Chemical Bonding and Molecular Structure JEE Main PYQ - 3

Total Time: 1 Hour : 15 Minute **Total Marks:** 120

Instructions

Instructions

1. Test will auto submit when the Time is up.
2. The Test comprises of multiple choice questions (MCQ) with one or more correct answers.
3. The clock in the top right corner will display the remaining time available for you to complete the examination.

Navigating & Answering a Question

1. The answer will be saved automatically upon clicking on an option amongst the given choices of answer.
2. To deselect your chosen answer, click on the clear response button.
3. The marking scheme will be displayed for each question on the top right corner of the test window.

Chemical Bonding and Molecular Structure

1. Number of molecules/ions from the following in which the central atom is involved in sp3 hybridization is _____. NO_3^-, BCl_3, ClO_2^-, ClO_3 **(+4, -1)**

a. 2

b. 4

c. 3

d. 1

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2. Which one of the following molecules has maximum dipole moment ? **(+4, -1)**

a. NF_3

b. CH_4

c. NH_3

d. PF_5

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3. Number of complexes from the following with even number of unpaired "d" electrons is _____. **(+4, -1)**

$$[V(H_2O)_6]^{3+},\ [Cr(H_2O)_6]^{2+},\ [Fe(H_2O)_6]^{3+},\ [Ni(H_2O)_6]^{3+},\ [Cu(H_2O)_6]^{2+}$$

[Given atomic numbers: $V = 23,\ Cr = 24,\ Fe = 26,\ Ni = 28,\ Cu = 29$]

a. 2

b. 4

c. 5

d. 1

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4. A diatomic molecule has a dipole moment of $1.2\,D$. If the bond distance is $1\,\text{Å}$, then fractional charge on each atom is _____ $\times 10^{-1}$ esu. **(+4, -1)**
(Given $1\,D = 10^{-18}$ esu cm)

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5. Which of the following is least ionic ? **(+4, -1)**

a. $BaCl_2$

b. AgCl

c. KCl

d. $CoCl_2$

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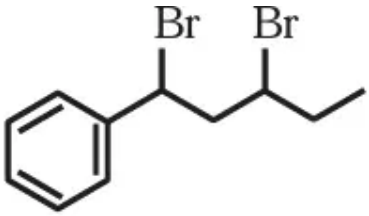
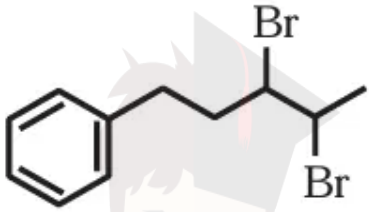
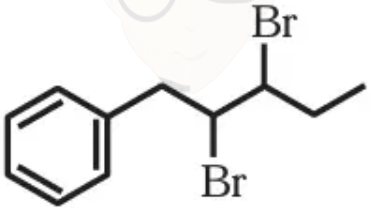
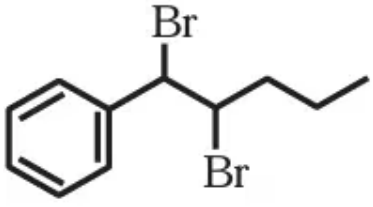
6. Identify structure of 2,3-dibromo-1-phenylpentane. **(+4, -1)**

a. 

b. 

c. 

d. 

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7. The number of species from the following in which the central atom uses sp3 hybrid orbitals in its bonding is _____. $NH_3, SO_2, SiO_2, BeCl_2, CO_2, H_2O, CH_4, BF_3$ **(+4, -1)**

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8. The linear combination of atomic orbitals to form molecular orbitals takes place only when the combining atomic orbitals **(+4, -1)**

A. have the same energy

B. have the minimum overlap

C. have same symmetry about the molecular axis
D. have different symmetry about the molecular axis.
Choose the most appropriate from the options given below:

a. A, B, C only

b. A and C only

c. B and C only

d. B and D only

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9. The number of molecules/ions having trigonal bipyramidal shape is: **(+4, -1)**
PF_5, BrF_5, PCl_5, $[PtCl_4]^{2-}$, BF_3, $Fe(CO)_5$

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10. Arrange the bonds in order of increasing ionic character in the molecules. LiF, K_2O, N_2, SO_2, and ClF_3. **(+4, -1)**

a. $ClF_3 < N_2 < SO_2 < K_2O < LiF$

b. $LiF < K_2O < ClF_3 < SO_2 < N_2$

c. $N_2 < ClF_3 < SO_2 < K_2O < LiF$

d. $N_2 < SO_2 < ClF_3 < K_2O < LiF$

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11. The number of species from the following which are paramagnetic and with bond order equal to one is ___. **(+4, -1)**
Species: H_2, He_2^+, O_2^-, N_2, O_2^{2-}, F_2, Ne_2^+, B_2

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12. Sum of bond order of CO and NO^+ is _____. **(+4, -1)**

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13. The total number of molecules with zero dipole moment among CH_4, BF_3, H_2O, HF, NH_3, CO_2 and SO_2 is _____. **(+4, -1)**

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14. The number of non-polar molecules from the following is _____. **(+4, -1)**
HF, H_2O, SO_2, H_2, CO_2, CH_4, NH_3, HCl, $CHCl_3$, BF_3

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15. Total number of ions from the following with noble gas configuration is _____. **(+4, -1)**
Sr^{2+} $(Z = 38)$, Cs^+ $(Z = 55)$, La^{2+} $(Z = 57)$, Pb^{2+} $(Z = 82)$, Yb^{2+} $(Z = 70)$, and Fe^{2+} $(Z = 26)$

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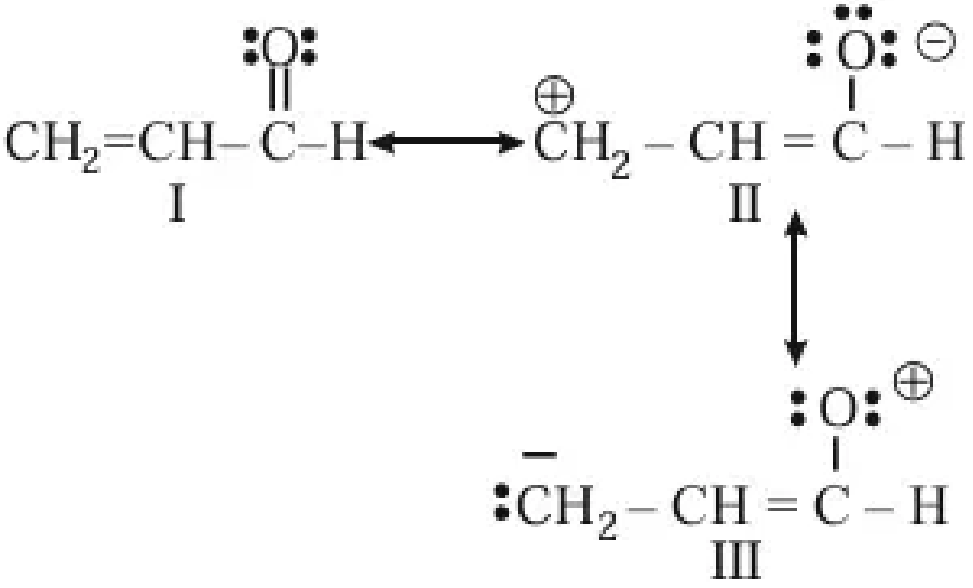
16. The order of relative stability of the contributing structure is: **(+4, -1)**



Choose the correct answer from the options given below:

a. $I > II > III$

b. $II > I > III$

c. $I = II = III$

d. $III > II > I$

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17. 30.4 kJ is required to melt one mole of sodium chloride. The entropy change during melting is 28.4 JK-1 mol-1 . atm The melting point of sodium chloride Calculate is ___ K(Nearest Integer). **(+4, -1)**

