With water:
F_2: **reacts vigorously** (oxidises water to O_2): $2F_2 + 2H_2O \rightarrow O_2 + 4HF$.
Cl_2: **reacts** to form HOCl + HCl (disproportionation).
Br_2: sparingly soluble; very slight reaction to form HOBr + HBr.
I_2: barely soluble; reaction with water non-spontaneous.

With CCl_4 (non-polar organic solvent):
Br_2: dissolves → orange/brown solution.
I_2: dissolves → violet solution (like I_2 vapour, same colour in non-polar solvent).
F_2/Cl_2: react/are too volatile to simply "dissolve" in CCl_4 the same way.

[Ans] Final Answer

Option (2) [correct is (4)]: F_2 and Cl_2 react with water; Br_2 and I_2 sparingly soluble in water but dissolve in CCl_4 giving coloured solutions

[!] Common Mistake

Option (1) says "all four dissolve freely in water." F_2 doesn't just dissolve — it *reacts* with water so vigorously it oxidises it to O_2. This is a key distinction: F_2 is such a powerful oxidant it attacks the solvent itself. Option (3) wrongly says "Br_2 and I_2 react vigorously with water."

SKILL: Diagram-Based Visual Literacy

Q.14 Table: Physical states and bp (F_2=gas/85K, Cl_2=gas/239K, Br_2=liquid/332K, I_2=solid/457K). Which statement is correctly supported?

[A] Conceptual Approach

State changes: gas→gas→liquid→solid (clear trend: more solid going down). Boiling points: 85<239<332<457 K (monotonically increasing). Both trends clearly go in one direction. Option (2) captures this perfectly.

[S] Step-by-Step Solution

Physical state trend: gas (F_2) → gas (Cl_2) → liquid (Br_2) → solid (I_2).
As MW increases, London dispersion forces strengthen ⇒ higher bp ⇒ changes state from gas to liquid to solid.

Boiling point trend: $85 < 239 < 332 < 457$ K — **steadily increases**.

Both trends are consistent with increasing atomic number and stronger van der Waals / London dispersion forces.

Option (1): wrong — bp INCREASES, not decreases. Option (3): wrong — Br_2 is liquid, I_2 is solid (reversed). Option (4): not all are gases.

[Ans] Final Answer

Option (3) [correct is (2)]: Physical state gas → liquid → solid; boiling points steadily increase with atomic number

[!] Common Mistake

Option (3) says "Br_2 is solid, I_2 is liquid." Exactly reversed! Bromine (Br_2) is the only liquid non-metal at room temperature; iodine (I_2) is a shiny solid. This is a famous fact — bromine is one of only two liquid elements at room temperature (the other being mercury).

SKILL: Time Management and Strategic Thinking

Q.15 Statements: I. mp and bp increase with atomic number. II. F_2 BDE larger than Cl_2 (F is lightest). III. Br and I sparingly soluble in water, dissolve in organic solvents. IV. Colour from visible radiation absorption. True ones?

[A] Conceptual Approach

I = TRUE. II = FALSE (F_2 BDE SMALLER than Cl_2 — the anomaly we just covered). III = TRUE. IV = TRUE. So I, III, IV are true.

[S] Step-by-Step Solution

Statement I: TRUE — mp and bp increase steadily from F_2 to I_2 (stronger van der Waals forces).

Statement II: **FALSE** — F_2 BDE (158 kJ/mol) is SMALLER than Cl_2 (242 kJ/mol), not larger. This is the lone pair repulsion anomaly. F being the lightest does NOT make its BDE larger.

Statement III: TRUE — Br_2 and I_2 are sparingly soluble in water but dissolve well in CCl_4, chloroform, and other organic solvents.

Statement IV: TRUE — Halogens absorb visible light, exciting electrons. The absorbed colour is removed, and we see the complementary colour.

True: I, III and IV.

[Ans] Final Answer

Option (2) [correct is (3)]: I, III and IV only

[!] Common Mistake

Statement II: "F_2 has larger BDE because fluorine is the lightest" uses faulty reasoning. Lighter = smaller, and smaller atoms should overlap better ⇒ stronger bond. But this ignores the lone pair repulsion! The actual result is the OPPOSITE of intuition — F_2 BDE is SMALLER. Always remember the F_2 BDE anomaly.

TOPIC 4 — Chemical Properties

(a) Oxidation States, Oxidising Behaviour & Reaction with Water

SKILL: Analytical Ability

Q.16 F_2 oxidises Cl^-, Br^-, I^-; Cl_2 oxidises Br^-, I^- only; Br_2 oxidises I^- only. Underlying principle?

[A] Conceptual Approach

Think of it like a school hierarchy: senior can bully juniors but not seniors. F_2 is the strongest oxidant — can displace all weaker halogens. Cl_2 is weaker than F_2 but stronger than Br_2 and I_2 — can displace those two. Br_2 can only displace I_2 (which is even weaker). The key: stronger oxidising halogen displaces halide of weaker oxidising halogens (higher in group = stronger oxidant).

[S] Step-by-Step Solution

Oxidising power order: $F_2 > Cl_2 > Br_2 > I_2$ (decreasing down the group).

Principle: A halogen can oxidise (displace) the halide ions of any halogen below it in the group (weaker oxidants), but cannot displace halide ions of halogens above it.

Examples:
- Cl_2 + 2KBr → 2KCl + Br_2 (Cl displaces Br: Cl > Br in oxidising power)
- Cl_2 + 2KF → No reaction (Cl cannot displace F: F > Cl in oxidising power)

This is a **redox competition**: the stronger electron-acceptor wins.

[Ans] Final Answer

Option (3): A stronger oxidising halogen displaces halide ions of weaker oxidising halogens (those below it in the group)

[!] Common Mistake

Option (2) says "a halogen can only oxidise the halide ion immediately above it." Too restrictive! F_2 can displace ALL three halides (Cl^-, Br^-, I^-) — not just Cl^- which is immediately above. It can skip multiple steps. The rule is any halogen BELOW in the group (weaker oxidant), not just the immediate neighbour.

SKILL: Conceptual Clarity

Q.17 F only shows −1 OS. Cl, Br, I also show +1, +3, +5, +7. Why is F restricted to −1?

[A] Conceptual Approach

Two separate reasons, both must be stated: (1) F is most electronegative element — NEVER loses electrons to a more electronegative partner (there's no partner more electronegative than F). (2) F has no d-orbitals (Period 2) — cannot expand its octet to accommodate 3, 5, or 7 bonding pairs.

[S] Step-by-Step Solution

Two reasons, both necessary:

Reason 1 — No positive OS possible:
F has the highest electronegativity (4.0). There is NO element more electronegative than F.
$\therefore$ F NEVER loses electrons to a partner. Positive OS requires being the less electronegative partner in a bond (like Cl in HOCl), but F is always the most electronegative.

Reason 2 — No d-orbitals:
F is in Period 2 (valence shell = 2s + 2p only; no 2d exists).
To show +3, +5, +7, an atom must promote electrons into d-orbitals to form more bonds.
F cannot do this (no d-orbitals available) $\Rightarrow$ max 1 bond pair beyond the 3 lone pairs.

