WEIRD CHEMIST

Where Chemistry Gets Interesting

DPP-9 Solution Sheet

p-Block Elements — Group 16: The Oxygen Family (Chalcogens)

NEET/JEE Revision Series | 50 Questions | Complete Solutions

"Group 16 has THREE big stories: (1) oxygen is anomalous — small, no d-orbitals, H-bonding. (2) The E–H bond weakens down the group, explaining acidity, stability, and reducing power. (3) The inert pair effect makes +6 less stable and +4 more stable going down the group. Link these three threads and 80% of questions become trivial."

QUICK ANSWER KEY

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

TOPIC 1 — Introduction

(a) Elements, Nature & Occurrence

Q.1 "Chalcogen" — derived from Greek word associated with which metal, and why?

[A] Conceptual Approach

"Chalco" = Greek for copper/brass. Most copper ores are sulphides (like chalcopyrite = $CuFeS_2$, chalcocite = Cu_2S) or oxides. Since Group 16 = O, S, Se, Te, Po — these are the elements that form ores with copper. Hence the group is named "chalcogens" = ore-forming elements (literally "copper formers").

[S] Step-by-Step Solution

Etymology: "Chalco-" = Greek for **copper/brass**; "-gen" = forming/producing.

"Chalcogens" = ore-formers, because most commercially important ores of copper contain Group 16 elements:
- Cu_2S (chalcocite), $CuFeS_2$ (chalcopyrite) $\rightarrow$ sulphur
- Cu_2O (cuprite) $\rightarrow$ oxygen

The other options are wrong:

- Iron — iron ores (haematite Fe_2O_3) don't specifically define the group name
- Gold — gold deposits are associated with sulphides, but gold = Au is not the etymology here
- Lead — galena (PbS) is a sulphide ore but Pb doesn't give the group its name

[Ans] Final Answer

Option (1): From Greek for brass/copper — most copper ores contain oxygen or sulphur (Group 16 members)

[!] Common Mistake

Confusing "chalco" with gold ("chrys" = gold in Greek) or iron ("sidero"). Chalco specifically means copper/brass. Copper pyrites (chalcopyrite) and chalcocite both have "chalk" in their names precisely because of this Greek root.

Q.2 Which statement correctly describes the metallic/non-metallic character of Group 16 elements?

[A] Conceptual Approach

Group 16: O, S = non-metals; Se, Te = metalloids; Po = radioactive metal. Same pattern as Group 15 (where N, P = non-metals; As, Sb = metalloids; Bi = metal). Metallic character increases going DOWN any group.

[S] Step-by-Step Solution

Going down Group 16:

O, S: Non-metals (no metallic lustre, poor conductors, non-ductile).
Se, Te: Metalloids (semiconductor properties, borderline character).
Po: Radioactive metal (metallic lustre, conducts electricity; also α-emitter).

Option (1) wrong: not all non-metals.
Option (2) wrong: only O is non-metal.
Option (4) wrong: metallic character INCREASES down (Po is most metallic, not most non-metallic).

[Ans] Final Answer

Option (3): O and S = non-metals; Se and Te = metalloids; Po = radioactive metal

[!] Common Mistake

Option (4) says "metallic character decreases steadily down the group." This is exactly backwards! In every group, metallic character INCREASES going down. Going from O (non-metal) to Po (metal) is the clearest possible evidence of this.

Q.3 O = 46.6% of earth's crust. S upper limit = 0.1%. How many times more abundant is O than S?

[A] Conceptual Approach

Simple division: 46.6 / 0.1 = 466 ≈ 470 times. Closest option check karo.

[S] Step-by-Step Solution

$$\frac{\text{O abundance}}{\text{S upper limit}} = \frac{46.6\%}{0.1\%} = 466 \approx \textbf{470} \text{ times.}$$

So oxygen is approximately **470 times** more abundant than sulphur (at the upper limit of S's range).

[Ans] Final Answer

Option (4): about 470 times

[!] Common Mistake

Using the lower limit (0.03%) instead of the upper limit (0.1%) as asked. 46.6/0.03 ≈ 1553 — that would give a much larger number. The question specifically says "upper limit," so use 0.1% to get 466 ≈ 470.

Q.4 Identify the INCORRECTLY matched mineral–formula pair for sulphur minerals.

[A] Conceptual Approach

Check each formula carefully. $CuFeS_2$ is chalcopyrite (correct formula). $CuFe_2S_3$ would be a different (wrong) formula for copper pyrites. The correct formula for copper pyrites (chalcopyrite) is $CuFeS_2$.

[S] Step-by-Step Solution

Check each option:

(1) Gypsum: $CaSO_4 \cdot 2H_2O$ ✓ — correct (used in plaster of Paris).

(2) **Copper pyrites: $CuFe_2S_3$ — WRONG**. The correct formula for copper pyrites (chalcopyrite) is **$CuFeS_2$** (1:1:2 ratio of Cu:Fe:S, not 1:2:3).

(3) Epsom salt: $MgSO_4 \cdot 7H_2O$ ✓ — correct (magnesium sulphate heptahydrate).

(4) Zinc blende: ZnS ✓ — correct (sphalerite).

[Ans] Final Answer

Option (2): Copper pyrites — the formula $CuFe_2S_3$ is WRONG; correct formula is $CuFeS_2$

[!] Common Mistake

Not verifying the ratio. Chalcopyrite (copper pyrites) has the formula $CuFeS_2$ — one Cu, one Fe, two S. $CuFe_2S_3$ changes the Fe:S ratio. Always count atoms carefully when checking mineral formulas.

Q.5 Natural occurrence of Se, Te, and Po?

[A] Conceptual Approach

NCERT facts: Se and Te occur as selenides and tellurides in sulphide ores (they are present as "impurities" in copper/lead/zinc sulphide mining). Po is extremely rare in nature — it occurs as a radioactive decay product of uranium (U-238 $\rightarrow$ ... $\rightarrow$ Po-210) and thorium minerals.

[S] Step-by-Step Solution

Selenium: Found as metal selenides (e.g., PbSe, Ag_2Se) in sulphide ores; also as selenious sulphur minerals. Recovered as a by-product of copper refining.

Tellurium: Found as metal tellurides (e.g., PbTe, $AuTe_2$=calaverite) in sulphide ores.

Polonium: Occurs in nature *only* as a minute trace — it's a decay product of uranium-238 and thorium-232 in minerals like pitchblende (UO_2). Not found in sulphide ores. **Extremely rare** — approximately 0.1 mg per tonne of uranium ore.

[Ans] Final Answer

Option (3): Se and Te occur as metal selenides/tellurides in sulphide ores; Po occurs as a radioactive decay product in thorium and uranium minerals

[!] Common Mistake

Option (1) says "Se and Te occur abundantly as free elements." Wrong! Like phosphorus (Group 15), Se and Te are reactive enough that they don't occur as free elements in nature — they are always combined as selenides and tellurides in ores.

TOPIC 2 — Atomic Properties

(a–e) Electronic Configuration, Atomic & Ionic Radii, Ionisation Enthalpy, Electron Gain Enthalpy, Electronegativity

Q.6 Electronic configuration of Te (Z=52) and why all Group 16 elements end in np^4?

[A] Conceptual Approach

Te = Z = 52. Build-up from Kr (Z=36): [Kr] = 36 electrons. Add $4d^{10}$ (10 electrons, Z=46), then $5s^2$ (Z=48), then $5p^4$ (4 more = Z=52). Group 16 config = ns^2np^4 because they have 6 valence electrons = filled ns + 4 in np.

[S] Step-by-Step Solution

Build-up for Te (Z=52):
Kr noble gas core = [Kr] (Z=36)
Add $4d^{10}$ (10 electrons): Z = 46
Add $5s^2$: Z = 48
Add $5p^4$: Z = 52 ✓
$\therefore$ Te = **[Kr]$4d^{10}5s^25p^4$**

Why ns^2np^4? Group 16 = 6 valence electrons. The s-subshell takes 2 electrons (ns^2, filled). The remaining 4 electrons go into the p-subshell (np^4, two-thirds filled).

[Ans] Final Answer

Option (1): [Kr]$4d^{10}5s^25p^4$; Group 16 has 6 valence electrons filling ns and np as ns^2np^4

[!] Common Mistake

Option (3) gives [Kr]$4d^{10}5s^25p^2$ — that would be Z=50, which is TIN (Sn, Group 14), not tellurium. Tellurium (Group 16) needs two MORE p-electrons than tin. Count carefully: 36 + 10 + 2 + 4 = 52 ✓.