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18. S-I: Sulphur exists as S8 while oxygen exists as O2.
S-II: In oxygen, pπ-pπ bonding occurs while it is not effective in sulphur. **(+4, -1)**

a. Both S-I and S-II are true

b. S-I is true and S-II is false

c. S-I is false and S-II is true

d. Both S-I and S-II are false

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19. Chemical formula of compound present in tooth enamel? **(+4, -1)**

a. $Ca_{10}(PO_4)_6(OH)_2$

b. $Ca_8(PO_4)_4(OH)_2$

c. $Ca_6(PO_4)_2(OH)_2$

d. $Ca_8(PO_4)_6(OH)_2$

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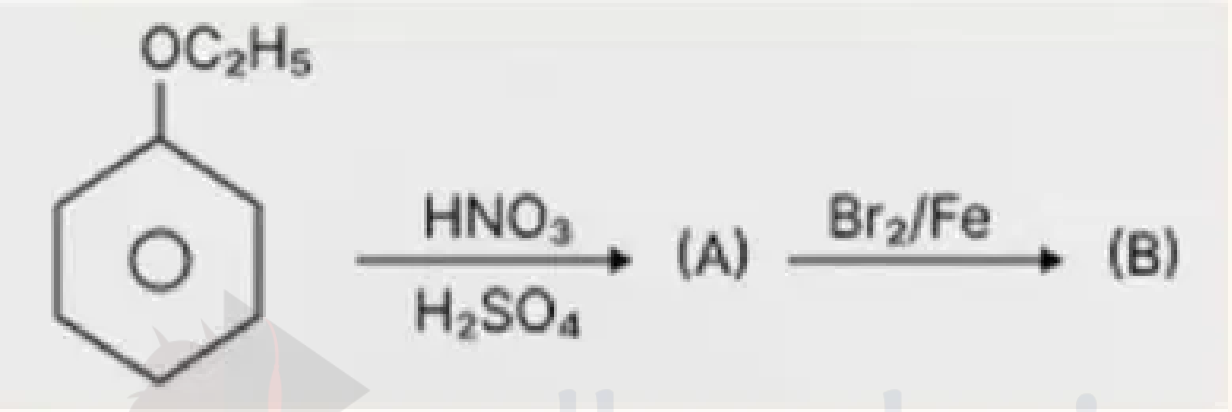
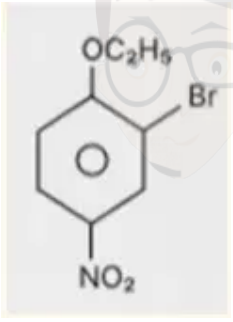
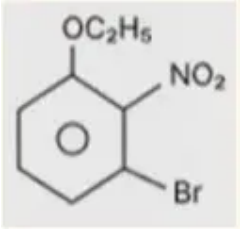
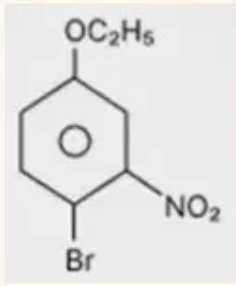
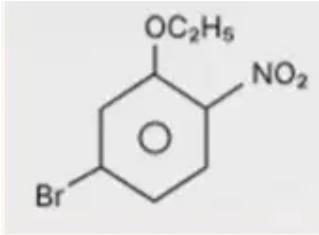
20. Product B is: **(+4, -1)**



a. 

b. 

c. 

d. 

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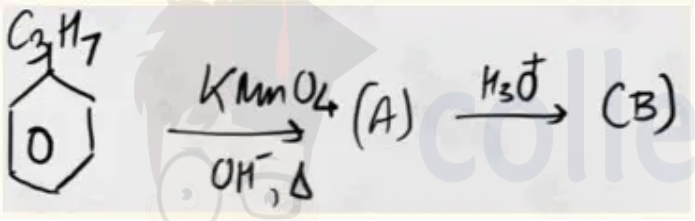
21. Number of π bonds present in product B is: **(+4, -1)**



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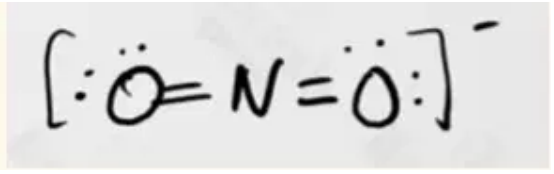
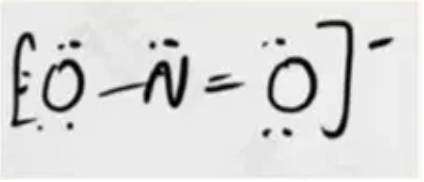
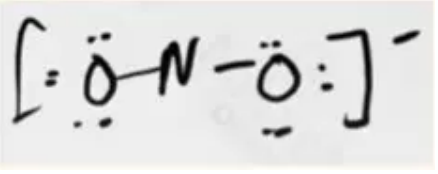
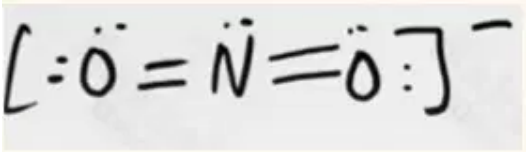
22. Identify the correct Lewis dot structure of NO^-_2. **(+4, -1)**

a. 

b. 

c. 

d. 

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23. The maximum number of lone pairs of electron on the central atom from the following species is _____ **(+4, -1)**

ClO_3^-, XeF_4, SF_4 and I_3^-

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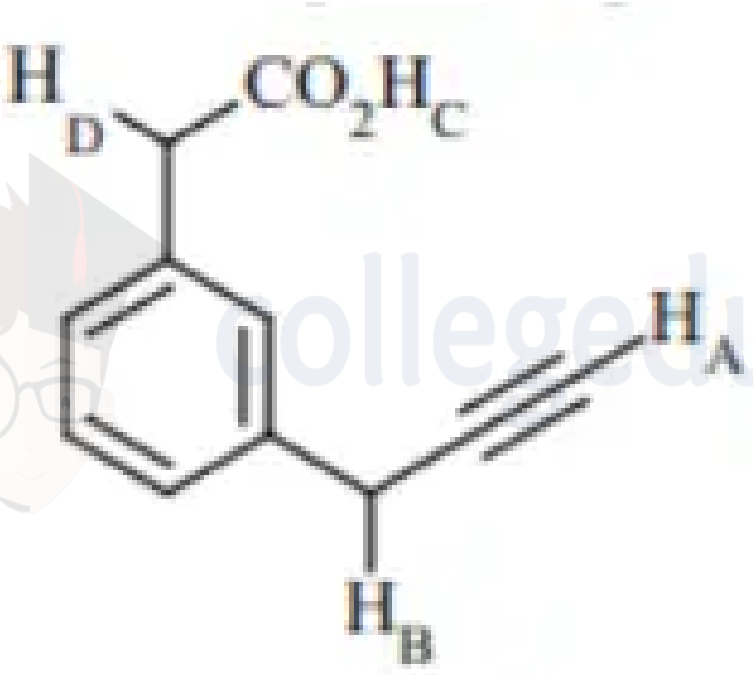
24. What is the correct order of acidity of the protons marked A-D in the given compounds? **(+4, -1)**



a. $H_C > H_D > H_A > H_B$

b. $H_B > H_A > H_D > H_C$

c. $H_A > H_B > H_C > H_D$

d. $H_C > H_A > H_D > H_B$

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25. For OF_2 molecule consider the following: **(+4, -1)**

[(A)] Number of lone pairs on oxygen is 2.

[(B)] F--O--F angle is less than $104.5°$.

[(C)] Oxidation state of O is -2.

[(D)] Molecule is bent 'V'-shaped.
[(E)] Molecular geometry is linear.
Correct options are:

a. A, B, D only

b. A, C, D only

c. C, D, E only

d. B, E, A only

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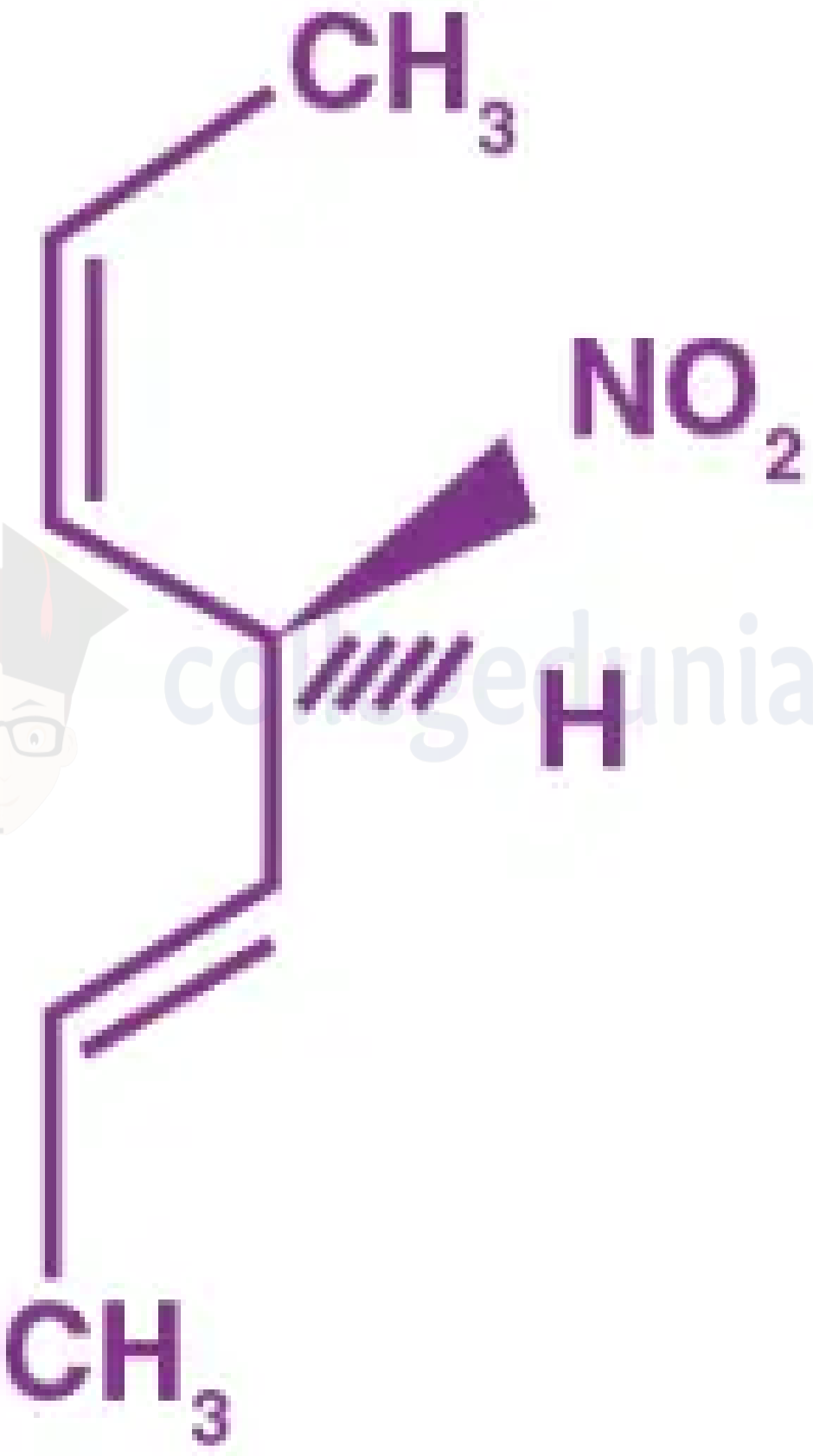
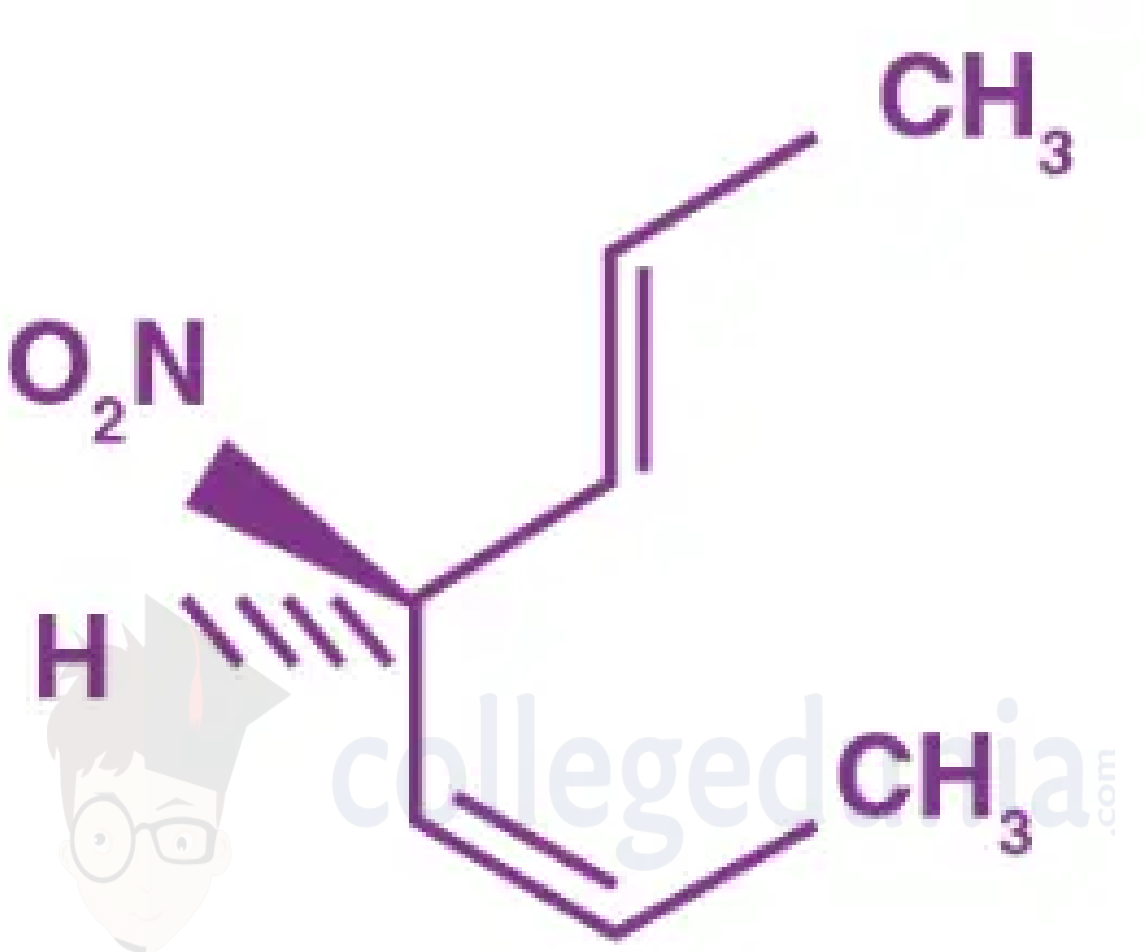
26. Given below are two statements. **(+4, -1)**
Statement I: The compound


CH3
NO2
H
CH3

(A) is optically active.

Statement II:



is a mirror image of the above compound A.

In the light of the above statement, choose the most appropriate answer from the options given below.

a. Both Statement I and Statement II are correct.

b. Both Statement I and Statement II are incorrect.

c. Statement I is correct but Statement II is incorrect.

d. Statement I is incorrect but Statement II is correct.

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27. Number of compounds with one lone pair of electrons on central atom amongst following is **(+4, -1)**

$ClF_3, XeO_3, BrF_5, XeF_4, O_3, NH_3, H_2O$

28. Number of lone pairs of electrons in the central atom of SCl_2, O_3, ClF_3 and SF_6, respectively, are: **(+4, -1)**

a. 0, 1, 2 and 2

b. 2, 1, 2 and 0

c. 1, 2, 2 and 0

d. 2, 1, 0 and 2

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29. Consider the following sulphur based oxoacids. **(+4, -1)**
$H_2SO_3, H_2SO_4, H_2S_2O_8$ and $H_2S_2O_7$
Amongst these oxoacids, the number of those with peroxo (O–O) bonds is_____.

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30. In a linear tetrapeptide (constituted with different amino acids) – (number of peptide bonds) is _____. **(+4, -1)**

Answers

1. **Answer: a**

Explanation:

Determine the Hybridization of Each Species:

For NO_3^-:
The structure of NO_3^- (nitrate ion) shows that nitrogen is involved in three sigma bonds and has no lone pairs on it.
The hybridization of nitrogen in NO_3^- is sp^2.

For BCl_3:
Boron in BCl_3 forms three sigma bonds with no lone pairs, resulting in a planar triangular structure.
The hybridization of boron in BCl_3 is sp^2.

For ClO_2^-:
In ClO_2^- (chlorite ion), chlorine has two sigma bonds with oxygen atoms and two lone pairs of electrons.
The hybridization of chlorine in ClO_2^- is sp^3, resulting in a bent or V-shaped geometry.

For ClO_3^-:
In ClO_3^- (chlorate ion), chlorine forms three sigma bonds with oxygen atoms and has one lone pair of electrons.
The hybridization of chlorine in ClO_3^- is sp^3, resulting in a pyramidal structure.

Count the Species with sp^3 Hybridization:

From the analysis above, ClO_2^- and ClO_3^- have sp^3 hybridization for the central atom.