Cl, Br, I (Periods 3, 4, 5): have d-orbitals (3d, 4d, 5d) accessible for excitation.

[Ans] Final Answer

Option (1): F is most electronegative (never forms positive OS) AND has no d-orbitals (cannot expand octet) — both factors together restrict F to -1 only

[!] Common Mistake

Option (2) says "F has too many electrons, cannot remove any." F has only 7 valence electrons — not "too many." And ALL elements can theoretically lose electrons under extreme conditions (IEs can be overcome). The issue is not electron count but the combination of highest EN and absence of d-orbitals.

SKILL: Application Skills

Q.18 X oxidises water to O_2 (very strong oxidant); Y forms HY + HOY with water (moderate); Z's reaction with water is non-spontaneous (Z^- is oxidised by O_2). Identify X, Y, Z.

[A] Conceptual Approach

Who oxidises water? Only F_2 is strong enough to oxidise $H_2O \rightarrow O_2$. So X = F_2. Who forms HOCl + HCl (disproportionation)? Y = Cl_2 (or Br_2, but typically Cl_2 is the textbook example). Z's reaction is non-spontaneous AND Z^- gets oxidised by dissolved O_2 in acid — this is I^- being oxidised to I_2 by atmospheric O_2. So Z = I_2.

[S] Step-by-Step Solution

Identifying X: Oxidises water to O_2 $\Rightarrow$ extremely strong oxidant $\Rightarrow$ only $\mathbf{F_2}$ does this: $2F_2 + 2H_2O \rightarrow O_2 + 4HF$.

Identifying Y: Forms HY + HOY $\Rightarrow$ disproportionation in water $\Rightarrow$ $\mathbf{Cl_2}$: $Cl_2 + H_2O \rightarrow HCl + HOCl$. ($Br_2$ also does this but less readily.)

Identifying Z: Reaction with water is non-spontaneous; Z^- can be oxidised by dissolved O_2 in acidic solution $\Rightarrow$ this is $\mathbf{I_2}$: $I^- + O_2$ (in acid) $\rightarrow I_2 + H_2O$. I_2 is such a weak oxidant that even atmospheric O_2 can reverse the reaction.

$\therefore$ X = F_2, Y = Cl_2 (or Br_2), Z = I_2.

[Ans] Final Answer

Option (4): X = F_2; Y = Cl_2 or Br_2; Z = I_2

[!] Common Mistake

Option (1) places X = Cl_2. Chlorine does NOT oxidise water to O_2 — it disproportionates in water ($Cl_2 + H_2O \rightarrow HCl + HOCl$). Only fluorine is powerful enough to oxidise water itself to O_2. This is one of the most unique properties of F_2.

SKILL: Diagram-Based Visual Literacy

Q.19 Table: Orbital excitation state → unpaired electrons → accessible OS. Which statement is correctly supported?

[A] Conceptual Approach

Table shows excited states allow +3, +5, +7 OS by promoting electrons into d-orbitals. Fluorine CANNOT reach excited states (no d-orbitals). So Cl, Br, I can access all four OS; F is stuck at ground state (-1 or $+1$ at most from the table, but actually only -1 as we discussed).

[S] Step-by-Step Solution

From the table: higher excited states (using d-orbitals) give more unpaired electrons ⇒ higher positive OS.

Cl, Br, I: Have 3d, 4d, 5d orbitals in their valence shells ⇒ can be excited to achieve +3, +5, +7 states.

F: Period 2, valence shell = 2s + 2p only. **No d-orbitals** in n=2 shell ⇒ cannot reach first, second, or third excited states in the table ⇒ maximum OS is determined by ground state configuration only (-1, and technically $+1$ in OF_2 due to special circumstance).

Option (2) correctly states: Cl, Br, I can reach +7 via d-orbitals; F cannot because its n=2 shell has no d-orbitals.

[Ans] Final Answer

Option (2) [correct is (2)]: Cl, Br, I can reach +7 via d-orbitals (third excited state); F cannot — its n=2 valence shell has no d-orbitals

[!] Common Mistake

Option (1) says "F can reach the third excited state to show +7 like iodine." Completely wrong! F has NO d-orbitals (Period 2). If you try to excite an F electron to a d-orbital, there is simply no d-orbital available in n=2. This is the same reason F cannot form NF_5-type compounds.

SKILL: Time Management and Strategic Thinking

Q.20 Statements: I. F_2 = strongest oxidising agent. II. Cl_2 can oxidise Br^- and I^-. III. I_2 reacts spontaneously with water to form HI + HOI. IV. Cl, Br can show +4, +6 in some compounds. True ones?

[A] Conceptual Approach

I = TRUE. II = TRUE. III = FALSE (I_2 reaction with water is NON-spontaneous; it's I^- that can get oxidised, not I_2 spontaneously reacting). IV = TRUE (in oxides and oxoacids, Cl and Br can show +4 and +6 OS). So I, II, IV correct.

[S] Step-by-Step Solution

Statement I: TRUE — F_2 is the strongest oxidising agent among halogens (and among all elements).

Statement II: TRUE — $Cl_2 + 2KBr \rightarrow 2KCl + Br_2$; also $Cl_2 + 2KI \rightarrow 2KCl + I_2$. Chlorine displaces both bromide and iodide.

Statement III: **FALSE** — I_2's reaction with water is *non-spontaneous*. In fact, iodide ion (I^-) can be oxidised back to I_2 by dissolved O_2 in acidic medium — the reverse happens, not I_2 reacting with water.

Statement IV: TRUE — Cl and Br exhibit +4 and +6 OS in their oxides (ClO_2, Cl_2O_6, BrO_2) and in combination with oxygen or fluorine.

True: I, II and IV.

[Ans] Final Answer

Option (1) [correct is (1)]: I, II and IV only

[!] Common Mistake

Statement III: "I_2 reacts spontaneously with water." This confuses I_2 with Cl_2. Chlorine *does* react with water (spontaneously). Iodine does NOT — I_2 is such a weak oxidant that the reaction with water is thermodynamically unfavourable. The dissolved I^- in acidic medium can actually be oxidised by atmospheric O_2 back to I_2 — the complete opposite.

(b) Reactivity towards Hydrogen

SKILL: Analytical Ability

Q.21 Acidic strength: HF < HCl < HBr < HI — counterintuitive since F is most electronegative. Correct resolution?

[A] Conceptual Approach

EN of F is highest ⇒ H–F bond is very polar ⇒ you might think HF donates H^+ easily. But the key is not polarity alone — it's the ENERGY COST to break the H–E bond (bond dissociation enthalpy). H–F bond is the strongest among all H–X bonds (574 kJ/mol). Very hard to break ⇒ very few H^+ ions released ⇒ weakest acid.

[S] Step-by-Step Solution

The key factor is H–X bond dissociation enthalpy, not polarity:

H–F BDE = 574 kJ mol^{-1} (strongest H–X bond; very short, compact, hard to break).
H–Cl BDE = 432; H–Br = 363; H–I = 295 kJ mol^{-1} (decreasing).

For an acid to donate H^+: the H–X bond must first break (in solution, assisted by hydration).

H–F bond is so strong that very little of it dissociates in water ⇒ low H^+ concentration ⇒ weakest acid.

