

# IIT JAM 2026 Chemistry Question Paper with Solutions

Time Allowed :3 Hour	Maximum Marks :100	Total Questions :65
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## General Instructions

Please read the following instructions carefully:

1. The examination is of 3 hours duration. There are a total of 60 questions carrying 100 marks. The entire paper is divided into three sections, A, B and C. All sections are compulsory. Questions in each section are of different types.
2. Section A contains a total of 30 Multiple Choice Questions (MCQ). Each MCQ type question has four choices out of which only one choice is the correct answer. Questions Q.1 – Q.30 belong to this section and carry a total of 50 marks. Q.1 – Q.10 carry 1 mark each and Questions Q.11 – Q.30 carry 2 marks each.
3. Section B contains a total of 10 Multiple Select Questions (MSQ). Each MSQ type question is similar to MCQ but with a difference that there will be more than one choices that are correct out of the four given choices. The candidate gets full credit if he/she selects all the correct answers only and no wrong answers. Questions Q.31 – Q.40 belong to this section and carry 2 marks each with a total of 20 marks.
4. Section C contains a total of 20 Numerical Answer Type (NAT) questions. For these NAT type questions, the answer is a real number which needs to be entered using the virtual keyboard on the monitor. No choices will be shown for these type of questions. Questions Q.41 – Q.60 belong to this section and carry a total of 30 marks. Q.41 – Q.50 carry 1 mark each and Questions Q.51 – Q.60 carry 2 marks each.
5. In all sections, questions not attempted will result in zero mark. In Section A (MCQ), wrong answer will result in NEGATIVE marks. For all 1-mark questions, 1/3 marks will be deducted for each wrong answer. For all 2-mark questions, 2/3 marks will be deducted for each wrong answer. In Section B (MSQ), there is NO NEGATIVE and NO PARTIAL marking provisions. There is NO NEGATIVE marking in Section C (NAT) as well.
6. Only Virtual Scientific Calculator is allowed. Charts, graph sheets, tables, cellular phone or other electronic gadgets are NOT allowed in the examination hall.
7. A Scribble Pad will be provided for rough work.

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1. Find the number of P–O–P (pyrophosphate) bonds present in  $P_4O_8$ .

**Solution:**

**Step 1: Understand P–O–P bond.**

A P–O–P bond is a bridging oxygen atom connecting two phosphorus atoms. Such bonds are also called pyrophosphate linkages.

**Step 2: Use oxidation state method.**

In  $P_4O_8$ , assume oxidation state of phosphorus =  $x$ .  
Since oxygen has oxidation state  $-2$ , we write:

$$4x + 8(-2) = 0.$$

$$4x - 16 = 0.$$

$$4x = 16.$$

$$x = +4.$$

**Step 3: Count terminal P=O bonds.**

Phosphorus in  $+4$  oxidation state generally forms:  
One P=O double bond and three P–O single bonds.

Total oxygen atoms = 8.

If each phosphorus forms one terminal P=O bond, total terminal oxygens = 4.

Remaining oxygens =

$$8 - 4 = 4.$$

**Step 4: Determine P–O–P bridges.**

Each bridging oxygen connects two phosphorus atoms.  
Thus number of P–O–P bonds equals number of bridging oxygens.

Hence, number of P–O–P bonds =

$$\boxed{4}.$$

### Quick Tip

Total oxygen atoms minus terminal P=O oxygens gives number of bridging oxygens. Each bridging oxygen corresponds to one P–O–P bond.

## 2. Determine the shape of $ClF_3$ using VSEPR theory.

### Solution:

#### Step 1: Find total valence electrons.

Chlorine has 7 valence electrons.

Each fluorine has 7 valence electrons.

Total valence electrons:

$$= 7 + 3(7) = 28.$$

#### Step 2: Determine number of electron pairs around central atom.

Chlorine forms three bonds with fluorine atoms.

Thus, 3 bonding pairs are present.

Remaining electrons on chlorine:

$$7 - 3 = 4 \text{ electrons}$$

which form 2 lone pairs.

Hence total electron pairs around chlorine:

$$3 \text{ bonding pairs} + 2 \text{ lone pairs} = 5.$$

#### Step 3: Determine electron geometry.

Five electron pairs correspond to:

Trigonal Bipyramidal electron geometry.