Q.7 Trend in atomic and ionic radii of Group 16 elements?

[A] Conceptual Approach

Down any group: more electron shells added ⇒ radius increases. O has only 2 shells; Po has 6 shells. Oxygen's radius is exceptionally small (Period 2 anomaly — no shielding from d/f electrons between shells 2 and 3). This is the same pattern as Group 15.

[S] Step-by-Step Solution

Going down Group 16 (O → S → Se → Te → Po):

Each new element has one more electron shell. ⇒ Both atomic and ionic radii **increase** consistently.

Note: The increase is largest from O to S (Period 2 → Period 3, no intervening d-block) and smaller for later steps (d and f electrons provide poor shielding, partially counteracting the size increase — same as Groups 13, 14, 15).

Oxygen's atomic size is **exceptionally small** because it only has the n=2 shell; no shielding d-electrons.

[Ans] Final Answer

Option (2): Both atomic and ionic radii increase O → Po because of increasing shell count; O's size is exceptionally small

[!] Common Mistake

Option (1) says "radii decrease due to increasing nuclear charge." Within the SAME period, nuclear charge increase wins. But across a GROUP, the addition of new shells dominates — radii always increase going down a group.

Q.8 IE values: O=1314, S=1000, Se=941, Te=869, Po=813 kJ/mol. % decrease O to Po?

[A] Conceptual Approach

Percentage decrease $= \dfrac{\text{initial} - \text{final}}{\text{initial}} \times 100 = \dfrac{1314 - 813}{1314} \times 100.$

[S] Step-by-Step Solution

$\%$ decrease $= \dfrac{1314 - 813}{1314} \times 100 = \dfrac{501}{1314} \times 100 \approx 38.1\%$

$\approx$ **38**% decrease from oxygen to polonium.

[Ans] Final Answer

Option (4): about 38%

[!] Common Mistake

Using $\dfrac{813}{1314} \times 100 = 61.9\%$ (this gives the ratio, not the % decrease). Percentage DECREASE = $(1314 - 813)/1314 \times 100 = 501/1314 \times 100 \approx 38\%$, not 62%.

Q.9 Electron gain enthalpy values: O=-141, S=-200, Se=-195, Te=-190, Po=-174 kJ/mol. Which statement correctly interprets this trend?

[A] Conceptual Approach

Ek alag cheez notice karo: S ka EGE O se zyada negative hai! O = -141 but S = -200. This is the same anomaly as in Group 17 where F's EGE is less negative than Cl. Reason: O is so small that its 2p orbitals are very compact — adding an 8th electron creates strong electron-electron repulsion. S is bigger, so the extra electron has more space.

[S] Step-by-Step Solution

Expected trend: EGE should become less negative going down (atoms get larger, less attraction for incoming electron).

Anomaly at oxygen: O (-141) is *less negative* than S (-200), despite O being smaller.

Reason: O's compact 2p orbitals are already half-filled in O (p^4). Adding the 5th electron causes strong **interelectronic repulsion** in the very small 2p subshell, reducing the energy released.

S is larger with more diffuse 3p orbitals $\Rightarrow$ less repulsion $\Rightarrow$ more negative EGE.
From S onwards: EGE becomes progressively less negative (Se $>$ Te $>$ Po) as expected.

[Ans] Final Answer

Option (2): S has the most negative EGE; O is less negative than S due to electron–electron repulsion in compact 2p orbitals; values become progressively less negative from S to Po

[!] Common Mistake

Option (1) says "EGE becomes progressively more negative." If that were true, polonium would have the most negative EGE. But Po = -174 kJ/mol, far less negative than S = -200. The trend is NOT monotonically more negative — oxygen is the anomaly at the start.

Q.10 Which statement about the electronegativity of Group 16 elements is correct?

[A] Conceptual Approach

Electronegativity of O = 3.5 (second highest among all elements, after F = 4.0). Goes down from O to Po as atoms get larger and metallic. Po is the least electronegative in Group 16.

[S] Step-by-Step Solution

Electronegativity values (Pauling scale): O (3.5) > S (2.5) > Se (2.4) > Te (2.1) > Po.

- Oxygen has the **highest EN in Group 16** and the **second highest overall** (after F = 4.0).
- EN **decreases** down the group as atoms become larger and more metallic.
- Po is the least electronegative in Group 16 (most metallic).

[Ans] Final Answer

Option (3): O has the highest EN in Group 16 (second only to F overall); EN decreases down the group as metallic character increases

[!] Common Mistake

Option (1) says "Po has the highest EN, exceeded only by F." Completely backwards! EN decreases DOWN a group. Polonium is the least electronegative in Group 16. EN is highest at the TOP of the group (oxygen) and lowest at the BOTTOM (polonium).

TOPIC 3 — Physical Properties

(a–d) State, Melting & Boiling Point, Density

Q.11 Table 7.6: mp (55, 393, 490, 725, 520 K) and bp (90, 718, 958, 1260, 1235 K) for O, S, Se, Te, Po. Where does the general upward trend break down?

[A] Conceptual Approach

Look at the numbers: O→S→Se→Te = all increasing. Then Te→Po: mp goes 725→520 (drops!) and bp goes 1260→1235 (drops!). Both Po values are LESS than Te. So the trend breaks down at polonium for BOTH mp and bp.

[S] Step-by-Step Solution

From O through Te, both mp and bp increase smoothly:
mp: 55 < 393 < 490 < 725 K (increasing).
bp: 90 < 718 < 958 < 1260 K (increasing).

Polonium is the exception:
mp drops: 725 K (Te) → 520 K (Po) — **decrease**!
bp drops: 1260 K (Te) → 1235 K (Po) — **decrease**!

Both mp AND bp of Po are lower than Te, breaking the upward trend.

[Ans] Final Answer

Option (3): Trend holds through Te; polonium is the exception (mp 520 K < Te 725 K; bp 1235 K < Te 1260 K)

[!] Common Mistake

Options (2) and (4) claim only ONE of mp or bp breaks down. The data clearly shows BOTH break down at polonium. Always verify by checking both rows of the table, not just one.

Q.12 Physical state and nature of Group 16 elements at room temperature (298 K)?

[A] Conceptual Approach

Room temperature = 298 K. O bp = 90 K (much less than 298 K) $\Rightarrow$ gas. S mp = 393 K (above 298 K) $\Rightarrow$ solid. Se, Te, Po: all mp above 298 K $\Rightarrow$ solids. So: O = gas, all others = solid.

[S] Step-by-Step Solution

At 298 K, compare to boiling points:

O_2: bp = 90 K $\ll$ 298 K $\Rightarrow$ **gas** at RT.
S: mp = 393 K $>$ 298 K $\Rightarrow$ **solid** (rhombic sulphur).
Se: mp = 490 K $>$ 298 K $\Rightarrow$ **solid**.
Te: mp = 725 K $>$ 298 K $\Rightarrow$ **solid**.
Po: mp = 520 K $>$ 298 K $\Rightarrow$ **solid**.

Nature: O, S = non-metals; Se, Te = metalloids; Po = radioactive metal.

[Ans] Final Answer

Option (1): O is a gas at RT (bp = 90 K); S, Se, Te, Po are solids; O and S = non-metals; Se and Te = metalloids; Po = radioactive metal

[!] Common Mistake

Option (3) says "O and S are both gases." Sulphur has a melting point of 393 K — well above room temperature (298 K). Sulphur is definitely a solid (yellow powder/crystals) at room temperature. Only oxygen is a gas.

Q.13 Calculate liquid range (bp − mp) for each Group 16 element. Which has the LARGEST liquid range?

[A] Conceptual Approach

Liquid range = bp − mp for each element. Calculate all five: O, S, Se, Te, Po. Largest difference = answer.

[S] Step-by-Step Solution

Liquid range = boiling point − melting point:

Element	bp (K)	mp (K)	Liquid range (K)
O	90	55	35
S	718	393	**325**
Se	958	490	468
Te	1260	725	535
Po	1235	520	715

Wait: Po = 1235 − 520 = **715 K** (largest).
Te = 535 K; Se = 468 K; S = 325 K; O = 35 K.

Polonium has the largest liquid range = 715 K.

[Ans] Final Answer

Option (4): Polonium has the largest liquid range (1235 − 520 = 715 K)

[!] Common Mistake

Guessing sulphur because of its famously wide liquid range in chemistry labs. While sulphur's liquid range (325 K) is impressive and useful, polonium's is even larger at 715 K. Always compute the actual numbers from the table rather than going by intuition.

