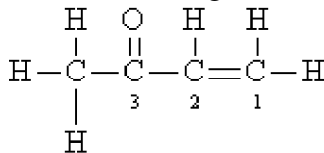


Ch. 11 Practice Questions

- The hybridization of the carbon atom in the cation CH_3^+ is:
 - sp^2
 - sp^3
 - dsp
 - sp
 - none of these
- In the molecule C_2H_4 the valence orbitals of the carbon atoms are assumed to be
 - not hybridized
 - sp hybridized
 - sp^2 hybridized
 - sp^3 hybridized
 - dsp hybridized
- The hybridization of the central atom in I_3^- is:
 - sp
 - sp^2
 - sp^3
 - dsp^3
 - d^2sp^3

4. Consider the following Lewis structure:

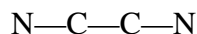


Which statement about the molecule is false?

- There are 10 sigma and 2 pi bonds.
 - C-2 is sp^2 hybridized with bond angles of 120° .
 - Oxygen is sp^3 hybridized.
 - This molecule contains 28 valence electrons.
 - There are some H-C-H bond angles of about 109° in the molecule.
5. The hybridization of the central atom, Al, in AlBr_3 is
- sp
 - sp^2
 - sp^3
 - dsp^3
 - d^2sp^3

6. In which of the compounds below is there more than one kind of hybridization (sp , sp^2 , sp^3) for carbon?
- I. $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_3$
 - II. $\text{CH}_3\text{CH}=\text{CHCH}_3$
 - III. $\text{CH}_2=\text{CH}-\text{CH}=\text{CH}_2$
 - IV. $\text{H}-\text{C}\equiv\text{C}-\text{H}$
- A) II and III
 - B) II only
 - C) III and IV
 - D) I, II, and III
 - E) III only
7. A π (pi) bond is the result of the
- A) overlap of two s orbitals
 - B) overlap of an s orbital and a p orbital
 - C) overlap of two p orbitals along their axes
 - D) sidewise overlap of two parallel p orbitals
 - E) sidewise overlap of two s orbitals

Consider the skeletal structure shown below:



Draw the Lewis structure and answer the following:

8. How many of the atoms are sp hybridized?
- A) 0
 - B) 1
 - C) 2
 - D) 3
 - E) 4
9. How many pi bonds does the molecule contain?
- A) 0
 - B) 2
 - C) 4
 - D) 6
 - E) 7
10. Which of the following molecules contains the shortest C–C bond?
- A) C_2H_2
 - B) C_2H_4
 - C) C_2H_6
 - D) C_2Cl_4
 - E) b and d

11. The electron configuration of a particular diatomic species is $(\sigma_{2s})^2(\sigma_{2s}^*)^2(\sigma_{2p})^2(\pi_{2p})^4(\pi_{2p}^*)^2$. What is the bond order for this species?
- 3.5
 - 3
 - 2.5
 - 2
 - 1.5
12. What is the bond order of He_2^+ ?
- 0
 - $\frac{1}{2}$
 - 1
 - $1\frac{1}{2}$
 - 2
13. If four orbitals on one atom overlap four orbitals on a second atom, how many molecular orbitals will form?
- 1
 - 4
 - 8
 - 16
 - none of these
14. When comparing Be_2 and H_2 :
- Be_2 is more stable because it contains both bonding and antibonding valence electrons.
 - H_2 has a higher bond order than Be_2 .
 - H_2 is more stable because it only contains σ_{1s} electrons.
 - H_2 is more stable because it is diamagnetic, whereas Be_2 is paramagnetic.
- I, II
 - III only
 - II, III
 - II, III, IV
 - III, IV
15. The configuration $(\sigma_{2s})^2(\sigma_{2s}^*)^2(\pi_{2py})^1(\pi_{2px})^1$ is the molecular orbital description for the ground state of
- Li_2^+
 - Be_2
 - B_2
 - B_2^{2-}
 - C_2