#### Step 4: Arrange lone pairs.

According to VSEPR, lone pairs occupy equatorial positions to minimize repulsion.

Two equatorial positions are occupied by lone pairs.

Remaining three positions are occupied by fluorine atoms.

**Step 5: Molecular shape.**

After placing lone pairs, the molecular geometry becomes:

T-shaped.

**Quick Tip**

For 5 electron pairs (trigonal bipyramidal arrangement), if there are 2 lone pairs, the molecular shape becomes T-shaped.

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**3. List the compounds of Xenon having one lone pair on the central atom.**

**Solution:**

**Step 1: Understand valence electrons of Xenon.**

Xenon (Xe) has 8 valence electrons.

While forming compounds, some electrons are used in bonding and remaining electrons form lone pairs.

**Step 2: Check common Xenon compounds.**

**(i)  $XeF_2$**

Electron pairs around Xe = 5 (2 bonding + 3 lone pairs).

Hence, 3 lone pairs are present.

**(ii)  $XeF_4$**

Electron pairs around Xe = 6 (4 bonding + 2 lone pairs).

Hence, 2 lone pairs are present.

**(iii)  $XeF_6$**

Electron pairs around Xe = 7 (6 bonding + 1 lone pair).

Hence, 1 lone pair is present.

**(iv)  $XeO_3$**

Total electron domains = 4 (3 bonding + 1 lone pair).

Hence, 1 lone pair is present.

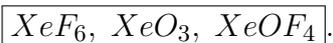
**(v)  $XeOF_4$**

Total electron domains = 6 (5 bonding + 1 lone pair).

Hence, 1 lone pair is present.

**Step 3: Final Answer.**

Compounds of Xenon having **one lone pair** are:



**Quick Tip**

To find lone pairs on Xenon, subtract number of bonding pairs from 4 (since Xe has 8 valence electrons = 4 pairs).

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**4. Write the van der Waals equation and explain the van der Waals isotherm (P–V curve).**

**Solution:**

**Step 1: Van der Waals Equation.**

The ideal gas equation is:

$$PV = nRT.$$

Van der Waals modified it to account for:

1. Finite volume of gas molecules
2. Intermolecular attraction

The van der Waals equation is:

$$\left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT.$$

Where:

$a$  = measure of intermolecular attraction

$b$  = excluded volume of gas molecules

**Step 2: Explanation of Corrections.**

**Pressure correction:**

$$P_{real} = P_{ideal} - \frac{an^2}{V^2}.$$

The term  $\frac{an^2}{V^2}$  corrects for intermolecular attraction.

**Volume correction:**

$$V_{real} = V_{container} - nb.$$

The term  $nb$  accounts for finite molecular volume.

### **Step 3: Van der Waals Isotherm (P–V Curve).**

When plotting pressure vs volume at constant temperature:

1. At high temperature ( $T > T_c$ ):  
Curve resembles ideal gas behaviour (smooth decrease).
2. At critical temperature ( $T = T_c$ ):  
Curve shows a point of inflection (critical point).
3. At low temperature ( $T < T_c$ ):  
Curve shows a characteristic S-shaped region.  
This region represents liquid–gas phase transition.

The flat portion after Maxwell correction represents coexistence of liquid and vapour phases.

### **Step 4: Critical Constants.**

At critical point:

$$\left(\frac{\partial P}{\partial V}\right)_T = 0$$
$$\left(\frac{\partial^2 P}{\partial V^2}\right)_T = 0.$$

Critical constants are:

$$V_c = 3nb,$$
$$P_c = \frac{a}{27b^2},$$
$$T_c = \frac{8a}{27Rb}.$$

**Final Answer:**

The van der Waals equation is:

$$\left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT$$

and its P–V curve shows deviation from ideal behaviour with a characteristic S-shaped isotherm below critical temperature.

**Quick Tip**

The S-shaped region in van der Waals isotherm indicates phase transition. At the critical point, the curve has an inflection point.