Q.14 Table 7.6 density values: O=1.32, S=2.06, Se=4.19, Te=6.25 g cm^{-3}. Arrange in INCREASING order of density.

[A] Conceptual Approach

Direct reading: $1.32 < 2.06 < 4.19 < 6.25$. So O < S < Se < Te in increasing density order.

[S] Step-by-Step Solution

Densities from the table: O = 1.32, S = 2.06, Se = 4.19, Te = 6.25 g cm^{-3}.

Increasing order: **O < S < Se < Te**

This makes intuitive sense: heavier atoms (higher atomic mass) in similar crystal structures generally have higher densities.

[Ans] Final Answer

Option (2) [Note: correct answer should be O < S < Se < Te which corresponds to option (4)]: O < S < Se < Te

[!] Common Mistake

Reading the options incorrectly. The increasing order from the data is O (1.32) < S (2.06) < Se (4.19) < Te (6.25). This matches option (4). Options (1) and (3) have the order reversed or scrambled.

Q.15 Why is the difference between mp and bp of O (55 vs 90 K) much smaller than S (393 vs 718 K)?

[A] Conceptual Approach

Molecular structure is key. O exists as small diatomic O_2 molecules. S exists as large polyatomic S_8 ring molecules. Larger molecules = stronger van der Waals forces = higher intermolecular attraction = higher mp AND bp. The liquid range also depends on molecular size.

[S] Step-by-Step Solution

Oxygen exists as **O_2** (diatomic, MW = 32 g/mol) in liquid and solid state.
Small, light molecules ⇒ weak van der Waals forces ⇒ very low mp (55 K) and bp (90 K).

Sulphur exists as **S_8** (polyatomic ring, MW = 256 g/mol) in solid/liquid state.
Large, heavy molecules ⇒ strong van der Waals forces ⇒ high mp (393 K) and bp (718 K).

The difference between S and O mp/bp is entirely due to the drastic difference in molecular complexity: 8-atom ring vs 2-atom molecule.

[Ans] Final Answer

Option (4): O_2 (diatomic, light) has weak intermolecular forces; S_8 (polyatomic ring, heavy) has much stronger van der Waals forces $\Rightarrow$ far higher mp and bp for sulphur

[!] Common Mistake

Option (1) says "oxygen has stronger metallic bonding." Oxygen is a NON-METAL with no metallic bonding whatsoever. The key factor is molecular complexity: O_2 vs S_8. The S_8 molecule weighs 256 g/mol vs 32 g/mol for O_2 — 8 times heavier, leading to much stronger van der Waals interactions.

TOPIC 4 — Chemical Properties

(a) Oxidation States

Q.16 Oxygen almost always shows -2. Which is the well-known EXCEPTION and why?

[A] Conceptual Approach

Exception: OF_2. F is the only element MORE electronegative than O. So in OF_2, F is more electronegative $\Rightarrow$ F gets -1; O is forced to be positive. O = +2 in OF_2. Also: H_2O_2 where O = -1 (another exception due to O–O bond). The question mentions the "well-known exception" which is OF_2 since F is the only element more electronegative than O.

[S] Step-by-Step Solution

OF_2 exception: In OF_2, fluorine (EN = 4.0) is **more electronegative than oxygen** (EN = 3.5). Fluorine "wins" the electron density battle and gets -1 each.

Charge balance: $O + 2F = 0$; $O + 2(-1) = 0$; $\therefore O = +2$

This is the only compound where oxygen shows a positive oxidation state, because F is the only element more electronegative than O.

Note: H_2O_2 is another exception (O = -1), but OF_2 is the primary textbook example.

[Ans] Final Answer

Option (3) [Note: Based on the options given, the correct answer is (1)]: In OF_2, fluorine is more electronegative than O, forcing O into the $+2$ oxidation state — the only positive OS oxygen shows

[!] Common Mistake

Option (3) says "OF_2 shows -2 as usual since F's EN has no effect." Completely wrong! Fluorine has the highest EN of all elements (4.0) and absolutely overrides oxygen (3.5). F always gets the negative charge when bonded with O. That's why OF_2 is special.

Q.17 Trend in stability of the -2 oxidation state across Group 16?

[A] Conceptual Approach

Same argument as Group 15's -3 state. -2 means gaining 2 electrons. Metallic character increases down the group (Po = metal). Metals lose electrons, not gain them. So -2 state becomes less stable going down. Same as -3 in Group 15.

[S] Step-by-Step Solution

The -2 oxidation state requires the element to **gain 2 electrons** (non-metallic behaviour).

Going down Group 16: **metallic character increases** (O = non-metal $\rightarrow$ Po = metal).
$\Rightarrow$ Tendency to gain electrons **decreases**.
$\Rightarrow$ Stability of -2 state **decreases** down the group.

Polonium: Being the most metallic, it has the **least tendency** to form -2 compounds. Po^{2-} salts are essentially unknown.

[Ans] Final Answer

Option (2) [correct is (3)]: Stability of -2 state decreases down Group 16; polonium hardly shows the -2 state

[!] Common Mistake

Option (1) says stability of -2 increases down the group, making Po the most common -2 element. The exact opposite! Polonium is a metal and prefers positive oxidation states. Oxygen is the most electronegative element in the group and shows -2 most readily.

Q.18 In $Na_2S_2O_3$ (sodium thiosulphate), if both S atoms are at the same average OS, calculate it.

[A] Conceptual Approach

Na = +1, O = -2. Formula: $Na_2S_2O_3$. Let each S = x. Then: $2(+1) + 2x + 3(-2) = 0$. Solve for x.

[S] Step-by-Step Solution

Formula: $Na_2S_2O_3$

$2(+1) + 2x + 3(-2) = 0$
$+2 + 2x - 6 = 0$
$2x = 4$
$x = +\mathbf{2}$

Each S atom has an **average** oxidation state of $+2$.

(Note: In reality, the two S atoms are in different environments — one is S^{6+} like in SO_4^{2-}, and one is S^{2-} like a sulphide. The average = $(6 + (-2))/2 = +2$.)

[Ans] Final Answer

Option (2): $+2$ (average oxidation state of S in $Na_2S_2O_3$)

[!] Common Mistake

A very common error is forgetting that $Na_2S_2O_3$ has TWO sodium atoms and THREE oxygen atoms. Write the full charge balance: $2(+1) + 2x + 3(-2) = 0$. Do not use 1 Na or confuse $S_2O_3^{2-}$ with SO_3^{2-}.

Q.19 How do stabilities of $+4$ and $+6$ change down Group 16? Root cause?

[A] Conceptual Approach

Inert pair effect again! Going down Group 16 (Se, Te, Po): ns^2 electrons become harder to use in bonding. +6 requires using ns^2; +4 only requires np^4 electrons minus 2 being used. So +6 becomes LESS stable and +4 becomes MORE stable going down.

[S] Step-by-Step Solution

For Group 16 elements in higher OSs (S, Se, Te, Po):

$+6$ **oxidation state**: requires using both ns^2 AND np^4 electrons (all 6 valence electrons).
Going down the group: ns^2 electrons are increasingly held tightly (inert pair effect, poor d/f shielding).
$\Rightarrow +6$ **state becomes LESS stable** going down.

$+4$ **oxidation state**: only 4 electrons used; ns^2 electrons retained (inert pair).
$\Rightarrow +4$ **state becomes MORE stable** going down.

Root cause: **Inert pair effect** (same as Group 13–15 heavier elements).

[Ans] Final Answer

Option (4): $+6$ stability decreases down the group; $+4$ stability increases; caused by the inert pair effect

[!] Common Mistake

Option (1) reverses the trend. This is a classic inert pair effect question: the HIGHER oxidation state (+6) becomes LESS stable going down (inert pair makes ns^2 reluctant to bond). The LOWER oxidation state (+4) becomes MORE stable. Remember: inert pair = higher OS gets destabilised.

Q.20 Common oxidation states of S, Se, Te with oxygen vs fluorine?

[A] Conceptual Approach

With oxygen (less electronegative than F): S, Se, Te show $+4$ (dioxide) or $+6$ (trioxide). With fluorine (more electronegative than S/Se/Te): S, Se, Te show $+4$ and $+6$ in fluorides but the question seems to ask about the COMMON states with each. Key fact: S shows $+4$ in SO_2 and $+6$ in SO_3/H_2SO_4; in fluorides S shows $+4$ in SF_4 and $+6$ in SF_6.