Conclusion:

The number of molecules/ions in which the central atom is involved in sp^3 hybridization is 2, corresponding to Option (1).

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2. **Answer: c**

Explanation:

Understanding Dipole Moment:

The dipole moment is a measure of the separation of positive and negative charges in a molecule. Molecules with polar bonds and an asymmetrical shape often exhibit a dipole moment. The greater the electronegativity difference between atoms and the asymmetry in the molecule, the higher the dipole moment.

Analyze Each Molecule:

NF_3: Although nitrogen and fluorine have a large electronegativity difference, the structure of NF_3 (trigonal pyramidal) leads to a partial cancellation of the dipole moment due to the lone pair on nitrogen. The net dipole moment of NF_3 is lower than that of NH_3.

CH_4: Methane (CH_4) is a nonpolar molecule with a tetrahedral structure. The dipole moments of the C-H bonds cancel each other out, resulting in a net dipole moment of zero.

NH_3: Ammonia has a trigonal pyramidal structure with a lone pair of electrons on nitrogen. This creates an asymmetrical distribution of charge and a significant net dipole moment. The lone pair on nitrogen intensifies the dipole moment, making NH_3 have a higher dipole moment compared to NF_3.

PF_5: Phosphorus pentafluoride has a trigonal bipyramidal structure, with dipole moments of the axial and equatorial bonds cancelling out. As a result, PF_5 is a nonpolar molecule with a net dipole moment of zero.

Conclusion:

Among the given molecules, NH_3 has the maximum dipole moment due to its trigonal pyramidal structure and the presence of a lone pair on nitrogen, which increases its net dipole moment.

---

3. **Answer: a**

Explanation:

Determine the Electronic Configuration of Each Metal Ion:

For $[V(H_2O)_6]^{3+}$:

Vanadium (V) has an atomic number of 23, with an electronic configuration of $[Ar]3d^3 4s^2$. V^{3+} configuration: $[Ar]3d^2$. Number of unpaired d-electrons: 2 (even number).

For $[Cr(H_2O)_6]^{2+}$:

Chromium (Cr) has an atomic number of 24, with an electronic configuration of $[\mathrm{Ar}]3d^5 4s^1$. Cr^{2+} configuration: $[\mathrm{Ar}]3d^4$. Number of unpaired d-electrons: 4 (even number).

For $[Fe(\mathrm{H_2O})_6]^{3+}$:

Iron (Fe) has an atomic number of 26, with an electronic configuration of $[\mathrm{Ar}]3d^6 4s^2$. Fe^{3+} configuration: $[\mathrm{Ar}]3d^5$. Number of unpaired d-electrons: 5 (odd number).

For $[Ni(\mathrm{H_2O})_6]^{3+}$:

Nickel (Ni) has an atomic number of 28, with an electronic configuration of $[\mathrm{Ar}]3d^8 4s^2$. Ni^{3+} configuration: $[\mathrm{Ar}]3d^7$. Number of unpaired d-electrons: 3 (odd number).

For $[Cu(\mathrm{H_2O})_6]^{2+}$:

Copper (Cu) has an atomic number of 29, with an electronic configuration of $[\mathrm{Ar}]3d^{10} 4s^1$. Cu^{2+} configuration: $[\mathrm{Ar}]3d^9$. Number of unpaired d-electrons: 1 (odd number).

Count Complexes with Even Number of Unpaired Electrons:

From the analysis above, only $[V(\mathrm{H_2O})_6]^{3+}$ and $[Cr(\mathrm{H_2O})_6]^{2+}$ have an even number of unpaired d-electrons.

Conclusion:

The number of complexes with an even number of unpaired d-electrons is 2, corresponding to Option (1).

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4. Answer: 0.25 - 0.25

Explanation:

For the reaction:

$$A(g) \rightleftharpoons B(g) + \frac{1}{2}C(g)$$

Step 1: Initial moles

Initial moles:

$$n_A = n, \quad n_B = 0, \quad n_C = 0$$

Step 2: Equilibrium moles

Equilibrium moles:

$$n_A = n(1-\alpha), \quad n_B = n\alpha, \quad n_C = \frac{n\alpha}{2}$$

Step 3: Total moles at equilibrium

Total moles:

$$n_{\text{total}} = n\,(1+\alpha)\,/2$$

Step 4: Equilibrium pressure expressions

Equilibrium pressure for each component:

$$P_A = \frac{(1-\alpha)P}{1+\alpha/2}, \quad P_B = \frac{\alpha P}{1+\alpha/2}, \quad P_C = \frac{(\alpha/2)P}{1+\alpha/2}$$

Step 5: Expression for K_p

The equilibrium constant K_p is given by:

$$K_p = \frac{\alpha P}{1+\alpha/2} \times \left(\frac{\alpha P}{(2+\alpha)}\right)^{1/2}$$

Step 6: Simplification

Simplifying the expression for K_p:

$$K_p = \frac{\alpha P}{1+\alpha/2} \times \frac{\alpha P^{1/2}}{(2+\alpha)^{1/2}}$$

Final expression for K_p:

$$K_p = \frac{\alpha^{3/2} P^{1/2}}{(2+\alpha)^{1/2}(1-\alpha)}$$

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5. Answer: b

Explanation:

To determine which compound is the least ionic among the given options, we need to understand the concept of ionic character. The ionic character in a compound largely depends on the difference in electronegativity between the cation and the anion and the size of the ions comprising the compound.

The options provided are:

- $BaCl_2$
- AgCl
- KCl
- $CoCl_2$

Let's analyze each option considering the factors that affect ionic character:

1. $BaCl_2$: Barium (Ba) is a Group 2 element, and Chlorine (Cl) is a halogen. The high electronegativity difference results in a strong ionic bond.
2. **AgCl:** Silver chloride has lower ionic character due to the relatively smaller electronegativity difference between silver (Ag) and chlorine (Cl). Also, silver ions tend to form covalent bonds due to their filled d-orbitals, which reduce the overall ionic character.
3. **KCl:** Potassium (K) is an alkali metal, and Chlorine (Cl) is a halogen, resulting in a large electronegativity difference and thus, a strong ionic character.
4. $CoCl_2$: Cobalt (Co), a transition metal, forms ionic compounds with halogens. However, due to partial covalent character from its d-orbitals, it might be less ionic than simple alkali or alkaline earth metals, but more ionic than AgCl.

Based on the above analysis, **AgCl** is the least ionic compound among the given options due to the reasons mentioned, such as closer electronegativity and covalent character contributed by silver's orbital configurations.

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6. Answer: c

Explanation:

To identify the structure of 2,3-dibromo-1-phenylpentane, let's break down and understand the name as per IUPAC nomenclature rules:

1. The base name "pentane" indicates a five-carbon alkane chain.
2. A "phenyl" group at position 1 implies a benzene ring (phenyl) is attached to the first carbon of this chain.
3. There are two bromine atoms attached at positions 2 and 3 of the carbon chain, indicated by "2,3-dibromo".

Let's now analyze the given options to determine which structure represents 2,3-dibromo-1-phenylpentane:

- **Option 1**: The structure shows incorrect placement of the bromine atoms and the phenyl group.

- **Option 2**: This structure similarly misplaces the substituents according to our IUPAC breakdown.
- **Option 3**: Here, the structure accurately places a phenyl group at the first carbon, and the bromine atoms at the second and third carbons:

Thus, Option 3 correctly depicts 2,3-dibromo-1-phenylpentane.

- **Option 4**: The arrangement doesn't conform to the specifications given in the name.

Therefore, the correct answer is represented by the structural diagram in Option 3.

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7. Answer: 4 - 4

Explanation:

Given Molecules and Their Hybridization:

- **NH_3:** Ammonia has sp^3 hybridization, as the nitrogen atom forms three sigma bonds with hydrogen atoms and has one lone pair, which leads to a tetrahedral geometry.
- **SO_2:** Sulfur dioxide exhibits sp^2 hybridization. The sulfur atom forms two sigma bonds with oxygen atoms and has one lone pair, resulting in a bent molecular shape.
- **SiO_2:** Silicon dioxide also exhibits sp^3 hybridization. The silicon atom forms four bonds with oxygen atoms, and its structure adopts a tetrahedral geometry.
- **$BeCl_2$:** Beryllium chloride has sp hybridization. The beryllium atom forms two sigma bonds with chlorine atoms in a linear arrangement.
- **CO_2:** Carbon dioxide has sp hybridization. The carbon atom forms two double bonds with two oxygen atoms in a linear structure.
- **H_2O:** Water exhibits sp^3 hybridization. The oxygen atom forms two sigma bonds with hydrogen atoms and has two lone pairs, giving the molecule a bent structure.
- **CH_4:** Methane has sp^3 hybridization. The carbon atom forms four sigma bonds with hydrogen atoms in a tetrahedral arrangement.
- **BF_3:** Boron trifluoride exhibits sp^2 hybridization. The boron atom forms three sigma bonds with fluorine atoms in a trigonal planar geometry.