As bond enthalpy decreases F → I: bond breaks more easily ⇒ more H^+ released ⇒ acidity increases.

∴ HF is a weak acid (pKa = 3.2); HCl, HBr, HI are all strong acids.

[Ans] Final Answer

Option (4): H–F BDE is highest (574 kJ/mol); very hard to break; little H^+ released ⇒ weakest acid. Decreasing BDE F→I means increasing acidity

[!] Common Mistake

"F is most electronegative ⇒ H–F most polar ⇒ strongest acid." This confuses *polarity* with *bond strength*. Yes, H–F is the most polar bond. But polarity doesn't mean the bond is easy to break — H–F is actually the **strongest** and **shortest** H–X bond. Strength (BDE) determines acid strength, not polarity.

SKILL: Conceptual Clarity

Q.22 Boiling points: HF=293K, HCl=189K, HBr=206K, HI=238K. HF is anomalously high. Why?

[A] Conceptual Approach

Exactly the same anomaly as H_2O in Group 16 and NH_3 in Group 15. HF has strong F–H...F hydrogen bonding because F is highly electronegative (4.0) and very small. The H-bond in HF is so strong that even solid HF and liquid HF exist as zig-zag chains of hydrogen-bonded molecules.

[S] Step-by-Step Solution

Expected bp order (van der Waals only, increasing with MW): HF < HCl < HBr < HI.
Actual bp: HCl (189 K) < HBr (206 K) < HI (238 K) < **HF (293 K)**.

HF anomaly: F is the most electronegative element (4.0) and very small.
⇒ H–F bond is highly polarised.
⇒ HF forms very strong **intermolecular F–H$\cdots$F hydrogen bonds** in the liquid state.
⇒ Much more energy needed to break these H-bonds ⇒ anomalously high bp.

HCl, HBr, HI: Cl/Br/I are larger and less electronegative; H-bonding is absent or very weak.
These three show the expected van der Waals trend (bp increases with MW).

[Ans] Final Answer

Option (2): HF forms strong intermolecular F–H$\cdots$F hydrogen bonds (F = highest EN + very small); much more energy needed to break them ⇒ anomalously high bp

[!] Common Mistake

Option (3) says "shorter H–X bond ⇒ higher bp." H–F is indeed the shortest bond, but bond *length* does not determine boiling point. Boiling point is determined by *intermolecular* forces, not intramolecular bond length. The reason is hydrogen bonding, not bond length.

SKILL: Application Skills

Q.23 Table: H–X bond length (91.7, 127.4, 141.4, 160.9 pm) and BDE (574, 432, 363, 295 kJ/mol) for HF, HCl, HBr, HI. Which statement is correctly supported?

[A] Conceptual Approach

Bond length INCREASES (91.7→160.9 pm); BDE DECREASES (574→295 kJ/mol). Longer bond = weaker bond = lower BDE. This consistent inverse relationship confirms the data.

[S] Step-by-Step Solution

From the table:

Bond length: HF (91.7) < HCl (127.4) < HBr (141.4) < HI (160.9) pm — **increases** down the group (as halogen size increases).

Bond dissociation enthalpy: HF (574) > HCl (432) > HBr (363) > HI (295) kJ mol^{-1} — **decreases** down the group.

Relationship: As bond length increases, BDE decreases.
Longer bonds have weaker overlap between H and X orbitals ⇒ less stable ⇒ easier to break ⇒ lower BDE.

This also explains why HI is the strongest acid (lowest BDE, easiest to break ⇒ most H^+ released).

[Ans] Final Answer

Option (3): H–X bond length increases HF → HI; BDE decreases; longer bonds are weaker — consistent with decreasing stability down the group

[!] Common Mistake

Option (1) says "BDE increases as bond length increases." The data literally shows the exact opposite: as bond length goes from 91.7 to 160.9 pm (increases), BDE goes from 574 to 295 kJ/mol (decreases). Never say longer bonds are stronger — that's true only in very specific resonance situations, not for simple covalent bonds.

SKILL: Diagram-Based Visual Literacy

Q.24 Table: pKa values: HF=3.2, HCl=−7.0, HBr=−9.5, HI=−10.0. More negative pKa = stronger acid. Which statement is correctly supported?

[A] Conceptual Approach

pKa: more negative = stronger acid. HF pKa=3.2 (positive) = weakest. HI pKa=−10.0 = most negative = strongest. Order is HF < HCl < HBr < HI in acidity. This perfectly matches option (1).

[S] Step-by-Step Solution

pKa and acidity: More negative pKa ⇒ stronger acid (larger Ka, more H^+ in solution).

From data:

- HF (pKa = 3.2): **weakest acid** (positive pKa means Ka < 1, weak acid)
- HCl (pKa = −7.0): strong acid
- HBr (pKa = −9.5): stronger acid
- HI (pKa = −10.0): **strongest acid** (most negative pKa)

Acidic strength order: HF < HCl < HBr < HI

This is consistent with decreasing H–X BDE down the group (easier to break bond ⇒ more H^+ released).

[Ans] Final Answer

Option (1): Acidic strength HF $<$ HCl $<$ HBr $<$ HI; HF weakest (pKa=3.2), HI strongest (pKa=-10.0); consistent with decreasing H–X BDE

[!] Common Mistake

Option (2) says "HF is the strongest acid because it has the highest pKa." This reverses the logic. A higher (more positive) pKa means WEAKER acid (smaller Ka). Lower (more negative) pKa means stronger acid. HF with pKa = 3.2 is a **weak acid**; HI with pKa = -10.0 is an extremely **strong acid**.

SKILL: Time Management and Strategic Thinking

Q.25 Statements: I. All halogens react with H_2, affinity decreases F→I. II. Acid strength: HF<HCl<HBr<HI III. Stability of HX INCREASES down group (BDE increases F→I). IV. HX dissolve in water to form hydrohalic acids. True ones?

[A] Conceptual Approach

I = TRUE. II = TRUE. III = FALSE (stability DECREASES down group because BDE decreases). IV = TRUE. So I, II, IV correct.

[S] Step-by-Step Solution

Statement I: TRUE — All halogens react with H_2: F_2 (explosive), Cl_2 (with UV), Br_2 (on heating), I_2 (reversibly at high T). Reactivity with H_2 decreases F → I.

Statement II: TRUE — Acidic strength: HF $<$ HCl $<$ HBr $<$ HI (verified by pKa data).

Statement III: **FALSE** — Thermal stability **decreases** down the group (BDE decreases: 574 → 432 → 363 → 295 kJ/mol). Lower BDE = decomposes more easily = less stable. HF is most stable; HI decomposes on heating.

Statement IV: TRUE — All HX gases dissolve in water: HF → hydrofluoric acid (weak); HCl → hydrochloric acid (strong); HBr → hydrobromic acid; HI → hydroiodic acid.

True: I, II and IV.

[Ans] Final Answer

Option (2) [correct is (2)]: I, II and IV only

[!] Common Mistake

Statement III is a critical reversal. Students who mix up the acidity trend (increases H→I) with the stability trend sometimes write both increase. They are OPPOSITE: acidity increases F→I; thermal stability DECREASES F→I. Both are explained by the same cause (decreasing BDE), just in opposite directions of effect.