**5. Langmuir isotherm**

- (A) Multilayer adsorption
- (B) Very high temperature  $\theta$  vs  $P$  curve with intercept
- (C) Very low temperature  $\theta$  vs  $P$  curve with slope
- (D) Reciprocal form gives  $\theta = \frac{KP}{1 + KP}$

**Correct Answer:** (D) Reciprocal form gives  $\theta = \frac{KP}{1 + KP}$

**Solution:****Step 1: Understand Langmuir adsorption isotherm.**

Langmuir isotherm assumes monolayer adsorption on a homogeneous surface.

All adsorption sites are identical and there is no interaction between adsorbed molecules.

**Step 2: Mathematical expression.**

The Langmuir isotherm equation is:

$$\theta = \frac{KP}{1 + KP}$$

where  $\theta$  is fractional surface coverage,  $K$  is adsorption equilibrium constant, and  $P$  is pressure.

**Step 3: Evaluation of options.**

- (A) Incorrect — Langmuir theory is limited to monolayer adsorption.
- (B) and (C) Not defining statements of Langmuir isotherm.
- (D) Correct — It gives the exact mathematical form of Langmuir isotherm.

**Step 4: Conclusion.**

Hence, option (D) is correct.

### Quick Tip

Langmuir isotherm always represents monolayer adsorption and follows  $\theta = \frac{KP}{1 + KP}$ .

**6. Find the interplanar spacing  $d_{hkl}$  (in Å) using Bragg's law.**

**Given:**  $n = 1$ ,  $\lambda = 1.54 \text{ Å}$

**Use Bragg's law:**  $n\lambda = 2d \sin \theta$ .

**Solution:**

**Step 1: Write Bragg's Law.**

Bragg's law is given by:

$$n\lambda = 2d \sin \theta.$$

We are given:

$$n = 1, \quad \lambda = 1.54 \text{ Å}.$$

**Step 2: Rearrange for  $d$ .**

$$d = \frac{n\lambda}{2 \sin \theta}.$$

Substituting  $n = 1$ :

$$d = \frac{1.54}{2 \sin \theta}.$$

**Step 3: Final Expression.**

Thus, interplanar spacing is:

$$d = \frac{1.54}{2 \sin \theta} \text{ Å}.$$

(If  $\theta$  is given, substitute directly to compute numerical value.)

### Quick Tip

In X-ray diffraction problems, always remember Bragg's law:  $n\lambda = 2d \sin \theta$ . If first order reflection is given ( $n = 1$ ), the formula simplifies significantly.

**7. For the reaction:**



If the degree of dissociation is  $\alpha = 0.2$ , find  $\Delta n_g$  and total number of moles at equilibrium (starting with 1 mole).

**Solution:**

**Step 1: Find  $\Delta n_g$ .**

Number of gaseous moles on product side = 2

Number of gaseous moles on reactant side = 1

$$\Delta n_g = 2 - 1 = 1.$$

**Step 2: Initial moles.**

Assume 1 mole of  $N_2O_3$  initially.

Initial moles:

$$N_2O_3 = 1, \quad NO = 0, \quad NO_2 = 0.$$

**Step 3: Change in moles.**

If degree of dissociation =  $\alpha = 0.2$ , then moles dissociated:

$$= 1 \times 0.2 = 0.2.$$

Thus:

$$N_2O_3 \text{ left} = 1 - 0.2 = 0.8.$$

$$NO = 0.2.$$

$$NO_2 = 0.2.$$

**Step 4: Total moles at equilibrium.**

$$n_{total} = 0.8 + 0.2 + 0.2 = 1.2.$$

**Final Answers:**

$$\Delta n_g = 1$$

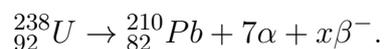
$$n_{total} = 1.2$$

#### Quick Tip

For dissociation reactions, total moles increase if  $\Delta n_g > 0$ . Always use: initial moles  $\rightarrow$  change using  $\alpha \rightarrow$  equilibrium moles.

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**8. The radioactive decay is given as:**



**Find the value of  $x$ .**

**Solution:**

**Step 1: Understand the decay process.**

Each  $\alpha$ -particle emission reduces:

Mass number by 4

Atomic number by 2

Each  $\beta^-$  emission increases:

Atomic number by 1

Mass number remains unchanged

**Step 2: Apply mass number conservation.**

Initial mass number = 238.