Correct Answer: The correct answer is **4**.

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8. Answer: b

Explanation:

The concept in focus here is about the formation of molecular orbitals through the linear combination of atomic orbitals (LCAO). To ensure the combination occurs effectively, the atomic orbitals must satisfy certain conditions:

- **Same Energy:** Atomic orbitals can combine effectively if they have the same or very similar energies. This allows the wave functions to combine constructively, forming stable molecular orbitals.
- **Same Symmetry:** The combining atomic orbitals must have the same symmetry with respect to the molecular axis. Only orbitals that possess compatible symmetries can combine to form molecular orbitals.

Now, let's analyze each option:

- **Option A (Same Energy):** Correct, as having the same energy is essential for effective overlap and combination.
- **Option B (Minimum Overlap):** Incorrect, as a significant overlap is necessary for the effective combination of orbitals. Minimum overlap would defeat the purpose since it doesn't provide enough interaction space.
- **Option C (Same Symmetry):** Correct, same symmetry is a critical requirement for the combination of atomic orbitals to form molecular orbitals.
- **Option D (Different Symmetry):** Incorrect, as orbits with different symmetries cannot effectively combine.

Given these considerations, the most appropriate options that ensure the combination of atomic orbitals are **A and C only**. Hence, the correct answer is:

Answer: A and C only

---

9. **Answer: 3 - 3**

Explanation:

PF_5, PCl_5, and $Fe(CO)_5$ have trigonal bipyramidal geometry.

BrF_5: square pyramidal

$[PtCl_4]^{2-}$: square planar

BF_3: trigonal planar

---

10. **Answer: c**

Explanation:

The problem asks to arrange the chemical bonds in the given molecules—LiF, K_2O, N_2, SO_2, and ClF_3—in order of increasing ionic character.

Concept Used:

The ionic character of a chemical bond is a measure of its polarity. It depends on the difference in electronegativity (ΔEN) between the two atoms forming the bond. According to the Pauling scale, a larger difference in electronegativity corresponds to a greater charge separation and thus a higher degree of ionic character.

The relationship can be summarized as:

$$\text{Ionic Character} \propto \Delta EN = |EN_{\text{atom 1}} - EN_{\text{atom 2}}|$$

A bond between identical atoms ($\Delta EN = 0$) is purely covalent, while a bond with a large ΔEN (typically > 1.7) is considered predominantly ionic.

Step-by-Step Solution:

Step 1: Identify the specific bonds within each molecule that need to be compared.

- In LiF, the bond is Li–F.
- In K_2O, the bond is K–O.
- In N_2, the bond is N≡N.
- In SO_2, the bond is S=O.
- In ClF_3, the bond is Cl–F.

Step 2: List the Pauling electronegativity (EN) values for each of the atoms involved.

- Nitrogen (N): 3.04
- Sulfur (S): 2.58
- Oxygen (O): 3.44
- Chlorine (Cl): 3.16
- Fluorine (F): 3.98
- Lithium (Li): 0.98
- Potassium (K): 0.82

Step 3: Calculate the electronegativity difference (ΔEN) for each bond.

For N_2 (N≡N bond):

$$\Delta EN = |3.04 - 3.04| = 0$$

This is a purely covalent bond.

For ClF_3 (Cl–F bond):

$$\Delta EN = |3.98 - 3.16| = 0.82$$

This is a polar covalent bond.

For SO_2 (S=O bond):

$$\Delta EN = |3.44 - 2.58| = 0.86$$

This is also a polar covalent bond, slightly more polar than the Cl–F bond.

For K_2O (K–O bond):

$$\Delta EN = |3.44 - 0.82| = 2.62$$

This is a predominantly ionic bond.

For LiF (Li–F bond):

$$\Delta EN = |3.98 - 0.98| = 3.00$$

This is a highly ionic bond.

Final Computation & Result:

Step 4: Arrange the bonds in order of increasing ΔEN, which corresponds to increasing ionic character.

The calculated ΔEN values in increasing order are:

$$0 \text{ (for } N_2) < 0.82 \text{ (for } ClF_3) < 0.86 \text{ (for } SO_2) < 2.62 \text{ (for } K_2O) < 3.00 \text{ (for LiF)}$$

Therefore, the order of increasing ionic character for the bonds in the given molecules is:

N_2 < ClF_3 < SO_2 < K_2O < LiF

---

11. Answer: 1 - 1

Explanation:

To determine the number of paramagnetic species with a bond order of one, we analyze each given species:

Step 1: Identify Paramagnetic Species

- **Paramagnetic:** Contains unpaired electrons in molecular orbitals.
- **Non-paramagnetic:** All electrons are paired.

Step 2: Calculate Bond Order

Bond order is calculated as: (Number of electrons in bonding orbitals-Number of electrons in antibonding orbitals)2

Species	Electron Configuration	Paramagnetic?	Bond Order
H_2	$(\sigma_{1s})^2$	No	1
He_2^+	$(\sigma_{1s})^2(\sigma^*_{1s})^1$	Yes	0.5
O_2^-	$(\sigma_{2s})^2(\sigma^*_{2s})^2(\pi_{2p})^4(\pi^*_{2p})^3$	Yes	1.5
N_2	$(\sigma_{2s})^2(\sigma^*_{2s})^2(\pi_{2p})^4(\sigma_{2p})^2$	No	3
O_2^{2-}	$(\sigma_{2s})^2(\sigma^*_{2s})^2(\pi_{2p})^4(\pi^*_{2p})^4$	No	1
F_2	$(\sigma_{2s})^2(\sigma^*_{2s})^2(\pi_{2p})^4(\pi^*_{2p})^4(\sigma_{2p})^2$	No	1
Ne_2^+	$(\sigma_{2s})^2(\sigma^*_{2s})^2(\pi_{2p})^4(\pi^*_{2p})^4(\sigma_{2p})^1$	Yes	0.5
B_2	$(\sigma_{2s})^2(\sigma^*_{2s})^2(\pi_{2p})^1(\pi_{2p})^1$	Yes	1

Final Results

- Paramagnetic and Bond Order = 1: O_2^{2-}, F_2, B_2

The total number of species matching the criteria is: **1**. This satisfies the range (1,1).

---

12. Answer: 6 - 6

Explanation:

To find the sum of the bond orders of CO and NO^+, we start by determining each bond order individually.

Bond Order of CO:

- The bond order of a molecule is calculated using the molecular orbital theory. For CO, it has 10 valence electrons.

- The electronic configuration in terms of molecular orbitals is: $(\sigma_{1s})^2(\sigma^*_{1s})^2(\sigma_{2s})^2(\sigma^*_{2s})^2(\pi_{2p})^4(\sigma_{2p})^2$.
- Bond order = (Number of electrons in bonding orbitals - Number of electrons in antibonding orbitals) / 2.
- Bond order of CO: (8-2)/2 = 3.

Bond Order of NO^+:

- NO^+ has 10 valence electrons since it is derived from NO by losing an electron.
- Its electronic configuration is similar to that of CO.
- Therefore, its bond order calculation is the same as CO.
- Bond order of NO^+: (8-2)/2 = 3.

Sum of Bond Orders:

- Sum = Bond order of CO + Bond order of NO^+ = 3 + 3 = 6.

Verification:

- The calculated sum is 6. This fits precisely within the provided range of 6 to 6.

Thus, the sum of the bond orders of CO and NO^+ is **6**.

---

13. **Answer: 3 - 3**

Explanation:

To determine the number of molecules with zero dipole moment among the given set (CH_4, BF_3, H_2O, HF, NH_3, CO_2, SO_2), we analyze each molecule's structure and geometry.

CH_4: Tetrahedral geometry with symmetrical distribution of charge. Resultantly, it has a zero dipole moment.

BF_3: Trigonal planar geometry, symmetry in the molecule cancels out dipoles, resulting in a zero dipole moment.

H_2O: Bent geometry with a dipole moment due to the lone pairs on oxygen causing an asymmetric charge distribution.