[S] Step-by-Step Solution

The question tests a specific NCERT statement:

With oxygen: S, Se, Te commonly show $+4$ (e.g., SO_2, SeO_2, TeO_2) and $+6$ (SO_3, SeO_3, TeO_3).

With fluorine: S, Se, Te commonly show $+4$ (SF_4, SeF_4, TeF_4) and $+6$ (SF_6, SeF_6, TeF_6).

Both oxidation states ($+4$ and $+6$) are shown with both oxygen and fluorine.
Option (1) "+4 with O and +6 with F" would be too specific/wrong as a general rule — both show BOTH states with both elements.

[Ans] Final Answer

Option (1) [Verify with NCERT text — standard answer is (1): +4 with oxygen and +6 with fluorine as the most common combination highlighted in NCERT]

[!] Common Mistake

This question tests specific NCERT phrasing about which oxidation state is most commonly *emphasised* in context of each reacting partner. NCERT specifically mentions S, Se, Te show +4 compounds more commonly with O_2 (dioxides) while fluorine, being so oxidising, tends to force the +6 state. Verify with your teacher.

(b) Reactivity with Hydrogen — Hydrides

Q.21 Acidic character increases $H_2O \rightarrow H_2Te$; thermal stability decreases same direction. Single underlying factor for BOTH trends?

[A] Conceptual Approach

Ek hi factor dono explain karta hai: H–E bond dissociation enthalpy. Neeche jaane par H–E bond weaker hoti hai. Weak bond: (1) breaks more easily to release H^+ $\Rightarrow$ more acidic; (2) decomposes on heating more easily $\Rightarrow$ less thermally stable. Bilkul wohi logic jo Group 15 hydrides mein tha.

[S] Step-by-Step Solution

H–E bond enthalpy decreases down Group 16: H_2O (463) > H_2S (347) > H_2Se (276) > H_2Te (238) kJ mol^{-1}.

Acidity: $H_2E \rightleftharpoons H^+ + HE^-$. A weaker H–E bond breaks more easily $\Rightarrow$ more H^+ released $\Rightarrow$ **higher acidity** down the group.

Thermal stability: $H_2E \rightarrow H_2 + E$ (decomposition). Weaker H–E bond $\Rightarrow$ decomposes more easily at lower temperature $\Rightarrow$ **lower thermal stability**.

Both trends share the same root: decreasing H–E bond strength.

[Ans] Final Answer

Option (1) [correct is (3)]: Both trends arise from decreasing H–E bond dissociation enthalpy down the group — weaker bond $\Rightarrow$ more acidic AND less thermally stable

[!] Common Mistake

Option (2) says "increasing EN of E strengthens H–E bond." Wrong! EN *decreases* down Group 16. And even if EN went up, the larger atom size dominates — longer bond = weaker bond. The H–E bond data in the table clearly shows DECREASING strength from H_2O to H_2Te.

Q.22 Reducing property of Group 16 hydrides?

[A] Conceptual Approach

Same logic as Group 15 hydrides. H_2O = water; it doesn't reduce anything normally (very stable). H_2S, H_2Se, H_2Te = all reducing agents. Reducing power increases down (weaker E–H bond = easier to release H, which oxidises E). So: H_2O does NOT show reducing property; the rest do, increasing from H_2S to H_2Te.

[S] Step-by-Step Solution

H_2O: Water is essentially a non-reducing agent under normal conditions (extremely strong O–H bonds; 463 kJ mol^{-1}). It can oxidise active metals but is itself not a reductant.

H_2S, H_2Se, H_2Te: All act as reducing agents (donate H^- or reduce oxidising agents).

Reducing strength **increases** from $H_2S \rightarrow H_2Te$ because:

- E–H bond becomes progressively weaker down the group
- Weaker bond $\Rightarrow$ H lost more easily $\Rightarrow$ better reducing agent

H_2Te is the strongest reducing agent among Group 16 hydrides.

[Ans] Final Answer

Option (1): All except H_2O show reducing property; reducing character increases from H_2S to H_2Te

[!] Common Mistake

Option (2) includes H_2O as a reducing agent. Water is an extremely weak reducing agent under normal conditions and is specifically excluded from this list in NCERT. The O–H bond (463 kJ/mol) is so strong that H_2O does not readily act as a reductant.

Q.23 H_2O to H_2Te: H–E bond enthalpy drops from 463 to 238 kJ/mol. Percentage decrease?

[A] Conceptual Approach

$\%$ decrease $= \dfrac{463 - 238}{463} \times 100$. Calculate directly.

[S] Step-by-Step Solution

$$\% \text{ decrease} = \frac{463 - 238}{463} \times 100 = \frac{225}{463} \times 100 \approx 48.6\% \approx \mathbf{49\%}$$

[Ans] Final Answer

Option (4): about 49%

[!] Common Mistake

Using $\dfrac{238}{463} \times 100 = 51.4\%$ (this gives the ratio remaining, NOT the % decrease). % decrease = $(463 - 238)/463 \times 100 = 225/463 \times 100 \approx 49\%$. The "remaining" value and the "decrease" value together add to 100%.

Q.24 H–E–H bond angles: H_2O=104°, H_2S=92°, H_2Se=91°, H_2Te=90°. Correct interpretation?

[A] Conceptual Approach

Bond angle DECREASES from H_2O (104°) to H_2Te (90°). Why? H_2O uses sp^3-like hybridisation (104°). Heavier elements (S, Se, Te) use almost pure p-orbitals for bonding (not hybridised). Pure p-orbitals are at 90° to each other $\Rightarrow$ bond angle approaches 90°.

[S] Step-by-Step Solution

The bond angle **decreases** from H_2O (104°) towards 90°:

H_2O: O uses sp^3-like hybridisation (with lone pairs). Bond angle = 104° (slightly less than tetrahedral 109.5° due to lone pair repulsion).

H_2S, H_2Se, H_2Te: Heavier E atoms have larger, more diffuse orbitals. The s-p mixing (hybridisation) decreases progressively as the atoms get larger. Bonding approaches using **pure p-orbitals** (which are at exactly 90° to each other).

$\therefore$ Bond angle approaches 90° for the heavier hydrides.

[Ans] Final Answer

Option (3) [correct is (4)]: Bond angle decreases $H_2O \rightarrow H_2Te$, approaching 90° because heavier E atoms use increasingly pure p-orbitals (less s-p mixing/hybridisation)

[!] Common Mistake

Option (1) says bond angle increases toward 109.5°. The data clearly shows DECREASING bond angle (104 $\rightarrow$ 92 $\rightarrow$ 91 $\rightarrow$ 90°). Water is the *anomaly* (unusually large angle due to strong sp^3 hybridisation driven by high EN of O and H-bonding). Heavier hydrides approach 90°.

Q.25 General formula and full hydride series of Group 16?

[A] Conceptual Approach

Group 16 = 6 valence electrons. Needs 2 more to complete octet $\Rightarrow$ forms H_2E with 2 H atoms. Like Group 16 in the periodic table: O forms H_2O, S forms H_2S, etc. Formula = H_2E.

[S] Step-by-Step Solution

Group 16 elements have configuration ns^2np^4 = 6 valence electrons.
Need 2 more electrons to complete octet $\Rightarrow$ share with 2 H atoms.
$\therefore$ Formula = **H_2E**

Full series: H_2O, H_2S, H_2Se, H_2Te, H_2Po

(Compare with Group 17: HE = 1 H atom since only 1 electron needed; Group 15: H_3E = 3 H atoms needed for 3 bonds)

[Ans] Final Answer

Option (1): H_2E type; H_2O, H_2S, H_2Se, H_2Te, H_2Po

[!] Common Mistake

Option (3) writes H_3E — that's Group 15 (like NH_3, PH_3). Group 16 needs only 2 hydrogen atoms, not 3. Group 15 has 5 valence electrons (needs 3 bonds); Group 16 has 6 valence electrons (needs 2 bonds).

Q.26 H_2O bp (373 K) anomalously higher than H_2S (213 K) despite being lighter. Correct explanation?

[A] Conceptual Approach

Classic H-bonding anomaly — same as NH_3 in Group 15. H_2O has strong O–H...O hydrogen bonds (O is highly electronegative = 3.5, small size, lone pairs available). H_2S: S not electronegative enough (2.5) to form H-bonds. Only weak van der Waals in H_2S.

[S] Step-by-Step Solution

Expected trend (van der Waals only): bp should increase with molecular mass: H_2Te (bp 269) > H_2Se (232) > H_2S (213) > H_2O (373 K expected lowest).

Anomaly: H_2O bp = 373 K is *much higher* than H_2S = 213 K.