(c) Reactivity towards Oxygen

SKILL: Analytical Ability

Q.26 F's compounds with oxygen (OF_2, O_2F_2) are called "oxygen fluorides" NOT "fluorine ox-

ides." Why?

[A] Conceptual Approach

Naming convention: in binary compounds, the MORE electronegative element is named last and gets the negative OS. F is MORE electronegative than O (F=4.0, O=3.5). So F gets the negative OS (-1), and O is forced into a positive OS. Positive O = O is the "special" atom here. So these are derivatives of **oxygen** (in positive state) combined with fluoride.

[S] Step-by-Step Solution

Naming rule: In binary compounds, the more electronegative atom gets the negative oxidation state.

F is more electronegative than O (F = 4.0, O = 3.5).
$\Rightarrow$ F takes -1 oxidation state.
$\Rightarrow$ O is forced into a **positive oxidation state** (+2 in OF_2).

Why "oxygen fluorides": In OF_2, oxygen has a positive OS (+2) — it is behaving like the "metal" or electropositive component. The compound is named from the perspective of oxygen being in a positive/unusual state combined with fluoride ion.

In normal "fluorine oxides," F would be positive and O negative — but that's impossible since F is ALWAYS more electronegative.

[Ans] Final Answer

Option (4): F is more electronegative than O, so F takes -1 and O is forced into a positive OS — these compounds are correctly named as oxygen fluorides (O in positive state)

[!] Common Mistake

Option (2) says "F is less electronegative than O in these compounds." Electronegativity is a fixed property of an element (Pauling scale) — it does NOT change depending on which compound the element is in. F is ALWAYS more electronegative than O. There is no compound in which F "becomes less electronegative."

SKILL: Conceptual Clarity

Q.27 Which halogen oxide matches which industrial/analytical application?

[A] Conceptual Approach

NCERT key facts to remember: ClO_2 = bleaching agent for paper pulp + water treatment. I_2O_5 = powerful oxidising agent used in estimation of CO (carbon monoxide). These two facts are NCERT-specific.

[S] Step-by-Step Solution

ClO_2 (chlorine dioxide):

- Used as a **bleaching agent** for paper pulp and textiles
- Used in **water treatment** (more effective and less harmful than Cl_2 alone)

I_2O_5 (iodine pentoxide):

- A powerful **oxidising agent**
- Used in the **estimation of carbon monoxide** (CO): $I_2O_5 + 5CO \rightarrow I_2 + 5CO_2$. The liberated I_2 is measured to quantify CO.

Option (3) is the only one that correctly states both applications.

[Ans] Final Answer

Option (3): ClO_2 = bleaching agent for paper pulp/textiles + water treatment; I_2O_5 = strong oxidising agent for CO estimation

[!] Common Mistake

Confusing the applications: I_2O_5 is the CO estimator (not ClO_2). A memory trick: "I_2O_5 counts CO" — iodine pentoxide is specifically used for carbon monoxide determination in gas analysis. ClO_2 is used in water treatment and paper bleaching.

SKILL: Application Skills

Q.28 O_2F_2 is a strong fluorinating agent used to oxidise Pu to PuF_6 (spent nuclear fuel). What dual role does O_2F_2 play?

[A] Conceptual Approach

"Fluorinating agent" = attaches F atoms to another element. "Oxidising agent" = raises OS of another element. For Pu $\rightarrow$ PuF_6: (1) Pu gains F atoms (fluorination); (2) Pu OS is raised to +6 (oxidation). Both happen simultaneously. O_2F_2 does both in one reaction.

[S] Step-by-Step Solution

Reaction: Pu + O_2F_2 $\rightarrow$ PuF_6 + O_2 (simplified).

Dual role of O_2F_2:
(1) **Fluorinating agent**: transfers F atoms to Pu, forming the Pu–F bonds in PuF_6.
(2) **Oxidising agent**: oxidises Pu from its ground state to the +6 oxidation state (Pu^{6+} in PuF_6).

Why PuF_6 is useful: It is volatile (a gas), so it can be separated from solid spent fuel components — essential for nuclear fuel reprocessing.

This dual role (fluorinate + oxidise simultaneously) is what makes O_2F_2 uniquely effective for this purpose.

[Ans] Final Answer

Option (4): O_2F_2 acts as both a fluorinating agent (attaches F to Pu) and oxidising agent (raises Pu to +6) to form volatile PuF_6

[!] Common Mistake

Option (1) says O_2F_2 acts as a reducing agent. A fluorinating agent of fluorine's strength is always an **oxidising agent** (F wants electrons, so it oxidises other elements). O_2F_2 cannot be a reducing agent — it is one of the most powerful fluorinating and oxidising agents known.

SKILL: Diagram-Based Visual Literacy

Q.29 Table: F oxides (stable, strong fluorinating); Cl oxides (explosive, ClO_2 for bleaching); Br oxides (least stable, only at low T); I oxides (solids, I_2O_5 for CO estimation). Which statement is correctly supported?

[A] Conceptual Approach

From the table: Br oxides = "least stable" (least of all 4). Cl oxides are reactive/explosive but more stable than Br. I oxides = insoluble solids, more stable than Br. So order of stability: I > Cl > Br. Bromine is the "middle row anomaly" — less stable than neighbours.

[S] Step-by-Step Solution

From the table:

- F oxides: only OF_2 stable at 298 K (limited stability overall).
- Cl oxides: reactive, explosive, but exist at room temperature.
- Br oxides: **least stable of all halogen oxides**; exist only at low temperatures ("middle row anomaly").
- I oxides: stable insoluble solids at room temperature.

Stability order: I > Cl > Br (with F having special considerations).

The anomaly is bromine — it sits between Cl and I but forms the *least stable* oxides of the group (classic second-row anomaly effect). Option (3) correctly captures this.

[Ans] Final Answer

Option (1) [correct is (3)]: Stability order I > Cl > Br; bromine oxides are the least stable (middle row anomaly), existing only at low temperatures

[!] Common Mistake

Expecting stability to simply increase I→Br→Cl→F down the group. Bromine is the anomaly here — it forms *less stable* oxides than both its neighbours (Cl and I). This "middle member anomaly" appears multiple times in p-block chemistry (Se in Group 16 also shows some anomalies).

SKILL: Time Management and Strategic Thinking

Q.30 Statements: I. F oxides called "oxygen fluorides" because F takes negative OS. II. General stability order I > Cl > Br. III. Cl oxides are insoluble solids; ClO_2 used for CO estimation. IV. I_2O_5 powerful oxidising agent for CO estimation. True ones?

[A] Conceptual Approach

I = TRUE. II = TRUE. III = FALSE (Cl oxides are gases/reactive compounds, NOT insoluble solids; AND ClO_2 is used for bleaching/water treatment, NOT CO estimation — that's I_2O_5). IV = TRUE. So I, II, IV correct.

[S] Step-by-Step Solution

Statement I: TRUE — F is more electronegative than O, takes -1, O forced into positive state, so compounds are "oxygen fluorides."

Statement II: TRUE — I_2O_5 = stable solid; Cl_2O, ClO_2 = reactive gases; Br oxides = least stable. So I > Cl > Br.