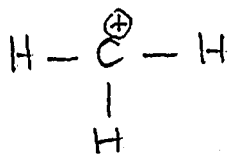
16. Which of the following species is paramagnetic?
- A) C_2
 - B) O_2
 - C) F_2
 - D) Li_2
 - E) none of these
17. Which of the following species has the largest dissociation energy?
- A) O_2
 - B) O_2^-
 - C) O_2^{2-}
 - D) O_2^+
 - E) O_2^{2+}
18. Which of the following diatomic molecules has a bond order of 2?
- A) B_2
 - B) C_2
 - C) P_2
 - D) F_2
 - E) Li_2
19. Consider the molecular orbital energy level diagrams for O_2 and NO . Which of the following is true?
- I. Both molecules are paramagnetic.
 - II. The bond strength of O_2 is greater than the bond strength of NO .
 - III. NO is an example of a homonuclear diatomic molecule.
 - IV. The ionization energy of NO is smaller than the ionization energy of NO^+ .
- A) I only
 - B) I and II
 - C) I and IV
 - D) II and III
 - E) I, II, and IV
20. The hybridization of the lead atom in $PbCl_4$ is
- A) dsp^2
 - B) sp^2
 - C) d^2sp^3
 - D) dsp^3
 - E) none of these

Answer Section

1. ANS: A
2. ANS: C
3. ANS: D
4. ANS: C
5. ANS: B
6. ANS: B
7. ANS: D
8. ANS: E
9. ANS: C
10. ANS: A
11. ANS: D
12. ANS: B
13. ANS: C
14. ANS: C
15. ANS: C
16. ANS: B
17. ANS: E
18. ANS: B
19. ANS: C
20. ANS: E

Solutions to Ch. 9 Practice Questions

① Lewis structure of CH_3^+ is



Note: C is violating the octet rule in this cation.

It happens when organic molecules form what's called a "carbocation".

Why did we put the (+) on C?

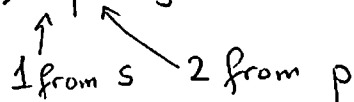
Because we can't put it on any H.

And we know C doesn't have a lone pair, using our shortcut:

$$\# \text{ of lone pair } e^- = (\text{no. of valence electrons}) - (\text{no. of bonds}) - (\text{formal charge})$$
$$= 4 - 3 - 1 = 0$$

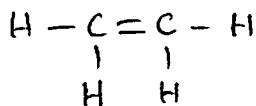
$$\Rightarrow \# \text{ of lone pairs} = \frac{0}{2} = 0$$

Therefore C has only 3 electron pairs $\Rightarrow sp^2$ hybridized



② The only way to write the skeletal structure is $\text{H} - \underset{\text{H}}{\text{C}} - \underset{\text{H}}{\text{C}} - \text{H}$ because hydrogens must be peripheral

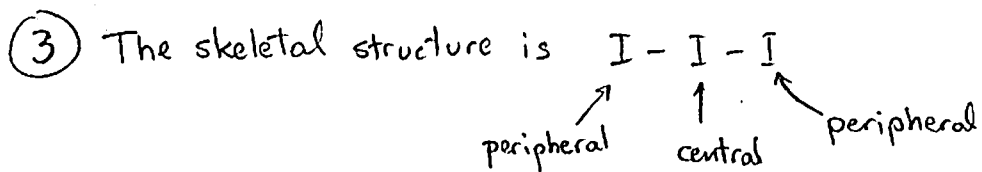
Carbon needs to make 4 bonds to avoid a (+) charge. That can be easily accomplished by simply adding a bond between the two carbons:



The second bond in the double bond is a π bond, and does not affect hybridization, which is a property of only the σ bonds and lone pairs.

Both carbon atoms have 3 electron groups (2 of the groups are simple σ bonds, and the third group is a σ sigma bond with a π bond tagging along (a double-bond)).

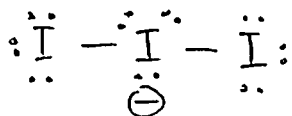
3 groups of $e^- \Rightarrow sp^2$ hybridized



Peripheral halogens make one bond, and act like "normal" halogens, following the octet rule. A peripheral halogen with a single bond has a zero formal charge. That leaves the central I as the only atom that can carry the (-) charge. Central halogens routinely violate the octet rule and thus carry a charge and make multiple bonds.