Cause: H_2O forms extensive **intermolecular hydrogen bonds** (O–H$\cdots$O):

- O has very high EN (3.5), small size, and two lone pairs
- H_2O molecules form a 3D H-bonded network
- H-bonding requires far more energy to break than van der Waals

H_2S: S's EN (2.5) too low for H-bonding $\Rightarrow$ only weak van der Waals $\Rightarrow$ low bp.

[Ans] Final Answer

Option (4) [correct is (3)]: Extensive H-bonding between H_2O molecules (absent in H_2S) raises water's bp anomalously

[!] Common Mistake

Option (4) says "higher molecular mass of H_2O compared to H_2S." H_2O (MW=18) has LOWER mass than H_2S (MW=34). So by van der Waals reasoning, H_2O should have the LOWER bp — making the anomaly (higher actual bp) even more remarkable. The ONLY explanation is hydrogen bonding.

Q.27 Thermal stability trend of Group 16 hydrides?

[A] Conceptual Approach

H–E bond enthalpy data (from the table): H_2O = 463, H_2S = 347, H_2Se = 276, H_2Te = 238 kJ/mol. Decreasing down the group. Weaker bond $\Rightarrow$ decomposes more easily $\Rightarrow$ less thermally stable.

[S] Step-by-Step Solution

H–E bond enthalpy decreases from H_2O (463) to H_2Te (238) kJ mol^{-1}.

Thermal stability $\propto$ H–E bond enthalpy.

$\therefore$ Thermal stability **decreases** from H_2O to H_2Po (and H_2Te/H_2Po are so unstable they decompose readily even at moderate temperatures).

H_2O is the most thermally stable; H_2Po is least stable (decomposes at room temperature essentially).

[Ans] Final Answer

Option (2): Thermal stability decreases $H_2O \rightarrow H_2Po$ due to decreasing H–E bond enthalpy

[!] Common Mistake

Option (1) says thermal stability increases. The H–E bond data directly contradicts this — it decreases consistently from 463 to 238 kJ/mol. Weaker bonds = decomposes more easily = less stable. This is directly readable from the table provided.

Q.28 Dissociation constants: $H_2O = 1.8 \times 10^{-16}$; $H_2Te = 2.3 \times 10^{-3}$. By how many orders of magnitude is H_2Te more acidic than H_2O?

[A] Conceptual Approach

Orders of magnitude = compare the powers of 10. H_2O: 10^{-16}; H_2Te: 10^{-3}. Difference in powers = $(-3) - (-16) = +13$ orders of magnitude.

[S] Step-by-Step Solution

$$\frac{K_a(H_2Te)}{K_a(H_2O)} = \frac{2.3 \times 10^{-3}}{1.8 \times 10^{-16}} \approx 10^{13}$$

H_2Te has a K_a about 10^{13} times larger than H_2O ⇒ **approximately 13 orders of magnitude** more acidic.

(Exact: $\log(2.3 \times 10^{-3}) - \log(1.8 \times 10^{-16}) = (-3 + 0.36) - (-16 + 0.26) = -2.64 - (-15.74) = 13.1$ orders)

[Ans] Final Answer

Option (2): about 13 orders of magnitude

[!] Common Mistake

Adding the exponents instead of subtracting: $(-16) + (-3) = -19$ is wrong. To find the ratio of K_a values, you SUBTRACT exponents: $(-3) - (-16) = +13$. H_2Te's K_a is 10^{13} times LARGER than H_2O's.

Q.29 H–E distances: H_2O=96 pm, H_2S=134 pm, H_2Se=146 pm, H_2Te=169 pm. Trend and relation to bond dissociation enthalpy?

[A] Conceptual Approach

As we go down: E gets larger ⇒ H–E bond gets longer. Longer bond = weaker bond = lower bond dissociation enthalpy. Consistent with data: distance increases, BDE decreases. Classic inverse relationship.

[S] Step-by-Step Solution

H–E bond distance trend: $96 < 134 < 146 < 169$ pm ⇒ **increases** down the group.

Reason: E gets progressively larger (more shells, larger atomic radius). The H–E distance is roughly the sum of H radius + E covalent radius.

Correlation with BDE: Longer bond ⇒ weaker overlap ⇒ lower bond dissociation enthalpy ($463 > 347 > 276 > 238$ kJ mol^{-1}).

Both trends (increasing distance AND decreasing BDE) are consistent: longer bonds are generally weaker bonds.

[Ans] Final Answer

Option (1) [correct is (2)]: H–E distance increases down the group; longer, weaker bonds are consistent with decreasing bond dissociation enthalpy

[!] Common Mistake

Option (1) says "H–E distance decreases, explaining increasing BDE." The data shows distance INCREASES (96 → 169 pm). Decreasing distance would mean the bonds are getting shorter as E gets bigger — physically impossible since E's covalent radius increases going down.

Q.30 Correct INCREASING order of acidic character of Group 16 hydrides?

[A] Conceptual Approach

Acidic character INCREASES going down the group (weaker H–E bond ⇒ easier H^+ release). H_2O is the weakest acid ($K_a = 1.8 \times 10^{-16}$); H_2Te is the strongest ($K_a = 2.3 \times 10^{-3}$). Increasing order = $H_2O < H_2S < H_2Se < H_2Te$.

[S] Step-by-Step Solution

From the dissociation constant data:
K_a: H_2O (1.8×10^{-16}) $\ll$ H_2S (1.3×10^{-7}) $\ll$ H_2Se (1.3×10^{-4}) $\ll$ H_2Te (2.3×10^{-3})

Increasing order of acidity: $H_2O < H_2S < H_2Se < H_2Te$

[Ans] Final Answer

Option (3): $H_2O < H_2S < H_2Se < H_2Te$

[!] Common Mistake

Writing the decreasing order instead of increasing. The question asks for INCREASING acidity. H_2O is the *weakest* acid (smallest K_a, goes first in an increasing sequence); H_2Te is the *strongest* (largest K_a, goes last).

(c) Reactivity with Oxygen

Q.31 Group 16 oxides follow EO_2 and EO_3 formulas. Which elements does "E" refer to, and what is ozone (O_3) classified as here?

[A] Conceptual Approach

E in EO_2/EO_3 = S, Se, Te, Po (they form SO_2, SeO_2, TeO_2; SO_3, SeO_3, TeO_3). Oxygen itself cannot form OO_2 or OO_3 as "oxides" in the usual sense. Ozone (O_3) is an allotrope of oxygen, not an oxide.

[S] Step-by-Step Solution

"E" in EO_2 and EO_3 refers to **S, Se, Te and Po** (oxygen is excluded because oxygen cannot be the central atom in its own oxide in the conventional sense).

Examples:
EO_2: SO_2, SeO_2, TeO_2, PoO_2
EO_3: SO_3, SeO_3, TeO_3

Ozone (O_3): This is an **allotrope of oxygen**, not an EO_2-type oxide. It is merely mentioned as a physical-state reference (a gas), not classified as an EO_2 oxide.

[Ans] Final Answer

Option (1): E = S, Se, Te, Po (oxygen excluded from its own oxide classification); ozone is an allotrope/gas, not an EO_2-type oxide

[!] Common Mistake

Option (4) says ozone is classified as EO_3-type (analogous to SO_3). O_3 is NOT an oxide — it is an allotrope of the pure element oxygen. An oxide requires two different elements (one being oxygen). O_3 is pure oxygen in a different bonding arrangement.

Q.32 Reducing/oxidising behaviour of EO_2 oxides (SO_2 vs TeO_2) going down Group 16?

[A] Conceptual Approach

SO_2 is a well-known reducing agent (used in bleaching). Going down the group, +4 state becomes more stable (inert pair effect). More stable +4 $\Rightarrow$ less tendency to get oxidised further $\Rightarrow$ less reducing, more oxidising. TeO_2 acts as an oxidising agent.

[S] Step-by-Step Solution

Reducing property of EO_2 oxides (E in +4 state):

SO_2: S in +4 state. The +6 state is stable for S $\Rightarrow$ SO_2 is readily oxidised to SO_3/H_2SO_4.
$\therefore$ SO_2 acts as a **reducing agent**.

TeO_2: Te in +4 state. As the inert pair effect increases down the group, Te^{4+} is very stable (reluctant to go to +6). Instead, TeO_2 tends to **accept electrons** (be reduced to Te^{2+} or Te^0).
$\therefore$ TeO_2 acts as an **oxidising agent**.

Reducing property decreases from SO_2 to TeO_2.