Statement III: **FALSE** — Two errors: (a) Cl oxides are NOT insoluble solids (Cl_2O, ClO_2, Cl_2O_7 are reactive gases/liquids). (b) ClO_2 is used for *bleaching*, not CO estimation. CO estimation uses $\mathbf{I_2O_5}$.

Statement IV: TRUE — $I_2O_5 + 5CO \rightarrow I_2 + 5CO_2$. The liberated I_2 quantifies CO. Correct application.

True: I, II and IV.

[Ans] Final Answer

Option (1) [correct is (1)]: I, II and IV only

[!] Common Mistake

Statement III contains a doubled trap: wrong physical state (Cl oxides are NOT solids) AND wrong application (ClO_2 bleaches, not estimates CO). The CO estimation one specifically is I_2O_5. These two specific facts (ClO_2=bleaching, I_2O_5=CO estimation) appear repeatedly in NEET/JEE exams.

(d) Reactivity towards Metals — Metal Halides

SKILL: Analytical Ability

Q.31 $SnCl_4$, $PbCl_4$, $SbCl_5$, UF_6 all MORE covalent than $SnCl_2$, $PbCl_2$, $SbCl_3$, UF_4. What links these comparisons?

[A] Conceptual Approach

Same principle as Group 14 ($SnCl_4$ more covalent than $SnCl_2$). Higher OS metal $\Rightarrow$ smaller ionic size + higher charge $\Rightarrow$ higher charge density $\Rightarrow$ polarises anion more $\Rightarrow$ more covalent character (Fajans' rules).

[S] Step-by-Step Solution

Underlying principle: Fajans' Rules

When a metal is in a higher oxidation state (e.g., Sn^{4+} vs Sn^{2+}):
- More electrons removed $\Rightarrow$ metal ion is **smaller**.
- Higher charge (4+ vs 2+).
- Combined: **higher charge density** (charge/volume ratio).

Higher charge density $\Rightarrow$ the metal cation polarises the halide anion more strongly (distorts the electron cloud towards itself).
More polarisation $\Rightarrow$ more **covalent character** (Fajans' rules).

This applies to ALL the pairs: $SnCl_4 > SnCl_2$; $PbCl_4 > PbCl_2$; $SbCl_5 > SbCl_3$; $UF_6 > UF_4$.

[Ans] Final Answer

Option (4) [correct is (2)]: Higher OS metal = smaller + higher charge = greater charge density $\Rightarrow$ stronger polarisation of halide $\Rightarrow$ more covalent (Fajans' rules)

[!] Common Mistake

Option (1) says "more atoms = more covalent." This is a made-up rule with no chemical basis. Covalency is about bonding character (electron sharing), not the number of atoms. $SnCl_4$ is more covalent than $SnCl_2$ because of the higher charge density of Sn^{4+}, not because it has more Cl atoms.

SKILL: Conceptual Clarity

Q.32 Ionic character of MX decreases: MF > MCl > MBr > MI. Why are iodides most covalent and fluorides most ionic?

[A] Conceptual Approach

Fajans' rules for the halide side: I^- is the LARGEST halide (most polarisable). Large, diffuse electron cloud is easily distorted (polarised) by the metal cation. More distortion = more covalent. F^- is the SMALLEST halide (least polarisable) = hardest to distort = stays ionic.

[S] Step-by-Step Solution

Fajans' Rules applied to halide ions:

For a given metal cation (M^+), the covalent character depends on how easily the halide ion is polarised (electron cloud distorted).

F^-: Smallest halide (ionic radius = 133 pm). Compact, tightly held electron cloud. **Least polarisable** $\Rightarrow$ resists distortion $\Rightarrow$ M–F bond remains predominantly **ionic**.

I^-: Largest halide (ionic radius = 220 pm). Diffuse, loosely held electron cloud. **Most polarisable** $\Rightarrow$ easily distorted by M^+ $\Rightarrow$ increased electron sharing $\Rightarrow$ M–I bond is most **covalent**.

Order: ionic character MF $>$ MCl $>$ MBr $>$ MI (decreasing).

[Ans] Final Answer

Option (1): I^- is largest and most polarisable; strong distortion by metal cation $\Rightarrow$ most covalent. F^- is smallest and least polarisable $\Rightarrow$ most ionic

[!] Common Mistake

Option (2) says "F is most electronegative, so all F compounds are most covalent." The electronegativity of F actually makes F PULL electrons towards itself strongly, which if anything increases the polarity of M–F bonds (more ionic character, not less). High EN $\Rightarrow$ more polar bond $\Rightarrow$ more ionic, not more covalent.

SKILL: Application Skills

Q.33 Which pair correctly identifies the MORE covalent compound?

[A] Conceptual Approach

Two Fajans' rules to apply: (a) higher OS = more covalent; (b) larger halide (I>Br>Cl>F) = more covalent. Option (4): $PbCl_4$ (+4) vs $PbCl_2$ (+2). Higher OS = more covalent. $PbCl_4$ wins.

[S] Step-by-Step Solution

Check each option using Fajans' rules:

(1) "UF_4 more covalent than UF_6": WRONG. U^{6+} has higher charge density than U^{4+} $\Rightarrow$ UF_6 is more covalent.

(2) "MgF_2 more covalent than MgI_2": WRONG. I^- is more polarisable than F^- $\Rightarrow$ MgI_2 is more covalent.

(3) "$PbCl_2$ more covalent than $PbCl_4$": WRONG. Pb^{4+} has higher charge density than Pb^{2+} $\Rightarrow$ $PbCl_4$ is more covalent.

(4) "**$PbCl_4$ more covalent than $PbCl_2$**": **CORRECT**. Pb^{4+} (higher charge, smaller size) has greater charge density than Pb^{2+} $\Rightarrow$ stronger polarisation of Cl^- $\Rightarrow$ more covalent character.

[Ans] Final Answer

Option (4): $PbCl_4$ is more covalent than $PbCl_2$ — Pb^{4+} has greater charge density than Pb^{2+}, polarises Cl^- more strongly

[!] Common Mistake

Option (2) says "MgF_2 more covalent because F has highest EN." Electronegativity of the halide is not what determines covalency in this context — it is the **polarisability** of the halide ion. I^- (large, diffuse) is much more polarisable than F^- (small, compact). MgI_2 is MORE covalent than MgF_2.

SKILL: Diagram-Based Visual Literacy

Q.34 Table: $SnCl_2/SnCl_4$, $PbCl_2/PbCl_4$, $SbCl_3/SbCl_5$ — all show higher OS halide is more covalent. Which statement is supported?

[A] Conceptual Approach

All three pairs show same trend: higher OS = more covalent. Option (2) states this generalisation correctly: in all three pairs, the higher OS compound is more covalent because higher charge density polarises the anion more.

[S] Step-by-Step Solution

From the table: in all three pairs ($SnCl_2/SnCl_4$, $PbCl_2/PbCl_4$, $SbCl_3/SbCl_5$), the **higher oxidation state halide is more covalent**.

Why: In each case, the metal in the higher OS is:
- Smaller (more electrons removed $\Rightarrow$ more contracted ionic radius)
- More highly charged
- $\Rightarrow$ Higher charge density $\Rightarrow$ stronger polarisation of Cl^- $\Rightarrow$ more covalent character

This is a **universal principle** (Fajans' rules), applicable to any metal in two different OS with the same halide.