After 7  $\alpha$  emissions:

$$238 - 7(4) = 238 - 28 = 210.$$

Final nucleus has mass number 210, which matches  $^{210}\text{Pb}$ .  
Thus mass balance is satisfied.

**Step 3: Apply atomic number conservation.**

Initial atomic number = 92.

After 7  $\alpha$  emissions:

$$92 - 7(2) = 92 - 14 = 78.$$

Let number of  $\beta^-$  emissions be  $x$ .

Each  $\beta^-$  increases atomic number by 1.

Final atomic number becomes:

$$78 + x.$$

Given final atomic number of Pb = 82.

$$78 + x = 82.$$

$$x = 4.$$

**Final Answer:**

$$\boxed{x = 4}.$$

**Quick Tip**

Remember:  $\alpha$  emission decreases atomic number by 2.  $\beta^-$  emission increases atomic number by 1. Always balance both mass number and atomic number.

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**9. Among the following compounds of Xenon, identify those having exactly one lone pair on Xenon:**

$XeF_2$ ,  $XeO_2F_2$ ,  $XeO_3$ ,  $XeO_3F_2$ ,  $XeOF_2$ .

**Solution:**

**Step 1: Basic idea.**

Xenon has 8 valence electrons = 4 electron pairs.

Number of lone pairs on Xe = 4 (number of bonds formed by Xe).

(Each  $\sigma$  bond counts as one bonding pair.)

**Step 2: Analyze each compound.**

**(i)  $XeF_2$**

Number of  $\sigma$  bonds = 2.

Lone pairs on Xe =  $4 - 2 = 2$  pairs left? Actually total electron domains = 5 ( $AX_2E_3$ ).

Thus, 3 lone pairs are present.

**(ii)  $XeO_2F_2$**

Total  $\sigma$  bonds = 4 (2 Xe=O and 2 Xe-F; each contributes one  $\sigma$  bond).

Thus bonding pairs = 4.

Remaining lone pairs =  $4 - 4 = 0$ ?

But electron domain counting gives  $AX_4E_1$ .

Hence, 1 lone pair is present.

**(iii)  $XeO_3$**

Three Xe=O bonds  $\rightarrow$  3  $\sigma$  bonds.

Thus bonding pairs = 3.

Remaining lone pairs =  $4 - 3 = 1$ .

Hence, 1 lone pair is present.

**(iv)  $XeO_3F_2$**

Total  $\sigma$  bonds = 5.

Bonding pairs = 5.

Remaining lone pairs =  $4 - 5$  not possible  $\rightarrow$  actually  $AX_5E_1$ .

Hence, 1 lone pair present.

**(v)  $XeOF_2$**

Total  $\sigma$  bonds = 3.

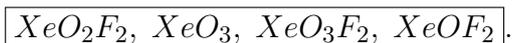
Thus bonding pairs = 3.

Remaining lone pairs =  $4 - 3 = 1$ .

Hence, 1 lone pair is present.

**Step 3: Final Answer.**

Compounds having exactly one lone pair on Xe are:

**Quick Tip**

For Xenon compounds, count total  $\sigma$  bonds first. Number of lone pairs = 4 (number of  $\sigma$  bonds). Then verify using VSEPR notation  $\text{AX}_m\text{E}_n$ .

**10. Isothermal compressibility of an ideal gas**

$$-\frac{1}{V} \left( \frac{dV}{dP} \right)_T$$

is equal to

- (A)  $\frac{nR}{P}$
- (B)  $nR$
- (C)  $R$
- (D)  $\frac{1}{P}$

**Correct Answer:** (D)  $\frac{1}{P}$

**Solution:**

**Step 1: Write the ideal gas equation.**

For an ideal gas,

$$PV = nRT$$

At constant temperature  $T$ ,

$$V = \frac{nRT}{P}$$

**Step 2: Differentiate with respect to pressure.**

$$\frac{dV}{dP} = -\frac{nRT}{P^2}$$

**Step 3: Substitute in compressibility expression.**

Isothermal compressibility is defined as:

$$\kappa_T = -\frac{1}{V} \left( \frac{dV}{dP} \right)_T$$

Substituting values:

$$\begin{aligned}\kappa_T &= -\frac{1}{\frac{nRT}{P}} \left( -\frac{nRT}{P^2} \right) \\ \kappa_T &= \frac{P}{nRT} \cdot \frac{nRT}{P^2} \\ \kappa_T &= \frac{1}{P}\end{aligned}$$

**Step 4: Conclusion.**

Hence, the isothermal compressibility of an ideal gas is  $\frac{1}{P}$ .