Using our shortcut: $\left(\begin{array}{c} \text{no. of} \\ \text{lone pair} \\ \text{electrons} \end{array} \right) = 7 - 2 - (-1) = 6 \Rightarrow \left(\begin{array}{c} \text{no. of} \\ \text{lone} \\ \text{pairs} \end{array} \right) = \frac{6}{2} = 3$

Thus we have:



The central I has 3 lone pairs and 2 bonds, for a total of 5 e^- groups.

$$5 e^- \text{ groups} \Rightarrow sp^3d$$

$$\begin{array}{c} \uparrow \uparrow \uparrow \\ 1 + 3 + 1 = 5 \end{array}$$

Since s and p orbitals can only create a maximum of $1 + 3 = 4$ hybrid orbitals, we need to bring in an extra d-orbital to the mix.

④ To count the no. of sigma bonds, we just count all the bonds regardless of their being single, double, or triple. A double bond has one σ bond and one π bond, so it contributes one σ bond. A triple bond again has only one σ bond and two π bonds, again contributing one σ bond to the total. There are 10 sigma bonds 2 π bonds (one π bond from each double bond).

C-2 has 3 e^- groups $\Rightarrow sp^2$ hybridized \Rightarrow bond angles = 120°

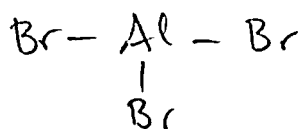
The oxygen has a double bond. To calculate the number of e^- groups around it, we need to know the no. of lone pairs. Since the double bond uses 2 e^- pairs (4 electrons) it leaves 4 e^- for lone pairs, which means the oxygen has $\frac{4}{2} = 2$ lone pairs, for a total of 3 e^- groups $\Rightarrow sp^2$

Oxygen has sp^2 hybridization, not sp^3

The molecule contains $\underbrace{(4)(4)}_C + \underbrace{(6)(1)}_H + \underbrace{(1)(6)}_O = 28$ valence e^-

The carbon in $\begin{array}{c} \text{H} \\ | \\ \text{H} - \text{C} - \\ | \\ \text{H} \end{array}$ has sp^3 hybridization (4 e^- groups), and bond angles of 109°

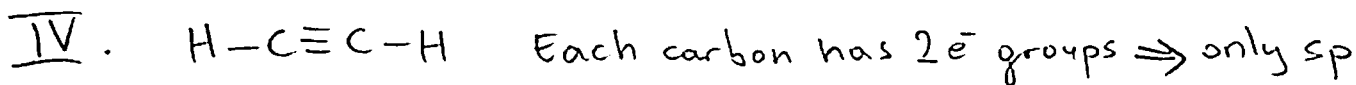
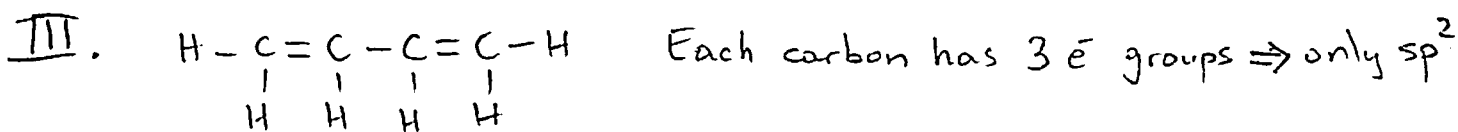
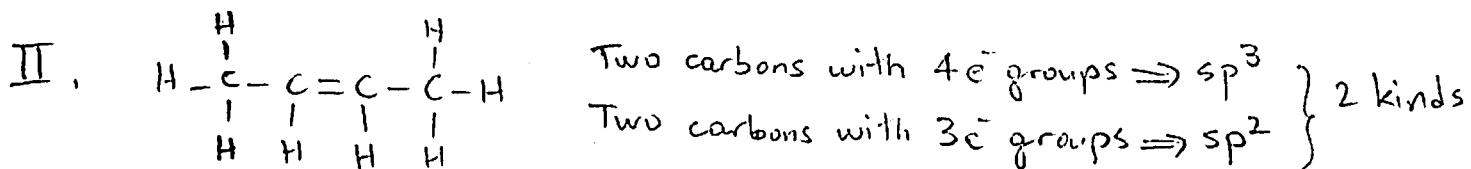
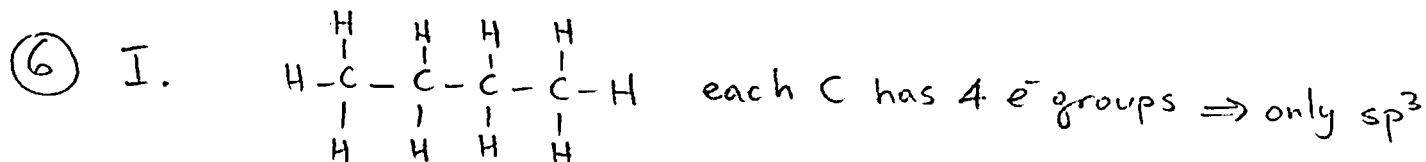
- (5) Al, like B, violates the octet rule and is "satisfied" with only six electrons. Therefore, Al in AlBr_3 does not need a lone pair:



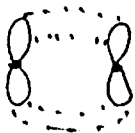
Our shortcut for lone pairs predicts the violation of the octet rule.

of lone pair $e^- = 3 - 3 - 0 = 0 \Rightarrow$ no lone pairs
We used a charge of 0 because the peripheral Br with their single bonds are "happy" and have no charge, and the molecule is neutral, so Al has zero formal charge.

Al has 3 e^- groups $\Rightarrow sp^2$ hybridized



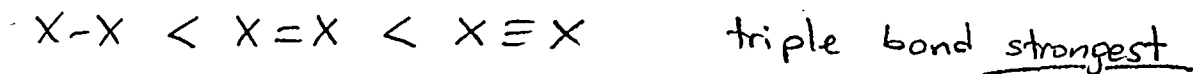
- (7) A π bond is the result of the sidewise overlap of two parallel p orbitals



- (8) The skeletal structure contains all the σ bonds. Since carbon cannot exceed the octet it is limited to hybrids without d orbitals. It is very rare for C to have lone pairs, and we should consider it only if there is no other way to satisfy the octets. Remember that C normally makes 4 bonds, and N makes 3. Let's see if we can satisfy the octets using them.
 $:\text{N} \equiv \text{C} - \text{C} \equiv \text{N}:$ Indeed the octets are satisfied, and all the atoms have 2 electron groups. So 4 atoms have sp hybridization.

⑨ $:N \equiv C - C \equiv N:$ contains 2 triple bonds. Each triple bond contains two π -bonds. Number of π bonds = $2 \times 2 = 4$

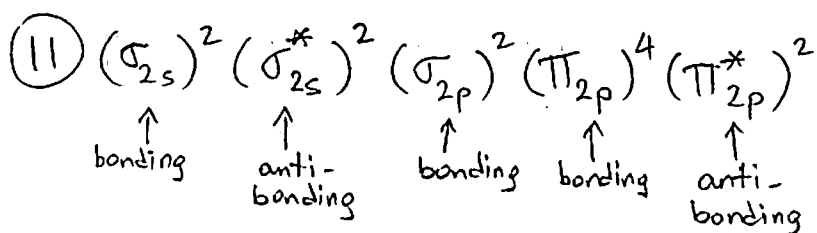
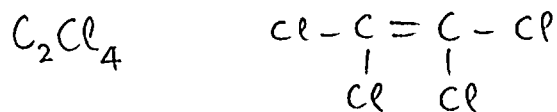
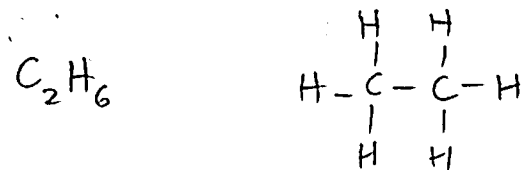
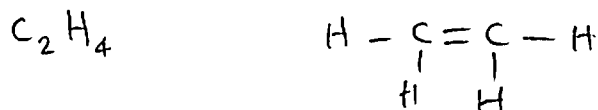
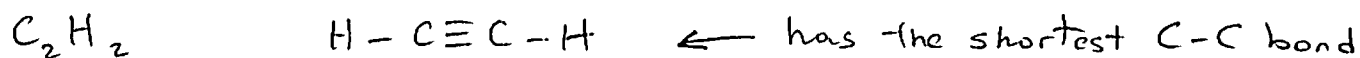
⑩ For bonds between two given atoms, the strength of the bond increases in the following order



and the length of the bond goes in the reverse order



The molecule with the triple bond contains the shortest C-C bond. We write the Lewis structures of each choice. We can use the long, standard procedure or use our rules of thumb (4 bonds for C, 2 bonds for O, 1 bond for H or peripheral halogens; use lone pairs or multiple bonds as needed).