[Ans] Final Answer

Option (2): In all three pairs, the higher OS compound is more covalent because higher metal charge density polarises the anion more strongly

[!] Common Mistake

Option (1) says "higher OS halides are more ionic." This is exactly backwards! Higher OS $\Rightarrow$ more charge density $\Rightarrow$ more covalent. The data in the table explicitly lists $SnCl_4$, $PbCl_4$, $SbCl_5$ (higher OS) as the "more covalent" ones — directly contradicting option (1).

SKILL: Time Management and Strategic Thinking

Q.35 Statements: I. Halogens react with metals to form metal halides. II. Ionic character of MX: MF>MCl>MBr>MI. III. Higher OS halides more covalent than lower OS. IV. $SnCl_2$ more covalent than $SnCl_4$. True ones?

[A] Conceptual Approach

I = TRUE. II = TRUE. III = TRUE. IV = FALSE (exact opposite — $SnCl_4$ more covalent than $SnCl_2$). So I, II, III correct.

[S] Step-by-Step Solution

Statement I: TRUE — Standard reaction: $2Na + Cl_2 \rightarrow 2NaCl$; $Fe + Br_2 \rightarrow FeBr_2$, etc.

Statement II: TRUE — Ionic character decreases as halide polarisability increases: F^- (least polarisable) $\rightarrow I^-$ (most polarisable).

Statement III: TRUE — Higher OS metal has greater charge density $\Rightarrow$ more covalent (Fajans' rules, confirmed by the table).

Statement IV: **FALSE** — $SnCl_4$ is MORE covalent than $SnCl_2$ (NOT the other way). Sn^{4+} has a higher charge density than Sn^{2+}, making $SnCl_4$ the more covalent compound.

True: I, II and III.

[Ans] Final Answer

Option (3) [correct is (3)]: I, II and III only

[!] Common Mistake

Statement IV flips the covalency comparison. The rule is clear: $SnCl_4$ (higher OS, Sn^{4+}) is MORE covalent; $SnCl_2$ (lower OS, Sn^{2+}) is MORE ionic. If you know the Fajans' rule direction (higher OS = more covalent), you immediately identify Statement IV as false.

(e) Interhalogens & NCERT In-Text Questions

SKILL: Analytical Ability

Q.36 Interhalogens: in IF_7, iodine is central (not fluorine). Why is the LARGER halogen always the central atom?

[A] Conceptual Approach

Central atom in any compound needs to accommodate multiple bonds. More bonds $\Rightarrow$ needs d-orbitals for expansion beyond 4. Larger halogen (like I) has accessible d-orbitals (Period 5, 5d available). Smaller halogen (F, Period 2) has NO d-orbitals. So only the larger halogen can be at the centre for XX'_3, XX'_5, XX'_7 types.

[S] Step-by-Step Solution

Why X (larger halogen) is always central:

In interhalogen compounds XX'_3, XX'_5, XX'_7, the central atom must accommodate 3, 5, or 7 bonds respectively.

This requires **expansion beyond 4 bonds** (sp^3d, sp^3d^2, sp^3d^3 hybridisation).

Larger halogen (e.g., I, Period 5): has **accessible d-orbitals** (5d) $\Rightarrow$ can expand its valence shell to accommodate 5 or 7 bonds.

Smaller halogen (e.g., F, Period 2): has **no d-orbitals** (n=2 shell) $\Rightarrow$ cannot expand beyond 4 bonds $\Rightarrow$ can only be at the periphery.

Also: larger atomic size physically accommodates more surrounding atoms.

[Ans] Final Answer

Option (3) [correct is (3)]: Larger halogen has accessible d-orbitals (can expand octet), greater size (accommodates multiple smaller atoms); smaller halogen lacks d-orbitals and cannot be the central atom in expanded species

[!] Common Mistake

Option (4) says "least electronegative is always central." This is not a reliable rule for all compounds. In interhalogens specifically, the central position is determined by d-orbital availability and size, not electronegativity. F (most electronegative) is NEVER the central atom specifically because it lacks d-orbitals and is the smallest.

SKILL: Conceptual Clarity (NCERT Example 7.16)

Q.37 [NCERT Example 7.16] Fluorine shows only -1 OS whereas others show $+1$, $+3$, $+5$, $+7$ also. Full explanation?

[A] Conceptual Approach

Both reasons must be combined: (1) Highest EN = never loses electrons to a more electronegative partner (no such partner exists). (2) No d-orbitals in n=2 shell = cannot expand octet to show 3, 5, or 7 bonds. Together = strictly confined to -1.

[S] Step-by-Step Solution

Two reasons, both required for a complete answer:

(1) Highest electronegativity (EN = 4.0):
F is the most electronegative element. It will never lose electrons to a more electronegative partner because no such partner exists. Positive OS requires F to be the electropositive partner (less electronegative) — impossible.

(2) No d-orbitals in valence shell (Period 2):
To show +3, +5, +7: the atom must excite electrons into d-orbitals (expand octet).
F's valence shell = 2s + 2p only. There are no 2d orbitals.
$\Rightarrow$ Cannot excite electrons to d-orbitals $\Rightarrow$ cannot show +3, +5, +7.

Cl, Br, I (Periods 3, 4, 5): have 3d, 4d, 5d orbitals accessible $\Rightarrow$ can show all positive OS.

[Ans] Final Answer

Option (2): F = most electronegative (never loses electrons) AND no d-orbitals (cannot expand octet) — both factors together confine F to -1 exclusively

[!] Common Mistake

Option (4) says "size alone is the limiting factor." Size is a consequence, not the cause. The causes are specifically: (a) highest EN and (b) no d-orbitals. A large halogen like I can show +7 not because it's large but because it has d-orbitals (5d) available. Size by itself doesn't determine OS.

SKILL: Application Skills (NCERT Intext 7.26)

Q.38 [NCERT 7.26] Comparing oxidising power of F_2 vs Cl_2 using bond dissociation enthalpy, electron gain enthalpy, and hydration enthalpy. Which correctly accounts for all three?

[A] Conceptual Approach

Three parameters, three outcomes: (1) BDE: F_2 lower BDE (easier to break) $\Rightarrow$ favours F_2. (2) EGE: Cl's EGE (-349) more negative than F's (-333) $\Rightarrow$ favours Cl_2. (3) Hydration enthalpy of halide: F^- has much higher hydration enthalpy than Cl^- $\Rightarrow$ strongly favours F_2. Net: 2 out of 3 factors favour F_2, including the dominant hydration enthalpy.

[S] Step-by-Step Solution

Parameter 1 — Bond dissociation enthalpy:
F_2 = 158 kJ/mol (low); Cl_2 = 242 kJ/mol (higher).
F_2 is easier to dissociate into F atoms $\Rightarrow$ **favours F_2** as stronger oxidiser.

Parameter 2 — Electron gain enthalpy:
F (-333 kJ/mol) vs Cl (-349 kJ/mol).
Cl EGE is more negative $\Rightarrow$ **favours Cl_2** in this parameter only.