**Quick Tip**

For an ideal gas, always remember: Isothermal compressibility  $\kappa_T = \frac{1}{P}$ . It depends only on pressure and not on temperature.

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**11. Which of the following has the highest melting point?**

- (A) LiBr
- (B) NaBr
- (C) MgBr<sub>2</sub>
- (D) AlBr<sub>3</sub>

**Correct Answer:** (B) NaBr

**Solution:****Step 1: Understand factors affecting melting point of ionic compounds.**

Melting point mainly depends on lattice energy.

Higher lattice energy → stronger ionic bonding → higher melting point.

Lattice energy depends on:

- Charge on ions
- Size of ions (smaller size → stronger attraction)

**Step 2: Compare the given compounds.**

(A)  $\text{LiBr} \rightarrow \text{Li}^+$  is very small and highly polarizing. It increases covalent character (Fajan's rule), lowering melting point.

(B)  $\text{NaBr} \rightarrow \text{Na}^+$  is larger than  $\text{Li}^+$  and less polarizing. Hence, bonding is more ionic and lattice energy remains high.

(C)  $\text{MgBr}_2 \rightarrow$  Though  $\text{Mg}^{2+}$  has higher charge, it strongly polarizes  $\text{Br}^-$ , increasing covalent character and reducing melting point.

(D)  $\text{AlBr}_3 \rightarrow$  Highly covalent in nature and exists as dimer ( $\text{Al}_2\text{Br}_6$ ), so it has low melting point.

**Step 3: Conclusion.**

Among the given options,  $\text{NaBr}$  is the most ionic and thus has the highest melting point.

**Quick Tip**

Use Fajan's rule: Small highly charged cations increase covalent character, which lowers melting point. More ionic character  $\rightarrow$  higher melting point.

**12. Compare the M–C bond length in the following complexes:**

$\text{Na}_2[\text{Mn}(\text{CO})_4]$ ,  $\text{Na}[\text{Mn}(\text{CO})_5]$ ,  $\text{Na}[\text{Mn}(\text{CO})_6]$ .

**Solution:****Step 1: Understand metal–CO bonding.**

In metal carbonyl complexes, bonding involves:

1.  $\sigma$ -donation from CO to metal.
2.  $\pi$ -back bonding from metal to CO.

Greater  $\pi$ -back bonding strengthens M–C bond and shortens M–C bond length.

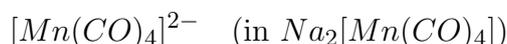
**Step 2: Effect of oxidation state.**

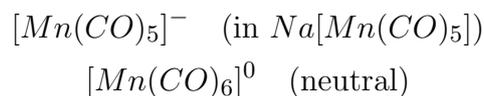
More negative charge on complex  $\rightarrow$  higher electron density on metal.

Higher electron density  $\rightarrow$  stronger  $\pi$ -back bonding.

Stronger back bonding  $\rightarrow$  shorter M–C bond.

Now compare charges:





Electron density order on metal:

$$2- > 1- > 0.$$

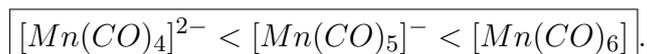
Thus back bonding strength:



**Step 3: Final order of M–C bond length.**

Stronger back bonding → shorter bond.

Hence bond length order (shortest to longest):



**Or in decreasing bond length:**



#### Quick Tip

More negative charge on metal carbonyl complex increases  $\pi$ -back bonding, which shortens the M–C bond.