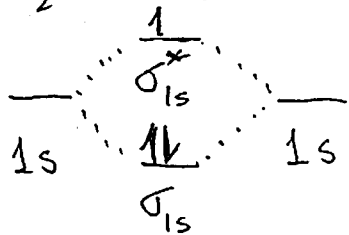


of e^- in bonding orbitals = $2 + 2 + 4 = 8$

of e^- in anti-bonding orbitals = $2 + 2 = 4$

Bond order = $\frac{8-4}{2} = 2$

(12) He_2^+ has 3 valence electrons. Two from each He, minus one e^- lost.



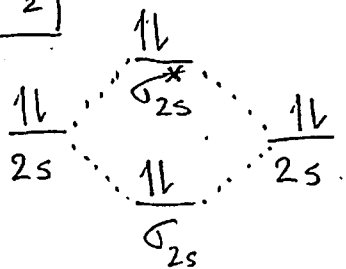
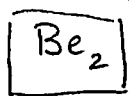
He_2^+ has a bond order of $\frac{1}{2}$, which is simply given by

$$\frac{1}{2}(\text{no. of bonding } e^- - \text{no. of antibonding } e^-)$$

(13) Number of orbitals is conserved. The total number of atomic orbitals is equal to the number of molecular orbitals they form.

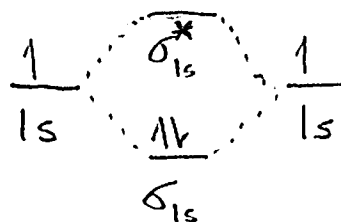
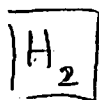
$$4 + 4 = 8 \text{ molecular orbitals}$$

(14) We need the molecular orbital diagrams for Be_2 and H_2 , and consider only the valence orbitals.



$$\text{bond order} = \frac{2 - 2}{2} = 0$$

unstable



$$\text{bond order} = \frac{2 - 0}{2} = 1$$

Because it has a bond order of zero, Be_2 is unstable. H_2 is more stable than Be_2 because it has a bond order of 1. Its stability has nothing to do with being diamagnetic. If H_2 had an e^- in σ_{1s}^* instead of having both its electrons in σ_{1s} , it would not be stable.

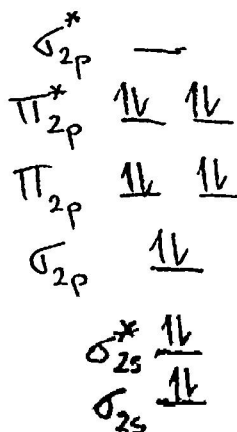
(15) While we can spend some time constructing the MO diagrams of all the species listed and compare it to the given one, it's wise to simply count the no. of e^- in the given MO configuration and see if we can eliminate some or all of the wrong choices.

$$(\sigma_{2s})^2 (\sigma_{2s}^*)^2 (\pi_{2p_y})^1 (\pi_{2p_x})^1 \text{ has } 2 + 2 + 1 + 1 = 6 \text{ valence } e^-$$

$$\text{Li}_2^+ \text{ has } (2)(1) - 1 = 1e^-, \text{ Be}_2 \text{ has } (2)(2) = 4e^-, \text{ B}_2 \text{ has } (2)(3) = 6e^-,$$

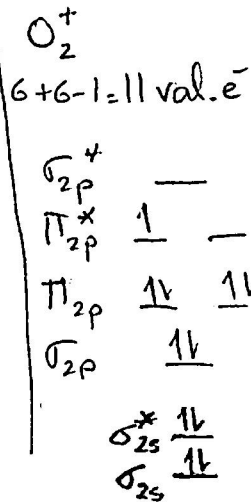
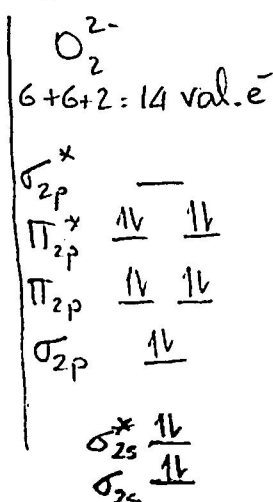
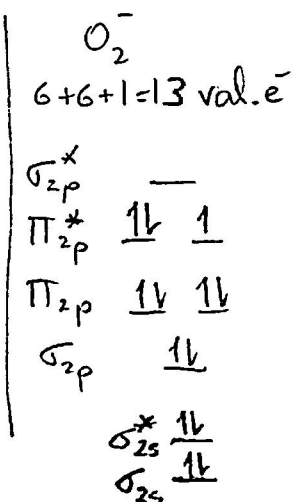
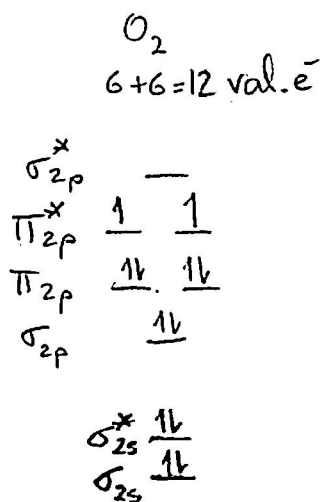
$$\text{B}_2^- \text{ has } (2)(3) + 2 = 8e^-, \text{ C}_2 \text{ has } (2)(4) = 8e^-$$