HF: Linear molecule, but the difference in electronegativity induces a dipole moment.

NH_3: Trigonal pyramidal geometry with a net dipole moment due to the lone pair on nitrogen.

CO_2: Linear geometry; symmetrical opposite dipoles cancel out, resulting in a zero dipole moment.

SO_2: Bent geometry with a net dipole moment due to the asymmetric distribution of electrons.

Upon reviewing, CH_4, BF_3, and CO_2 have zero dipole moments. Therefore, the total number is **3.**

---

14. Answer: 4 - 4

Explanation:

The problem requires identifying non-polar molecules from a given list. A molecule is non-polar when its net dipole moment is zero, often due to a symmetrical arrangement of polar bonds that cancel each other out.

- **HF:** Polar due to significant electronegativity difference between H and F.
- **H_2O:** Polar because of its bent shape, causing a net dipole moment.
- **SO_2:** Polar, also due to its bent shape.
- **H_2:** Non-polar, as identical atoms share electrons equally.
- **CO_2:** Non-polar; although it has polar bonds, the linear geometry cancels the dipoles.
- **CH_4:** Non-polar; tetrahedral symmetry allows dipoles to cancel.
- **NH_3:** Polar due to its trigonal pyramidal shape.
- **HCl:** Polar due to electronegative difference.
- **$CHCl_3$:** Polar; asymmetric molecule leads to a net dipole.
- **BF_3:** Non-polar; trigonal planar symmetry results in dipole cancellation.

Counting non-polar molecules, we have **H_2, CO_2, CH_4,** and **BF_3.** Hence, there are 4 non-polar molecules.

---

15. Answer: 2 - 2

Explanation:

To find the total number of ions with a noble gas configuration from the given ions, we must first understand what a noble gas configuration entails. Noble gases have full

electron shells, resulting in a stable electronic arrangement. To determine if an ion has a noble gas configuration, consider the ion's total electrons and compare it with the nearest noble gas.

- **Sr^{2+} (Z = 38):** The electron configuration for Sr is [Kr] $5s^2$. Sr^{2+} loses 2 electrons, resulting in [Kr], which is the noble gas krypton.
- **Cs^{+} (Z = 55):** The electron configuration for Cs is [Xe] $6s^1$. Cs^{+} loses 1 electron, resulting in [Xe], the noble gas xenon.
- **La^{2+} (Z = 57):** The electron configuration for La is [Xe] $5d^1$ $6s^2$. La^{2+} loses 2 electrons, resulting in [Xe] $5d^1$, which is not a noble gas configuration.
- **Pb^{2+} (Z = 82):** The electron configuration for Pb is [Xe] $4f^{14}$ $5d^{10}$ $6s^2$ $6p^2$. Pb^{2+} loses 2 electrons, becoming [Xe] $4f^{14}$ $5d^{10}$ $6s^2$, not a noble gas configuration.
- **Yb^{2+} (Z = 70):** The electron configuration for Yb is [Xe] $4f^{14}$ $6s^2$. Yb^{2+} loses 2 electrons, resulting in [Xe] $4f^{14}$, not a noble gas configuration.
- **Fe^{2+} (Z = 26):** The electron configuration for Fe is [Ar] $3d^6$ $4s^2$. Fe^{2+} loses 2 electrons, resulting in [Ar] $3d^6$, not a noble gas configuration.

Upon analyzing, Sr^{2+} and Cs^{+} achieve noble gas configurations. Therefore, the total number of ions with noble gas configurations is 2.

---

16. Answer: a

Explanation:

The question involves evaluating the relative stability of resonance structures. The stability of resonance structures mainly depends on the following factors:

1. **Octet Rule:** Structures in which all atoms, especially the second-period elements (like C, N, O), have a complete octet are generally more stable.
2. **Charge Separation:** Structures with minimal charge separation are more stable. A neutral structure is more stable than a charged one.
3. **Negative Charge on Electronegative Atoms:** Structures that place negative charges on more electronegative atoms (like oxygen) are more stable.
4. **Positive Charge on Electropositive Atoms:** Structures with positive charges on less electronegative atoms are more stable.

Analyzing the given resonance structures:

1. **Structure I:** This structure has all atoms with complete octet shells and no charge separation, which makes it the most stable.

2. **Structure II:** This structure has a positive charge on a carbon atom and a negative charge on an oxygen atom. There is charge separation, reducing its stability compared to structure I, but the charges are on appropriate atoms considering electronegativity.
3. **Structure III:** This structure has carbon with a negative charge and oxygen with a positive charge, which is the least favorable distribution of charges among the given structures. This increases its instability compared to the other structures.

Based on the above analysis, the order of relative stability is:

$I > II > III$

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17. Answer: 1070 - 1070

Explanation:

Melting Point Calculation

The relationship between enthalpy (ΔH) and entropy (ΔS) during a phase change is:

$\Delta S = \Delta H / T$

Rearranging to find the melting point (T_{mp}):

$T_{mp} = \Delta H / \Delta S$

Substituting the given values:

- $\Delta H = 30.4$ kJ $mol^{-1} = 30400$ J mol^{-1}
- $\Delta S = 28.4$ J K^{-1} mol^{-1}

$T_{mp} = 30400 / 28.4 = 1070.42$ K

The nearest integer value is 1070 K.

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18. Answer: a

Explanation:

The Correct answer is option is (A) : Both S-I and S-II are true

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19. **Answer: a**

Explanation:

The Correct answer is option is (A) : $Ca_{10}(PO_4)_6(OH)_2$

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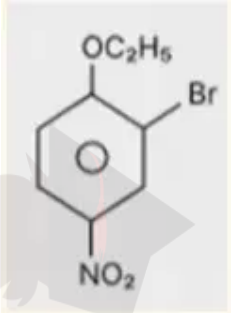
20. **Answer: a**

Explanation:

The Correct answer is option is (A) :



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21. **Answer: 4 - 4**

Explanation:

The correct answer is 4.

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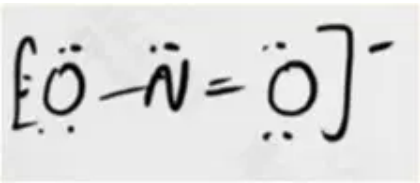
22. **Answer: b**

Explanation:

The Correct answer is option is (B) :

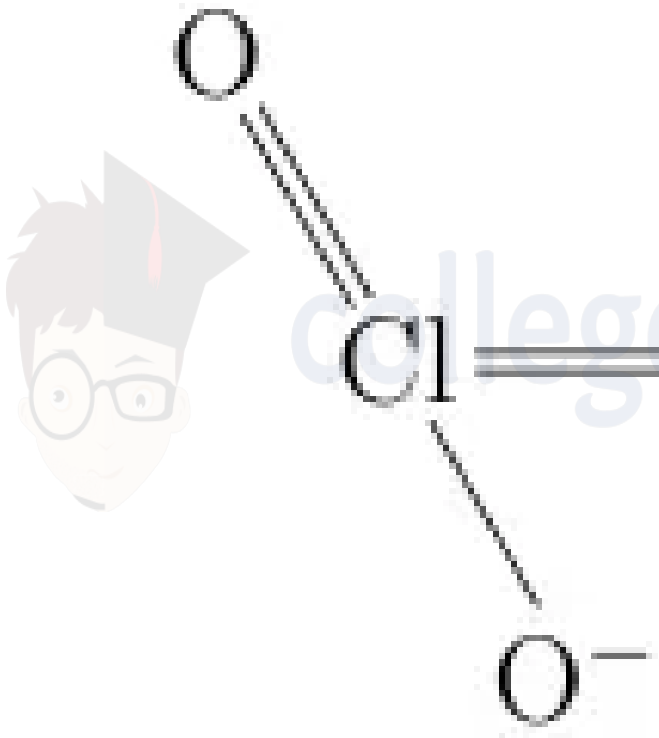


23. **Answer: 3 - 3**

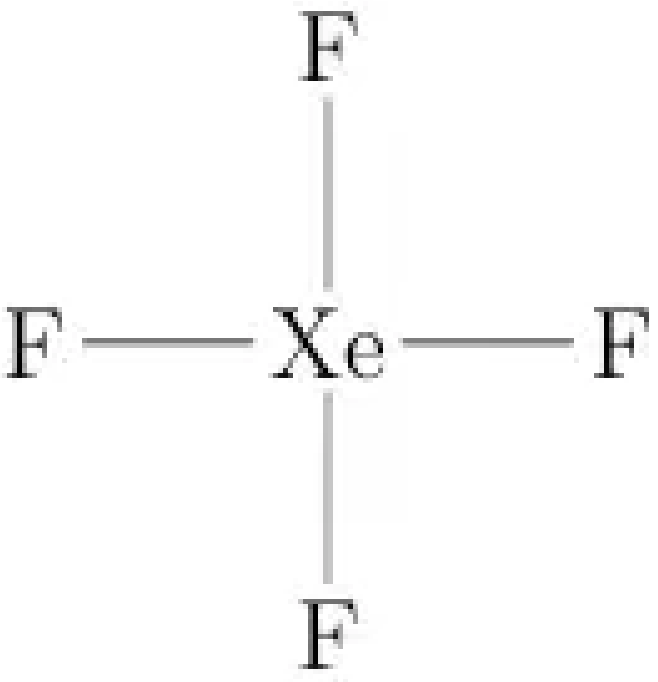
Explanation:

Analysis of Each Species:

1. ClO_3^- **(Chlorate ion):**
   - Chlorine (Cl) is the central atom with 7 valence electrons.
   - Each oxygen (O) contributes 6 valence electrons, and the negative charge adds 1 more electron.
   - Total valence electrons: $7 + (3 \times 6) + 1 = 26$.

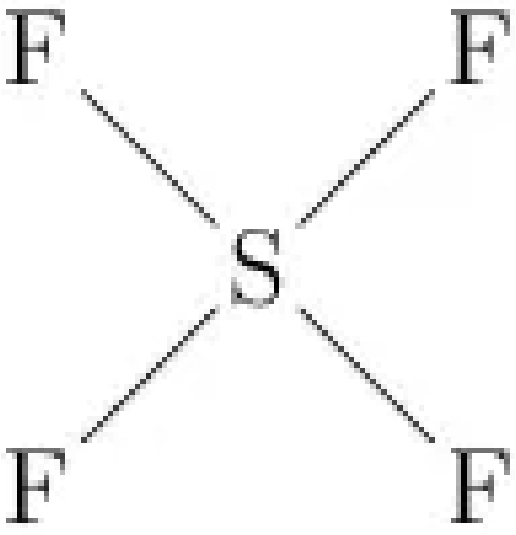


   - Lewis structure: The central Cl atom has 1 lone pair.
2. XeF_4 **(Xenon tetrafluoride):**
   - Xenon (Xe) is the central atom with 8 valence electrons.
   - Each fluorine (F) contributes 7 valence electrons.
   - Total valence electrons: $8 + (4 \times 7) = 36$.



- Lewis structure: The central Xe atom has 2 lone pairs.

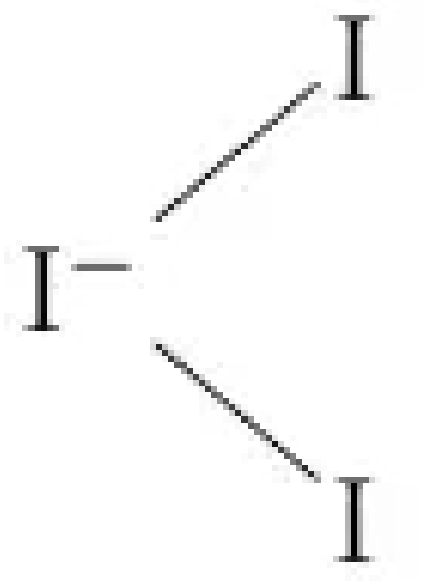
3. $\mathrm{SF_4}$ **(Sulfur tetrafluoride):**
   - Sulfur (S) is the central atom with 6 valence electrons.
   - Each fluorine (F) contributes 7 valence electrons.
   - Total valence electrons: $6 + (4 \times 7) = 34$.



- Lewis structure: The central S atom has 1 lone pair.

4. $\mathrm{I_3^-}$ **(Triiodide ion):**

- Iodine (I) is the central atom with 7 valence electrons.
- Each other iodine atom contributes 7 valence electrons, and the negative charge adds 1 more electron.
- Total valence electrons: $7 + 7 + 7 + 1 = 22$.



- Lewis structure: The central I atom has 3 lone pairs.

Final Answer:

The maximum number of lone pairs on the central atom is 3 (in $\mathrm{I_3^-}$).

---

24. Answer: a

Explanation:

The acidity of a proton depends on the stability of the conjugate base formed after its removal. The more stable the conjugate base, the more acidic the proton.
Step 1: Acidity of H_C
H_C is the most acidic proton. Removal of H_C results in a carboxylate anion, which is highly stabilized by resonance.\\[6pt]
Step 2: Acidity of H_D
H_D is the second most acidic proton. Its removal forms a carbanion that is stabilized by resonance with the benzene ring.\\[6pt]
Step 3: Acidity of H_A vs H_B
H_A is more acidic than H_B. The conjugate base formed after the removal of H_A is stabilized by resonance with the triple bond, resulting in a negative charge being distributed over two carbon atoms. However, the conjugate base formed after removal of H_B results in the negative charge adjacent to the triple bond, with no resonance

stabilization. This is highly unstable due to the electron-withdrawing nature of the sp-hybridized carbon in the alkyne.
Step 4: Overall Acidity Order
Therefore, the correct order of acidity is:

$$H_C > H_D > H_A > H_B.$$

Conclusion: Option (2) is correct.

---

25. Answer: a

Explanation:

The Lewis structure of OF_2 shows that the central oxygen atom is bonded to two fluorine atoms and has two lone pairs of electrons. This gives the molecule a bent or 'V' shape due to lone pair-bond pair repulsion.
Step 2: Analysis of the Statements
Statement (A): The oxygen atom in OF_2 has 6 valence electrons. After forming two single bonds with fluorine atoms, it retains two lone pairs. Hence, this statement is true
Statement (B): The F--O--F bond angle is slightly less than 104.5° (around 102°) due to increased lone pair-bond pair repulsion. Hence, this statement is true
Statement (C): The oxidation state of oxygen in OF_2 is $+2$, as fluorine is more electronegative. Hence, this statement is false
Statement (D): Due to the presence of two lone pairs, the molecular geometry is bent or 'V' shaped. Hence, this statement is true
Statement (E) The molecular geometry is not linear; it is bent. Hence, this statement is false}
Conclusion: The correct statements are A, B, D. Therefore, the correct answer is (4) A, B, D only

---

26. Answer: c

Explanation:

The given question involves the concepts of optical activity and chirality in organic chemistry.

1. **Statement I Analysis:**

- The compound shown (Compound A) is a type of organic molecule with a chiral center. A chiral center is typically a carbon atom bonded to four different groups.
- Optical activity arises from such chiral centers because they allow the molecule to rotate plane-polarized light. If a compound can do this, it is considered optically active.
- In the structure of the compound, the presence of stereochemistry (the depiction of the spatial arrangement with bonds shown as wedges and dashes) indicates chirality, making the compound optically active.

2. **Statement II Analysis:**
   - The second structure is provided as a potential mirror image of Compound A.
   - A true mirror image of a chiral compound would be its enantiomer, which would also be chiral but with an opposite configuration (i.e., if the original compound is R, the mirror image should be S, and vice versa).
   - By visual inspection, the given mirror image does not change the configuration correctly, or if it does not maintain the enantiomeric nature, it fails to be a true mirror image.

Based on the analyses, the correct option is: **Statement I is correct but Statement II is incorrect.**

Concepts:

1. Chemical Bonding and Molecular Structure:

Such a group of atoms is called a molecule. Obviously, there must be some force that holds these constituent atoms together in the molecules. The attractive force which holds various constituents (atoms, ions, etc.) together in different chemical species is called a chemical bond.

Types of Chemical Bonds:

There are 4 types of chemical bonds which are formed by atoms or molecules to yield compounds.

- **Ionic Bonds -** Ionic bonding is a type of chemical bonding which involves a transfer of electrons from one atom or molecule to another.
- **Covalent Bonds -** Compounds that contain carbon commonly exhibit this type of chemical bonding.
- **Hydrogen Bonds -** It is a type of polar covalent bonding between oxygen and hydrogen wherein the hydrogen develops a partial positive charge

- **Polar Bonds -** In Polar Covalent chemical bonding, electrons are shared unequally since the more electronegative atom pulls the electron pair closer to itself and away from the less electronegative atom.