[Ans] Final Answer

Option (4): Reducing property decreases from SO_2 to TeO_2; SO_2 = reducing agent; TeO_2 = oxidising agent

[!] Common Mistake

Option (1) reverses the trend. "Reducing property increases from SO_2 to TeO_2" would mean TeO_2 is a stronger reductant than SO_2. Wrong! As +4 becomes more stable going down (inert pair effect), the +4 oxide has less drive to get further oxidised, so it acts less as a reductant and more as an oxidant.

Q.33 Oxidation state of S in SO_3? Does it match the earlier-established common states (+4, +6)?

[A] Conceptual Approach

SO_3: S + 3O = 0. O = -2 each. S + $3(-2)$ = 0. S = +6. Match karo: yes, +6 is one of the common states for S.

[S] Step-by-Step Solution

SO_3 oxidation state calculation:
$x + 3(-2) = 0 \Rightarrow x - 6 = 0 \Rightarrow x = +6$

S in SO_3 = $+6$

Consistency check: Earlier established that S, Se, Te commonly show +4 and +6. $+6$ is indeed one of the two established common states.

This is consistent: SO_3 ($+6$), H_2SO_4 ($+6$), SO_2 ($+4$).

[Ans] Final Answer

Option (2) [correct is (2)]: S is $+6$ in SO_3, consistent with the established common oxidation states

[!] Common Mistake

Confusing EO_2 (+4) with EO_3 (+6). In EO_2: $x+2(-2) = 0 \Rightarrow x = +4$. In EO_3: $x+3(-2) = 0 \Rightarrow x = +6$. The trioxide always corresponds to the +6 state; the dioxide to the +4 state.

Q.34 Which oxide–physical state pair is CORRECTLY matched? SeO_2=solid, O_3=solid, SO_2=solid, SO_3=gas?

[A] Conceptual Approach

Physical states: SO_2 = gas (smells like burnt matches); O_3 = gas; SeO_2 = SOLID (white solid, used in oxidation reactions); SO_3 = liquid or solid (white fumes in moist air, bp = 45°C). At room temp SO_3 is actually a liquid/solid not a gas.

[S] Step-by-Step Solution

Physical states of Group 16 oxides at room temperature (298 K):

SO_2: gas (bp = -10°C = 263 K $<$ 298 K) $\Rightarrow$ gas at RT.
O_3: gas (bp = -112°C = 161 K $<$ 298 K) $\Rightarrow$ gas at RT.
SeO_2: solid (mp = 340°C $>$ 298 K) $\Rightarrow$ solid at RT. ✓
SO_3: liquid/solid at RT (bp = 45°C for α-form, mp = 17°C) $\Rightarrow$ NOT a gas at normal RT.

Correctly matched pair: **SeO_2 — solid**.

[Ans] Final Answer

Option (1): SeO_2 — solid

[!] Common Mistake

Thinking SO_2 is a solid because it is "denser than air." Being denser than air does NOT make SO_2 a solid — SO_2 is definitely a gas at room temperature (you can smell it from volcanoes and burning sulphur). Only SeO_2 is a solid among the options given.

Q.35 Which statement about EO_3-type oxides of Group 16 is CORRECT?

[A] Conceptual Approach

EO_3 = trioxide. NCERT mentions: SO_3, SeO_3, TeO_3 are known. No mention of PoO_3 (polonium's +6 state is not well characterised). So S, Se, Te form EO_3; Po is not mentioned.

[S] Step-by-Step Solution

Known EO_3 oxides:
SO_3: very well known (acid anhydride of H_2SO_4)
SeO_3: known (selenium trioxide)
TeO_3: known (tellurium trioxide)
PoO_3: not mentioned in the text as a characterised EO_3 compound.

Also: O_3 (ozone) is NOT classified as an EO_3 of oxygen.

So option (3) is correct: S, Se, Te form EO_3 oxides; the text does not mention a PoO_3.

[Ans] Final Answer

Option (3): SO_3, SeO_3, TeO_3 are known; the text doesn't mention an EO_3 oxide of polonium

[!] Common Mistake

Option (1) says all five elements form EO_3. Polonium's +6 state is very unstable (inert pair effect in Period 6). NCERT does not list PoO_3 as a characterised compound. Only S, Se, Te clearly form trioxides.

(d) Reaction with Halogens

Q.36 SF_6 is "exceptionally stable for steric reasons." Why, and why don't other hexahalides (like SCl_6) exist?

[A] Conceptual Approach

SF_6 stability = steric argument. S is a relatively small atom. F is the smallest halogen (atomic radius = 64 pm). Six small F atoms can pack snugly in an octahedron around S, **sterically shielding** S from attack. Larger halogens (Cl, Br, I) are too big — they would repel each other and can't all fit in octahedral positions around S.

[S] Step-by-Step Solution

SF_6 exceptional stability:

S (covalent radius $\approx$ 104 pm) is surrounded by 6 F atoms (smallest halogen, van der Waals radius $\approx$ 147 pm) in an **octahedral arrangement**.

F atoms are small enough to pack tightly $\Rightarrow$ they **sterically shield** the S atom at the centre from approach by reagents.
Result: reactions (like hydrolysis) cannot occur because reagents cannot reach the S centre.

Why SCl_6 doesn't exist:
Cl atoms (van der Waals radius $\approx$ 175 pm) are significantly larger than F. Six Cl atoms around S would be **sterically overcrowded** — they would repel each other and no stable structure can form.

[Ans] Final Answer

Option (4): S's small size allows 6 F atoms (smallest halogen) to pack octahedrally, sterically shielding S; larger halogens cannot achieve this close packing

[!] Common Mistake

Option (2) says "S has highest EN in the group, forming strongest S–F bonds." Wrong on two counts: (a) EN decreases down the group, so S does NOT have the highest EN in Group 16 (oxygen does); (b) the stability is specifically described as *steric*, not thermodynamic. S–F bonds are strong, but the *kinetic* stability is due to steric shielding.

Q.37 General stability trend of Group 16 halides (F vs Cl vs Br vs I)?

[A] Conceptual Approach

Standard halide stability order: F > Cl > Br > I. Why? E–F bond is the strongest (F is most electronegative, E–F bond enthalpy highest). E–I bond is weakest. So fluorides are most stable, iodides least stable. Among hexahalides: only hexafluorides (SF_6, SeF_6, TeF_6) are stable.

[S] Step-by-Step Solution

E–X bond strength: F > Cl > Br > I (F most electronegative, forms strongest bonds with E).

$\therefore$ **Stability order** of Group 16 halides (decreasing): F > Cl > Br > I

Hexahalides: Only **hexafluorides** are stable (SF_6, SeF_6, TeF_6). Hexachlorides, hexabromides, hexaiodides of Group 16 do not exist stably because:
(a) Larger halogen atoms create steric strain.
(b) E–Cl, E–Br, E–I bonds weaker $\Rightarrow$ less thermodynamic stabilisation.

[Ans] Final Answer

Option (1): Halide stability decreases F > Cl > Br > I; among hexahalides, only hexafluorides are stable

[!] Common Mistake

Option (2) says stability increases F < Cl < Br < I (iodides most stable). Exact opposite! Fluoride forms the strongest bond with most non-metallic elements (highest EN = strongest charge transfer in bond). Iodide bonds are weakest (I is large, its bonds are long and weak).

Q.38 In $2Se_2Cl_2 \rightarrow SeCl_4 + 3Se$: OS of Se in Se_2Cl_2, and which two OSs does it disproportionate into?

[A] Conceptual Approach

Se_2Cl_2: two Se and two Cl. Cl = -1 each. Let each Se = x. $2x + 2(-1) = 0 \Rightarrow 2x = 2 \Rightarrow x = +1$. Products: $SeCl_4$ (Se = +4) and Se (Se = 0). So Se^{+1} disproportionates into Se^{+4} and Se^{0}.

[S] Step-by-Step Solution

OS of Se in Se_2Cl_2:
Cl = -1; total Cl charge = $2(-1) = -2$
Molecule neutral: $2x + (-2) = 0 \Rightarrow x = +1$
$\therefore$ Se is $+1$ in Se_2Cl_2.

Products:
$SeCl_4$: Se $+4(-1) = 0 \Rightarrow$ Se = $+4$
3Se (elemental): Se = 0

Disproportionation: Se($+1$) simultaneously oxidised to $+4$ (in $SeCl_4$) and reduced to 0 (elemental Se).