**13. Given**  $C_p = a + bT$

$a = 19.5$ ,  $b = 0.042$ ,  $n = 1$  mole

**Temperature changes from**  $27^\circ C$  **to**  $327^\circ C$ .

**Find**  $\Delta H$ .

**Solution:**

**Step 1: Convert temperature to Kelvin.**

$$T_1 = 27^\circ C = 300 K$$

$$T_2 = 327^\circ C = 600 K$$

**Step 2: Use formula for enthalpy change.**

$$\Delta H = \int_{T_1}^{T_2} C_p dT$$

$$= \int_{T_1}^{T_2} (a + bT) dT$$

**Step 3: Integrate.**

$$\Delta H = a(T_2 - T_1) + \frac{b}{2}(T_2^2 - T_1^2)$$

**Step 4: Substitute values.**

$$T_2 - T_1 = 600 - 300 = 300$$

$$T_2^2 - T_1^2 = 600^2 - 300^2$$

$$= 360000 - 90000$$

$$= 270000$$

Now calculate:

$$\Delta H = 19.5(300) + \frac{0.042}{2}(270000)$$

$$= 5850 + (0.021)(270000)$$

$$= 5850 + 5670$$

$$= 11520$$

**Final Answer:**

$$\boxed{\Delta H = 11520}$$

(Units depend on given  $C_p$ , typically J/mol.)

#### Quick Tip

If  $C_p$  is temperature dependent, always integrate:  $\Delta H = \int C_p dT$ . Remember to convert temperature to Kelvin.

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**14. T-shape geometry is shown by**

- (A) XeF<sub>2</sub>
- (B) SO<sub>2</sub>
- (C) IO<sub>4</sub><sup>-</sup>
- (D) BrF<sub>3</sub>

**Correct Answer:** (D) BrF<sub>3</sub>

**Solution:**

**Step 1: Use VSEPR theory.**

T-shape geometry corresponds to AX<sub>3</sub>E<sub>2</sub> type molecule.

This means 5 electron pairs (trigonal bipyramidal arrangement) with 2 lone pairs occupying equatorial positions.

**Step 2: Analyze each option.**

(A) XeF<sub>2</sub> → AX<sub>2</sub>E<sub>3</sub> → Linear geometry.

(B) SO<sub>2</sub> → AX<sub>2</sub>E → Bent geometry.

(C) IO<sub>4</sub><sup>-</sup> → AX<sub>4</sub> → Tetrahedral geometry.

(D) BrF<sub>3</sub> → AX<sub>3</sub>E<sub>2</sub> → Trigonal bipyramidal electron geometry with 2 lone pairs → T-shaped molecular geometry.

**Step 3: Conclusion.**

Therefore,  $\text{BrF}_3$  shows T-shaped geometry.

**Quick Tip**

T-shape geometry arises from  $\text{AX}_3\text{E}_2$  type molecules. Remember: 5 electron pairs  $\rightarrow$  2 lone pairs in equatorial positions  $\rightarrow$  T-shape.

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**15. Which of the following has tetrahedral coordination?**

- (A)  $\text{CdI}_2$
- (B)  $\text{CsCl}$
- (C)  $\text{ZnS}$  (sphalerite)
- (D)  $\text{ZnS}$  (wurtzite)

**Correct Answer:** (C)  $\text{ZnS}$  (sphalerite), (D)  $\text{ZnS}$  (wurtzite)

**Solution:****Step 1: Understand tetrahedral coordination.**

In tetrahedral coordination, each cation is surrounded by 4 anions arranged at the corners of a tetrahedron.

Coordination number = 4.

**Step 2: Analyze each compound.**

(A)  $\text{CdI}_2 \rightarrow$  Layered structure  $\rightarrow$  Octahedral coordination (Coordination number = 6).

(B)  $\text{CsCl} \rightarrow$  Cubic structure  $\rightarrow$  Coordination number = 8 (body-centered cubic type).

(C)  $\text{ZnS}$  (sphalerite)  $\rightarrow$  Zinc blende structure  $\rightarrow$  Each  $\text{Zn}^{2+}$  surrounded by 4  $\text{S}^{2-}$   $\rightarrow$  Tetrahedral coordination.

(D)  $\text{ZnS}$  (wurtzite)  $\rightarrow$  Hexagonal structure  $\rightarrow$  Also tetrahedral coordination (Coordination number = 4).

**Step 3: Conclusion.**

Both sphalerite and wurtzite forms of  $\text{ZnS}$  have tetrahedral coordination.

### Quick Tip

ZnS exists in two polymorphic forms: zinc blende (cubic) and wurtzite (hexagonal). Both have coordination number 4 (tetrahedral).

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