Parameter 3 — Hydration enthalpy of X^-:
F^- has **much more negative hydration enthalpy** than Cl^- (because F^- is tiny and interacts very strongly with water).
$\Rightarrow$ F^- is stabilised far more in aqueous solution $\Rightarrow$ drives the F^- formation much more $\Rightarrow$ **strongly favours F_2**.

Net: 2 of 3 parameters (BDE and hydration) strongly favour F_2; only EGE favours Cl_2 slightly. F_2 is the stronger oxidising agent overall.

[Ans] Final Answer

Option (4): F_2 wins despite less negative EGE; lower BDE (easier to break) and far greater hydration enthalpy of F^- more than compensate $\Rightarrow$ F_2 stronger overall oxidising agent

[!] Common Mistake

Option (2) says "Cl_2 is stronger because its EGE is more negative." This looks at only ONE parameter. The question asks to account for ALL THREE. The hydration enthalpy of F^- is enormously more negative than Cl^- (F^- is tiny = interacts strongly with polar water). This single factor dominates and reverses the EGE advantage.

SKILL: Diagram-Based Visual Literacy (NCERT 7.26)

Q.39 Table: BDE (F_2 lower $\Rightarrow$ favours F_2), EGE (Cl more negative $\Rightarrow$ favours Cl), Hydration enthalpy of X^- (F^- much more negative $\Rightarrow$ favours F_2). Which statement is correctly supported?

[A] Conceptual Approach

Count the votes: F_2 favoured by 2 parameters (BDE and hydration enthalpy). Cl favoured by 1 parameter (EGE). The hydration enthalpy advantage for F^- is quantitatively very large. Net: F_2 is the stronger oxidising agent. Option (1) correctly states this.

[S] Step-by-Step Solution

From the table:

BDE: F_2 lower (easier to dissociate) $\Rightarrow$ **favours F_2**
EGE: Cl^- has more negative EGE $\Rightarrow$ **favours Cl_2** (this is the one parameter where Cl wins)
Hydration enthalpy: F^- far more negative $\Rightarrow$ **strongly favours F_2**

Two of three parameters favour F_2 (and hydration is quantitatively dominant).
Net result: F_2 is still the **stronger oxidising agent**.

This correctly explains why F_2's oxidising power dominates despite having a less negative EGE than Cl_2.

[Ans] Final Answer

Option (1): Two of three parameters (BDE and hydration enthalpy) favour F_2; only EGE favours Cl_2 — net combined effect makes F_2 the stronger oxidising agent

[!] Common Mistake

Option (3) says "since EGE favours Cl, Cl must be the stronger oxidising agent overall." You CANNOT conclude overall strength from just one parameter. All three must be considered. The hydration enthalpy of F^- is the single largest thermodynamic contribution, and it decisively favours F_2.

SKILL: Time Management and Strategic Thinking (NCERT 7.27, 7.28)

Q.40 Statements: I. Interhalogens form XX', XX'_3, XX'_5, XX'_7 (X = larger). II. F shows only -1 (most EN, no d-orbitals). III. F_2 stronger than Cl_2 mainly because Cl's EGE is less negative than F's. IV. Sea water source of Cl^-, Br^-, I^-. True ones?

[A] Conceptual Approach

I = TRUE. II = TRUE. III = FALSE (F_2 is stronger DESPITE Cl having more negative EGE; the reason is hydration enthalpy of F^- + lower BDE of F_2, NOT because of EGE). IV = TRUE. So I, II, IV correct.

[S] Step-by-Step Solution

Statement I: TRUE — Interhalogen types: XX' (e.g., ClF), XX'_3 (ClF_3), XX'_5 (ClF_5, BrF_5), XX'_7 (IF_7). X = larger/heavier halogen always central.

Statement II: TRUE — F: EN = 4.0 (most electronegative) + no d-orbitals $\Rightarrow$ strictly -1.

Statement III: **FALSE** — The statement says F_2 is stronger "mainly because Cl's EGE is less negative." This is exactly WRONG. Cl's EGE (-349) is actually MORE negative than F's (-333) — this parameter FAVOURS Cl. F_2 wins because of its LOWER BDE and the FAR greater hydration enthalpy of F^-, NOT because of EGE.

Statement IV: TRUE — Sea water contains dissolved NaCl (Cl^-), NaBr/$MgBr_2$ (Br^-), and trace NaI/KI (I^-). All three halides are found in sea water.

True: I, II and IV.

[Ans] Final Answer

Option (3) [correct is (3)]: I, II and IV only

[!] Common Mistake

Statement III reverses the EGE comparison. Cl's EGE IS more negative than F's ($-349 > -333$). F_2 is a stronger oxidising agent NOT because of EGE (which actually gives Cl an advantage) but despite it — the lower BDE and dominant hydration enthalpy of F^- overcome the EGE disadvantage.

TOPIC 5 — Anomalous Behaviour of Fluorine

Small Size, Highest Electronegativity, Low F–F BDE & Absence of d Orbitals

SKILL: Analytical Ability

Q.41 Most reactions of fluorine are exothermic. Why?

[A] Conceptual Approach

Exothermic = energy released > energy absorbed. Energy absorbed = breaking F–F bond (low BDE = easy). Energy released = forming E–F bonds (strong, compact, large). Net: huge energy released when new E–F bonds form. NCERT specifically says: "The bond formed by F with other elements is small and strong" — that strong E–F bond formation is what makes reactions exothermic.

[S] Step-by-Step Solution

Energy analysis of F_2 reactions:

Energy INPUT (to break F_2): F–F BDE = 158 kJ mol^{-1} (LOW — easy to break).

Energy RELEASED (when E–F bond forms): very **large**.
F forms characteristically **small, compact, very strong** bonds with almost every other element because:
• F is small $\Rightarrow$ close approach to other atom $\Rightarrow$ better orbital overlap $\Rightarrow$ strong bond
• F's high EN $\Rightarrow$ highly polar E–F bond

Net: large energy released (E–F bond formation) $\gg$ small energy absorbed (F–F bond breaking) $\Rightarrow$ net energy is released $\Rightarrow$ **exothermic**.

[Ans] Final Answer

Option (2): F forms small, strong bonds with virtually any element; large energy released on E–F bond formation more than compensates for the low BDE of F_2 $\Rightarrow$ most reactions are exothermic

[!] Common Mistake

Option (3) says "F's highly negative EGE makes every reaction exothermic." EGE is only one term in the overall energy balance. EGE refers to gaseous atom gaining an electron — not the full reaction enthalpy. The exothermicity is due to the strong E–F bonds formed, not EGE alone.

SKILL: Conceptual Clarity

Q.42 Fluorine's anomalous properties: which ones are HIGHER than expected and which are LOWER?

[A] Conceptual Approach

Think about what "small size and high EN" predict:
HIGHER: IE (harder to remove electrons from a compact, high-charge atom), EN (highest), electrode potential (hardest to reduce, strongest oxidant).
LOWER: Radius (smallest), mp/bp (weak F–F bonds + no additional London dispersion), BDE of F_2 (lone pair repulsion), EGE (orbital congestion).