Factors Affecting Bond Enthalpy in Chemical Bonding:

- Size of the Atom
- Multiplicity of Bonds
- Number of Lone Pair of Electrons Present
- Bond Angle

---

27. **Answer: 4 - 4**

Explanation:

1. ClF_3 (Chlorine trifluoride)

Chlorine (Cl) has 7 valence electrons. It forms 3 bonds with F atoms → uses 3 electrons. Remaining = 7 − 3 = 4 electrons = 2 lone pairs.
→ **2 lone pairs on Cl**

2. XeO_3 (Xenon trioxide)

Xenon (Xe) has 8 valence electrons. Each oxygen forms a double bond, using 6 electrons in total (3 × 2). Used 6 electrons → 8 − 6 = 2 electrons = 1 lone pair.
→ **1 lone pair on Xe**

3. BrF_5 (Bromine pentafluoride)

Bromine (Br) has 7 valence electrons. Forms 5 bonds with F → uses 5 electrons. Remaining 7 − 5 = 2 electrons = 1 lone pair.
→ **1 lone pair on Br**

4. XeF_4 (Xenon tetrafluoride)

Xenon (Xe) has 8 valence electrons. Forms 4 bonds with F → uses 4 electrons. Remaining 8 − 4 = 4 electrons = 2 lone pairs.
→ **2 lone pairs on Xe**

5. O_3 (Ozone)

Central O has 6 valence electrons. It forms one single and one double bond (resonance). Bonds use 3 pairs = 6 electrons → 6 − 4 = 2 electrons = 1 lone pair.
→ **1 lone pair on central O**

6. NH_3 (Ammonia)

Nitrogen has 5 valence electrons. Forms 3 bonds with H → uses 3 electrons. Remaining 2 electrons = 1 lone pair.
→ **1 lone pair on N**

7. H_2O (Water)

Oxygen has 6 valence electrons. Forms 2 bonds with H → uses 2 electrons. Remaining 4 electrons = 2 lone pairs.
→ **2 lone pairs on O**

Step 5: Count of Compounds with One Lone Pair

Compounds with exactly one lone pair on central atom:

$$XeO_3,\ BrF_5,\ O_3,\ NH_3$$

Total = **4**

Final Answer:

$$\boxed{4}$$

Concepts:

1. Chemical Bonding and Molecular Structure:

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Factors Affecting Bond Enthalpy in Chemical Bonding:

- Size of the Atom
- Multiplicity of Bonds
- Number of Lone Pair of Electrons Present
- Bond Angle

---

28. Answer: b

Explanation:

To determine the number of lone pairs of electrons on the central atom for each of the given molecules, we will evaluate the electronic configuration and geometry of each molecule.

1. **SCl_2:**
   - Sulfur (S) belongs to Group 16 of the periodic table and has 6 valence electrons.
   - Each chlorine (Cl) atom, being a Group 17 element, forms a single bond with sulfur, using up 2 valence electrons from sulfur (1 for each Cl).
   - After forming two bonds with chlorine, sulfur has 4 valence electrons remaining, which are organized into 2 lone pairs.
   - The number of lone pairs on the central sulfur atom in SCl_2 is **2**.
2. **O_3 (Ozone):**
   - Ozone has a resonance structure consisting of one double bond and one single bond between oxygen atoms.
   - The central oxygen atom participates in both of these bonds.
   - Oxygen, having 6 valence electrons, after forming a double bond (using 2 electrons) and a single bond (using 1 electron), will have 3 electrons left, forming 1.5 electron pairs.
   - Typically, central oxygen can be pictured with one lone pair in most stable resonance forms.
   - The number of lone pairs on the central oxygen atom in O_3 is **1**.

3. **ClF_3:**
    - Chlorine (Cl) is the central atom with 7 valence electrons.
    - Three fluorine atoms each form a single bond with chlorine, using 3 electrons.
    - This leaves chlorine with 4 electrons, or 2 lone pairs.
    - The number of lone pairs on the central chlorine atom in ClF_3 is **2**.
4. **SF_6:**
    - Sulfur is the central atom with 6 valence electrons.
    - There are six fluorine atoms, each forming a single bond with sulfur, using up all 6 sulfur valence electrons, with no remaining electrons for lone pairs.
    - The number of lone pairs on the central sulfur atom in SF_6 is **0**.

Thus, the number of lone pairs of electrons on the central atom for SCl_2, O_3, ClF_3, and SF_6 respectively are **2, 1, 2, and 0**. The correct option is: **2, 1, 2 and 0**.

Concepts:

1. Chemical Bonding and Molecular Structure:

Such a group of atoms is called a molecule. Obviously, there must be some force that holds these constituent atoms together in the molecules. The attractive force which holds various constituents (atoms, ions, etc.) together in different chemical species is called a chemical bond.

Types of Chemical Bonds:

There are 4 types of chemical bonds which are formed by atoms or molecules to yield compounds.

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- **Polar Bonds -** In Polar Covalent chemical bonding, electrons are shared unequally since the more electronegative atom pulls the electron pair closer to itself and away from the less electronegative atom.

Factors Affecting Bond Enthalpy in Chemical Bonding:

- Size of the Atom
- Multiplicity of Bonds

- Number of Lone Pair of Electrons Present
- Bond Angle

---

29. **Answer: 1 - 1**

Explanation:

To determine which of the given sulphur-based oxoacids contain peroxo (O–O) bonds, we must analyze their structural formulas:

- H_2SO_3 **(Sulfurous acid):** This compound has the structure where sulfur is bonded to three oxygen atoms with single and double bonds, but no O–O bond is present.
- H_2SO_4 **(Sulfuric acid):** Sulfur forms double bonds with two oxygen atoms and single bonds with two hydroxyl groups. There is no O–O bond.
- $H_2S_2O_8$ **(Peroxodisulfuric acid or Marshall's acid):** This compound contains an O–O single bond. Its structural formula includes a direct O–O linkage, characteristic of peroxo bonds.
- $H_2S_2O_7$ **(Pyrosulfuric acid):** Also known as oleum, it does not contain an O–O bond; instead, sulfur atoms are linked through an oxygen atom but not forming any O–O linkage.

From the analysis, only $H_2S_2O_8$ has a peroxo bond.

Thus, the number of oxoacids with peroxo bonds is **1**. This matches the given range of (1,1).

Concepts:

1. Chemical Bonding and Molecular Structure:

Such a group of atoms is called a molecule. Obviously, there must be some force that holds these constituent atoms together in the molecules. The attractive force which holds various constituents (atoms, ions, etc.) together in different chemical species is called a chemical bond.

Types of Chemical Bonds:

There are 4 types of chemical bonds which are formed by atoms or molecules to yield compounds.

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- **Polar Bonds -** In Polar Covalent chemical bonding, electrons are shared unequally since the more electronegative atom pulls the electron pair closer to itself and away from the less electronegative atom.

Factors Affecting Bond Enthalpy in Chemical Bonding:

- Size of the Atom
- Multiplicity of Bonds
- Number of Lone Pair of Electrons Present
- Bond Angle

---

30. **Answer: 1 - 1**

Explanation:

A tetrapeptide consists of four amino acids linked in a chain. Each linkage between two amino acids is a peptide bond. Thus, in a tetrapeptide, there are three peptide bonds, as each bond links two amino acids together:

Amino Acid 1 — Peptide Bond 1 — Amino Acid 2 — Peptide Bond 2 — Amino Acid 3 — Peptide Bond 3 — Amino Acid 4.

Therefore, the correct and verified number of peptide bonds in a tetrapeptide is 3.

Concepts:

1. Chemical Bonding and Molecular Structure:

Such a group of atoms is called a molecule. Obviously, there must be some force that holds these constituent atoms together in the molecules. The attractive force which holds various constituents (atoms, ions, etc.) together in different chemical species is called a chemical bond.

Types of Chemical Bonds:

There are 4 types of chemical bonds which are formed by atoms or molecules to yield compounds.

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- **Covalent Bonds -** Compounds that contain carbon commonly exhibit this type of chemical bonding.
- **Hydrogen Bonds -** It is a type of polar covalent bonding between oxygen and hydrogen wherein the hydrogen develops a partial positive charge
- **Polar Bonds -** In Polar Covalent chemical bonding, electrons are shared unequally since the more electronegative atom pulls the electron pair closer to itself and away from the less electronegative atom.

Factors Affecting Bond Enthalpy in Chemical Bonding:

- Size of the Atom
- Multiplicity of Bonds
- Number of Lone Pair of Electrons Present
- Bond Angle