[Ans] Final Answer

Option (4): Se is $+1$ in Se_2Cl_2; disproportionates into Se($+4$) in $SeCl_4$ and Se(0) as elemental Se

[!] Common Mistake

Option (1) says Se is $+2$ in Se_2Cl_2. Let's check: Se_2Cl_2 has 2 Se atoms and 2 Cl atoms. If Se = $+2$: charge = $2(+2) + 2(-1) = +4 - 2 = +2 \neq 0$. Not balanced! Must be Se = $+1$: $2(+1) + 2(-1) = 0$. Correct.

Q.39 Match halide type with hybridisation and geometry: EF_4 tetrafluorides, EX_2 dihalides?

[A] Conceptual Approach

EF_4 = central E with 4 F bonds + 1 lone pair = 5 electron groups = sp^3d hybridisation (trigonal bipyramidal arrangement, one equatorial LP = see-saw geometry). EX_2 = 2 bonds + 2 lone pairs = 4 electron groups = sp^3 hybridisation (V-shaped/bent, like H_2O).

[S] Step-by-Step Solution

EF_4 (tetrafluorides): E has 6 valence electrons. 4 used for bonds with F; 1 lone pair remaining.
Total electron groups = 5 $\Rightarrow$ **sp^3d hybridisation**.
Arrangement = trigonal bipyramidal; lone pair in equatorial position $\Rightarrow$ **see-saw geometry**.

EX_2 (dihalides): E has 6 valence electrons. 2 used for bonds; 2 lone pairs.

Total electron groups = 4 ⇒ **sp^3 hybridisation**.
2 bonds + 2 lone pairs ⇒ **V-shaped (bent) geometry**.

[Ans] Final Answer

Option (2): EF_4 = sp^3d (see-saw); EX_2 = sp^3 (bent)

[!] Common Mistake

Option (1) reverses them: "EF_4 = sp^3 tetrahedral." If EF_4 were sp^3, there would be 4 bonds + 0 lone pairs = perfect tetrahedron. But E (Group 16) has 6 valence electrons: with 4 bonds (4 electrons used), 2 remain as 1 lone pair. So 5 electron pairs = sp^3d, NOT sp^3.

Q.40 Physical states of SF_4, SeF_4, TeF_4 at room temperature?

[A] Conceptual Approach

NCERT text states: SF_4 is a gas, SeF_4 is a liquid, TeF_4 is a solid. The pattern: going down (heavier elements), boiling/melting points increase (stronger intermolecular forces), so physical state changes from gas → liquid → solid.

[S] Step-by-Step Solution

From NCERT Table / text on tetrafluorides:

SF_4: Gas at room temperature (MW = 108 g/mol; volatile, low boiling point).
SeF_4: Liquid at room temperature (MW = 155 g/mol; moderate volatility).
TeF_4: Solid at room temperature (MW = 203 g/mol; high intermolecular forces for heavier element).

The trend gas → liquid → solid reflects increasing molecular weight and London dispersion forces down the group.

[Ans] Final Answer

Option (4): SF_4 = gas; SeF_4 = liquid; TeF_4 = solid

[!] Common Mistake

Option (3) says all three tetrafluorides are gases. While SF_4 is a gas, SeF_4 and TeF_4 have higher molecular weights and stronger intermolecular forces (London dispersion forces scale with polarisability and MW), making them liquid and solid respectively.

Q.41 All Group 16 elements form dichlorides and dibromides EXCEPT one. Which exception, and why?

[A] Conceptual Approach

Oxygen forms OF_2 (fluoride only among dihalides) but does NOT form OCl_2, OBr_2, OI_2 as stable compounds under normal conditions. This is because oxygen almost exclusively shows −2 OS, and in EX_2 type dihalides, E is in a positive OS (+2). Oxygen doesn't like positive OS (it's too electronegative). The only "OX_2" = OF_2 (where F is more electronegative, forcing O positive).

[S] Step-by-Step Solution

Exception: Oxygen

In EX_2 dihalides (where X = Cl, Br), the central E atom would need to be in a positive oxidation state (E^{+2}, Cl is -1).

But oxygen's electronegativity (3.5) is higher than Cl (3.0) and Br (2.8).
$\Rightarrow$ O refuses to take a positive OS with Cl or Br (it would be the more electronegative partner, getting the negative charge).

OF_2 exists because F (EN = 4.0) $>$ O (EN = 3.5) — F forces O into +2.
But OCl_2-type structures are unstable under normal conditions because O $>$ Cl in EN.

This is consistent with oxygen showing almost exclusively -2 (except with F).

[Ans] Final Answer

Option (1): Oxygen is the exception; consistent with O showing only -2 OS (except OF_2 where F is more electronegative)

[!] Common Mistake

Option (2) says polonium is the exception because it's radioactive. Radioactivity doesn't prevent halide formation. Polonium does form $PoCl_2$, $PoBr_2$ etc. The exception is OXYGEN — which uniquely refuses positive oxidation states with halogens less electronegative than F.

Q.42 Which is NOT a dimeric monohalide mentioned in the text? S_2F_2, S_2Cl_2, Te_2Cl_2, Se_2Br_2?

[A] Conceptual Approach

NCERT text mentions these dimeric monohalides: S_2F_2, S_2Cl_2, Se_2Cl_2, Se_2Br_2. Check: Te_2Cl_2 — this is NOT in the NCERT list. NCERT specifically mentions dimeric monohalides of S and Se, not Te.

[S] Step-by-Step Solution

Dimeric monohalides (E_2X_2) mentioned in NCERT text:
- S_2F_2 (disulphur difluoride)
- S_2Cl_2 (disulphur dichloride)
- Se_2Cl_2 (diselenium dichloride)
- Se_2Br_2 (diselenium dibromide)

Te_2Cl_2 is **NOT** mentioned in NCERT's list of dimeric monohalides for Group 16 elements.

[Ans] Final Answer

Option (3): Te_2Cl_2 is NOT one of the dimeric monohalides mentioned in the text

[!] Common Mistake

Confusing Se_2Cl_2 with Te_2Cl_2. The text specifically lists S_2F_2, S_2Cl_2, Se_2Cl_2, Se_2Br_2 as dimeric monohalides. Tellurium's dimeric monohalides are not listed.

Q.43 Verify balance of $2Se_2Cl_2 \rightarrow SeCl_4 + 3Se$ by counting atoms.

[A] Conceptual Approach

Count Se and Cl on both sides. LHS: $2\times Se_2Cl_2$ = 4 Se + 4 Cl. RHS: $SeCl_4$ + 3Se = 1 Se + 4 Se = 4 Se (total); 4 Cl. So both balanced.

[S] Step-by-Step Solution

LHS: $2Se_2Cl_2 = 2 \times 2 = 4$ Se atoms; $2 \times 2 = 4$ Cl atoms.

RHS: $SeCl_4$ (1 Se + 4 Cl) + 3Se (3 Se + 0 Cl) = $1 + 3 = 4$ Se; $4 + 0 = 4$ Cl.

LHS: 4 Se, 4 Cl = RHS: 4 Se, 4 Cl ✓

The equation is **correctly balanced**.

[Ans] Final Answer

Option (4): LHS: 4 Se, 4 Cl = RHS: 4 Se, 4 Cl — correctly balanced

[!] Common Mistake

Forgetting to multiply both atoms by the coefficient 2 on the LHS. $2Se_2Cl_2$ means TWO molecules of Se_2Cl_2: so $2 \times 2 = 4$ Se and $2 \times 2 = 4$ Cl. Option (3) only counts 1 molecule (4 Se, 4 Cl on LHS — wait, option 3 says 4 Se but only 4 Cl but then 3 Se on RHS... Always redo the count step by step).

Q.44 Structure and physical state of Group 16 hexafluorides?

[A] Conceptual Approach

Hexafluorides: EF_6 = 6 bonds around E + 0 lone pairs = 6 electron groups = sp^3d^2 hybridisation = octahedral geometry. All are gases at room temperature (SF_6, SeF_6, TeF_6 are all gaseous).

[S] Step-by-Step Solution

EF_6 (hexafluorides): E with 6 F bonds, 0 lone pairs.
6 electron groups $\Rightarrow$ sp^3d^2 hybridisation $\Rightarrow$ **Octahedral geometry**.

Physical states: SF_6, SeF_6, TeF_6 are all **gases** at room temperature.
(SF_6 bp = $-64°C$; SeF_6 bp = $-34.5°C$; TeF_6 bp = $-36°C$ — all well below 25°C).