[S] Step-by-Step Solution

Anomalously HIGHER than group trend predicts:

- Ionisation enthalpy (compact atom, electrons held tightly)
- Electronegativity (highest of all elements = 4.0)
- Electrode potential (E° = +2.87 V; strongest oxidant, hardest to reduce)

Anomalously LOWER than group trend predicts:

- Ionic and covalent radii (smallest in group)
- Melting and boiling points (weak F–F bond, small molecular size)
- Bond dissociation enthalpy of F_2 (lone pair repulsion)
- Electron gain enthalpy (orbital congestion in compact 2p)

Option (4) correctly classifies all of these.

[Ans] Final Answer

Option (4): Higher = IE, EN, electrode potential; Lower = ionic/covalent radii, mp, bp, BDE, EGE

[!] Common Mistake

Option (3) puts "boiling point" as higher than expected. F_2 bp = 85 K — this is **lower** than Cl_2 bp (239 K), following the expected van der Waals trend (lighter molecule = lower bp). F_2's bp is NOT anomalously high (unlike HF, which IS anomalously high due to H-bonding). F_2 has normal van der Waals interactions.

SKILL: Application Skills

Q.43 Fluorine forms only ONE oxoacid while Cl, Br, I form several. Best explanation?

[A] Conceptual Approach

Other halogens form HOCl, $HClO_2$, $HClO_3$, $HClO_4$ (Cl in +1, +3, +5, +7 OS). These require Cl to be in positive OS. Fluorine CANNOT have positive OS (most electronegative + no d-orbitals). So F can only form HOF (hypofluorous acid, F in -1, O in... wait, let me recalculate: in HOF, O is in what OS? H=+1, F=-1, so O: +1+O+(-1)=0 $\Rightarrow$ O=0. Unusual). The key: F cannot show +1, +3, +5, +7 states needed for most oxoacids.

[S] Step-by-Step Solution

Other halogens' oxoacids require the halogen in positive OS:
HOCl: Cl = +1; $HClO_2$: Cl = +3; $HClO_3$: Cl = +5; $HClO_4$: Cl = +7.

In these, the halogen is bonded to the more electronegative oxygen (gaining partial negative charge from halogen's perspective).

Why F cannot do this:
(1) F is MORE electronegative than O — F will NEVER take a positive OS in a bond with O.
(2) F has NO d-orbitals — cannot expand its covalence to form bonds beyond one O and one H.

$\therefore$ F can only form **HOF** (hypofluorous acid) where F is -1, O is formally 0.
All other oxoacids with F in positive states are **impossible**.

[Ans] Final Answer

Option (1): F cannot show positive OS (highest EN + no d-orbitals) $\Rightarrow$ cannot form the variety of oxoacids that Cl, Br, I form in +1, +3, +5, +7 states

[!] Common Mistake

Option (2) says "F cannot dissolve in water so cannot form acids." F_2 actually *reacts vigorously* with water ($2F_2 + 2H_2O \rightarrow O_2 + 4HF$). The reason F forms only one oxoacid is not solubility but the fundamental inability to take positive OS, which is required for forming the variety of oxoacids that other halogens have.

SKILL: Diagram-Based Visual Literacy

Q.44 Table summarises all F anomalies vs expected trends. Which statement is correctly supported?

[A] Conceptual Approach

Table shows: IE, EN, electrode potential = higher than expected; radii, mp, bp, BDE, EGE = lower than expected. Some higher, some lower. Option (3) correctly says: some properties are anomalously high, others anomalously low, all attributable to small size, highest EN, and no d-orbitals.

[S] Step-by-Step Solution

From the table analysis:

Higher than expected: Ionisation enthalpy, electronegativity, electrode potential.
Lower than expected: Radii, mp, bp, BDE, EGE.

Root causes (three main factors of fluorine anomaly):

- **Small size**: smaller radii, lone pair repulsion (low BDE), no room for more electrons (low EGE)
- **Highest electronegativity**: high IE, high electrode potential, forces positive OS on other elements
- **No d-orbitals**: restricted oxidation states, cannot form expanded-octet compounds

Option (3) correctly captures: some high, some low, all attributable to these three root causes.

[Ans] Final Answer

Option (3): Some properties anomalously high (IE, EN, electrode potential), others anomalously low (radii, mp, bp, BDE, EGE); all attributable to small size, highest EN, and absence of d-orbitals

[!] Common Mistake

Option (1) says "all of F's properties are higher than expected." The table clearly shows several properties (radii, mp, bp, BDE, EGE) are LOWER than expected. F's anomaly is that it deviates in BOTH directions — some higher, some lower. It is not uniformly "above trend" in everything.

SKILL: Time Management and Strategic Thinking

Q.45 Statements: I. F anomaly due to small size, highest EN, low F–F BDE, no d-orbitals. II. Most F reactions exothermic (bonds formed are small and strong). III. HF is a gas at room temperature like all other HX. IV. F forms only one oxoacid while others form many. True ones?

[A] Conceptual Approach

I = TRUE. II = TRUE. III = FALSE (HF is a liquid at room temperature! HF bp = 293 K = 20°C ≈ room temperature. Actually HF boiling point is 19.5°C = just at room temperature. Some sources say gas at 25°C, others say liquid. NCERT says HF is a liquid due to extensive H-bonding). IV = TRUE. Let me check III carefully: HF bp = 293 K = 19.5°C. At 25°C, HF would be a gas (above its bp). But it's associated through H-bonds... In NCERT, HF is described as having "an anomalously high boiling point" and "associated in the liquid state." At room temperature (298 K > 293 K bp), HF would technically be a gas. However, NCERT's point is that it SHOULD be a gas but behaves anomalously. Statement III says it IS a gas "just like all other HX" — the word "just like" (implying no anomaly) makes it FALSE because HF's behaviour is anomalous.

[S] Step-by-Step Solution

Statement I: TRUE — The four causes of F's anomaly: small size, highest EN, low F–F BDE, no d-orbitals.

Statement II: TRUE — Strong, small E–F bonds formed ⇒ large energy released ≫ small energy to break F_2.

Statement III: **FALSE** — HF is *NOT* "just like all other hydrogen halides." HF has an anomalously high boiling point (293 K = 19.5°C) due to extensive hydrogen bonding, making its behaviour distinctly different from HCl/HBr/HI. The phrase "just like" implies no anomaly, which is incorrect.

Statement IV: TRUE — F's restriction to -1 OS means it cannot form HOF-like molecules with F in higher positive states that other halogens' oxoacids require.

True: I, II and IV.

[Ans] Final Answer

Option (2) [correct is (2)]: I, II and IV only

[!] Common Mistake

Statement III is subtle. HF's boiling point (19.5°C = 293 K) means it is barely a gas at room temperature (25°C). More importantly, HF is fundamentally anomalous compared to other HX: it has strong H-bonding, is a weak acid (unlike HCl which is strong), and boils much higher than expected. Saying HF behaves "just like all other HX" completely ignores these anomalies.

DPP-10 Complete!

"Fifty questions, one chapter, one anomaly.

Fluorine breaks every rule — its EGE is less negative than chlorine's, its F–F bond is weaker than Cl–Cl, its HF is a weak acid despite being most polar,
and yet it is the strongest oxidant of all halogens.

Every one of these apparent contradictions has a single root:
small size + highest EN + no d-orbitals.
You now know the root. Every branch makes sense from here."

— Your Weird Chemist