[Ans] Final Answer

Option (2): Octahedral structure; all gaseous at room temperature

[!] Common Mistake

Option (1) says "tetrahedral, all liquids." Tetrahedral = 4 bonds. Hexafluorides have 6 bonds. 6 bonds = octahedral. And physically all three hexafluorides are gases at room temperature — SF_6 is the "heavy gas" used in electrical insulation (but still a gas at RT!).

TOPIC 5 — Anomalous Behaviour of Oxygen

Anomalous Behaviour of Oxygen

Q.45 Which correctly ranks halide stability in DECREASING order?

[A] Conceptual Approach

Standard: F > Cl > Br > I in decreasing stability. Fluorides most stable; iodides least. This applies to Group 16 halides as well.

[S] Step-by-Step Solution

Bond strength: E–F > E–Cl > E–Br > E–I

∴ Stability in decreasing order: **F > Cl > Br > I**

[Ans] Final Answer

Option (3): F > Cl > Br > I

[!] Common Mistake

Option (1) says I > Br > Cl > F. This is the order of bond LENGTH (longer bonds with bigger halogens), but NOT stability. Longer bonds are weaker bonds in this context. Stability = bond strength = F > Cl > Br > I.

Q.46 Anomalous behaviour of O compared to other Group 16: two structural factors and specific example?

[A] Conceptual Approach

Oxygen is anomalous because of: (1) small size and (2) high electronegativity. Together these enable **hydrogen bonding** in H_2O — which is absent in H_2S and heavier hydrides. This is why H_2O has anomalously high bp compared to H_2S.

[S] Step-by-Step Solution

Two factors causing oxygen's anomalous behaviour:

(1) **Small size**: O has only n=2 electrons; very compact atom. No intervening d/f electrons.
(2) **High electronegativity** (3.5; second only to F): O creates highly polarised O–H bonds.

Together: O is small + highly electronegative ⇒ **strong hydrogen bonding** in H_2O.
H-bonds are intermolecular O–H$\cdots$O interactions (absent in H_2S because S is larger and less electronegative).

Specific example: H_2O bp = 373 K (anomalously high) vs H_2S bp = 213 K.

[Ans] Final Answer

Option (2): Small size + high electronegativity ⇒ strong H-bonding in H_2O; absent in H_2S

[!] Common Mistake

Option (4) says "presence of d-orbitals and high electronegativity." OXYGEN HAS NO D-ORBITALS (Period 2). This is a defining feature of oxygen's anomalous chemistry. The very *absence* of d-orbitals (not their presence) limits O's covalency. The two correct factors are small size and high EN.

Q.47 Covalency of oxygen vs other Group 16 elements?

[A] Conceptual Approach

O = Period 2, no d-orbitals. Max covalency = 4 theoretically, but in practice rarely exceeds 2. Other Group 16 (S, Se, Te) have d-orbitals ⇒ can expand valence shell ⇒ covalency can exceed 4 (like SF_6 = covalency 6, sp^3d^2).

[S] Step-by-Step Solution

Oxygen covalency:
O = Period 2: only 2s and 2p orbitals (no d-orbitals).
Maximum theoretical covalency = 4 (using sp^3, like in H_3O^+).
In practice, O rarely exceeds covalency **2** (like in H_2O, CO_2 etc.).

Other Group 16 (S, Se, Te) covalency:
S, Se, Te have accessible d-orbitals (3d, 4d, 5d).
Can use sp^3d (covalency 5) or sp^3d^2 (covalency 6).
Example: SF_6 (S covalency = 6, sp^3d^2).

[Ans] Final Answer

Option (1): O's covalency limited to max 4 (no d-orbitals) and in practice rarely exceeds 2; S, Se, Te can exceed 4 using d-orbitals

[!] Common Mistake

Option (2) says oxygen's covalency is "unlimited due to accessible d-orbitals." OXYGEN HAS NO D-ORBITALS. This is one of the most important facts about oxygen's anomalous chemistry. Never say oxygen has d-orbitals.

Q.48 Which scenario is IMPOSSIBLE for oxygen due to its covalency limitation?

[A] Conceptual Approach

Max covalency of O = 4. OF_6 would require O to have covalency 6 (bond with 6 F atoms). This would need sp^3d^2 hybridisation (requires d-orbitals). O has no d-orbitals ⇒ impossible.

[S] Step-by-Step Solution

Check each option:

(1) H_2O: O covalency = 2 (2 O–H bonds). Well within limit. Possible ✓
(2) H_3O^+: O covalency = 3 (3 bonds: 2 O–H + 1 coordinate bond to H). sp^3, within limit. Possible ✓
(3) **OF_6**: O would need **covalency = 6** (6 O–F bonds). This needs sp^3d^2 hybridisation requiring d-orbitals. O has **NO d-orbitals** ⇒ **IMPOSSIBLE**.
(4) OF_2: O covalency = 2 (2 O–F bonds). Well within limit (and it actually exists!). Possible ✓

[Ans] Final Answer

Option (3): OF_6 is impossible — O would need covalency 6, requiring sp^3d^2 (d-orbitals needed); O has no d-orbitals; maximum covalency for O is 4

[!] Common Mistake

Option (2) says H_3O^+ is impossible for O ("covalency of 3 exceeds its maximum"). H_3O^+ (the hydronium ion) is completely real and very important in acid-base chemistry. O can reach covalency 3 or even 4 (sp^3); it just cannot go BEYOND 4. The impossible one is OF_6 (covalency 6).

Q.49 Why can sulphur (Period 3) expand valence shell beyond 4 while oxygen (Period 2) cannot?

[A] Conceptual Approach

Ek hi answer: S ke paas 3d orbitals hain (Period 3 mein d-orbital hai); O ke paas koi bhi d-orbital nahi (Period 2 mein sirf s aur p). Expanding covalence beyond 4 requires d-orbitals. Same logic applies throughout p-block Period 2 vs Period 3+ comparisons.

[S] Step-by-Step Solution

Oxygen (Period 2): Valence shell = 2s + 2p orbitals **only**.
No d-orbitals in Period 2 $\Rightarrow$ cannot use sp^3d or sp^3d^2 hybridisation.
Maximum 4 bonds possible (sp^3).

Sulphur (Period 3): Valence shell = 3s + 3p + **3d** orbitals.
3d orbitals are energetically accessible for bonding.
Can use sp^3d (5 bonds) or sp^3d^2 (6 bonds) hybridisation.
Examples: SF_4 (sp^3d, 5 electron pairs), SF_6 (sp^3d^2, 6 bonds).

[Ans] Final Answer

Option (4) [correct is (1)]: S has accessible 3d orbitals (Period 3) allowing valence shell expansion; O's n=2 shell has no d-orbitals at all

[!] Common Mistake

Option (2) says "atomic radius alone determines maximum covalency." If that were true, O (smallest radius) would have the lowest covalency, and Po (largest) would have the highest — but maximum covalency is fundamentally about ORBITAL AVAILABILITY, not just size. The d-orbital presence/absence is the key factor.

Q.50 Where is strong hydrogen bonding present among Group 16 hydrides?

[A] Conceptual Approach

H-bonding needs: donor atom with high EN (N, O, F) bonded to H, AND acceptor atom with lone pair. O (EN = 3.5) = perfect. S (EN = 2.5) = not high enough for effective H-bonding. H_2O has H-bonding; H_2S does not (or only very weakly).

[S] Step-by-Step Solution

Conditions for effective hydrogen bonding: the atom bonded to H must be:

- Highly electronegative (creates highly polarised X–H bond)
- Small (allows close approach of H-bond donor and acceptor)

H_2O: O (EN = 3.5, very small) $\Rightarrow$ strong O–H$\cdots$O H-bonds. **Present.**

H_2S: S (EN = 2.5, much larger) $\Rightarrow$ S–H is only weakly polarised; S is too large.
H-bonding essentially **absent** in H_2S (only very weak van der Waals).

[Ans] Final Answer

Option (4): H-bonding present in H_2O; absent in H_2S

[!] Common Mistake

Option (2) says H-bonding present in both H_2O and H_2S equally. Wrong! S (EN = 2.5) is not electronegative enough to form effective hydrogen bonds. The benchmark for H-bonding is EN $\geq$ 3.0 approximately (N, O, F). S at 2.5 falls below this threshold.

DPP-9 Complete!

"Group 16 was never 50 separate facts.

It was always three connected ideas:

Oxygen is anomalous (small, no d-orbitals, H-bonding),
the E–H bond weakens down the group (acidity, stability, reducing power),
and the inert pair effect makes +6 less stable at the bottom.
You now see all three threads."

— Your Weird Chemist