(16 cont.) Two F atoms bring 7 valence electrons each (for a total of 14)



No unpaired $e^- \Rightarrow$ diamagnetic

(17) For a given pair of atoms, the highest bond order corresponds to the strongest bond, and therefore the largest dissociation energy.



$$\text{Bond order} = \frac{8-4}{2} = 2$$

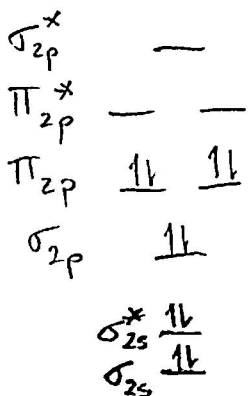
$$\text{Bond order} = \frac{8-5}{2} = 1.5$$

$$\text{Bond order} = \frac{8-6}{2} = 1$$

$$\text{Bond order} = \frac{8-3}{2} = 2.5$$



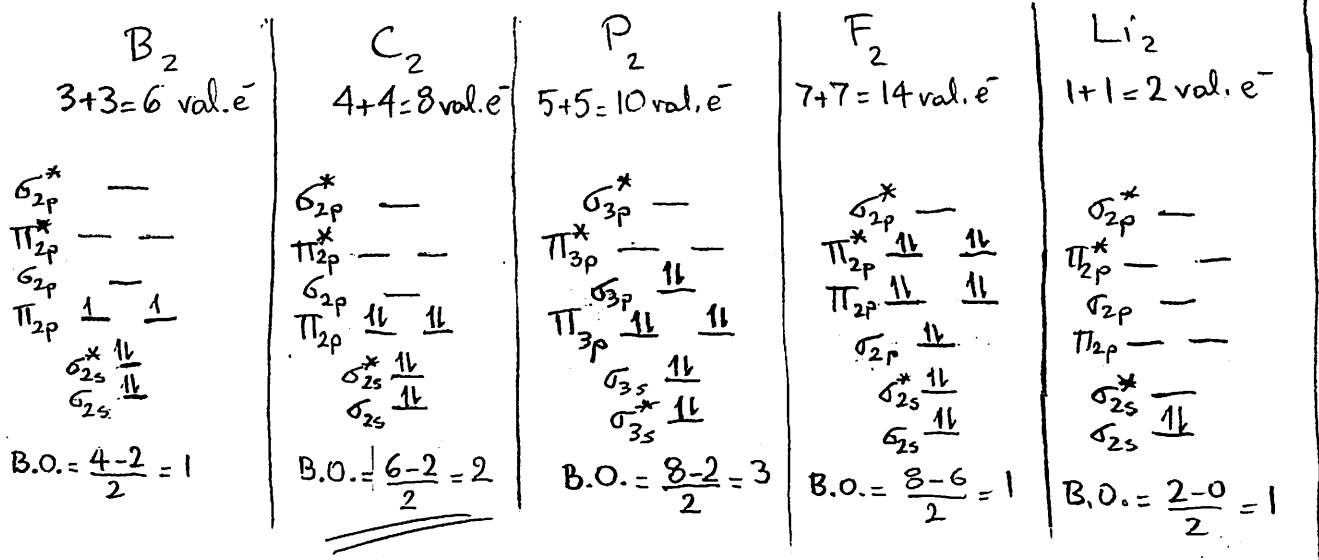
$$6+6-2=10 \text{ val. } e^-$$



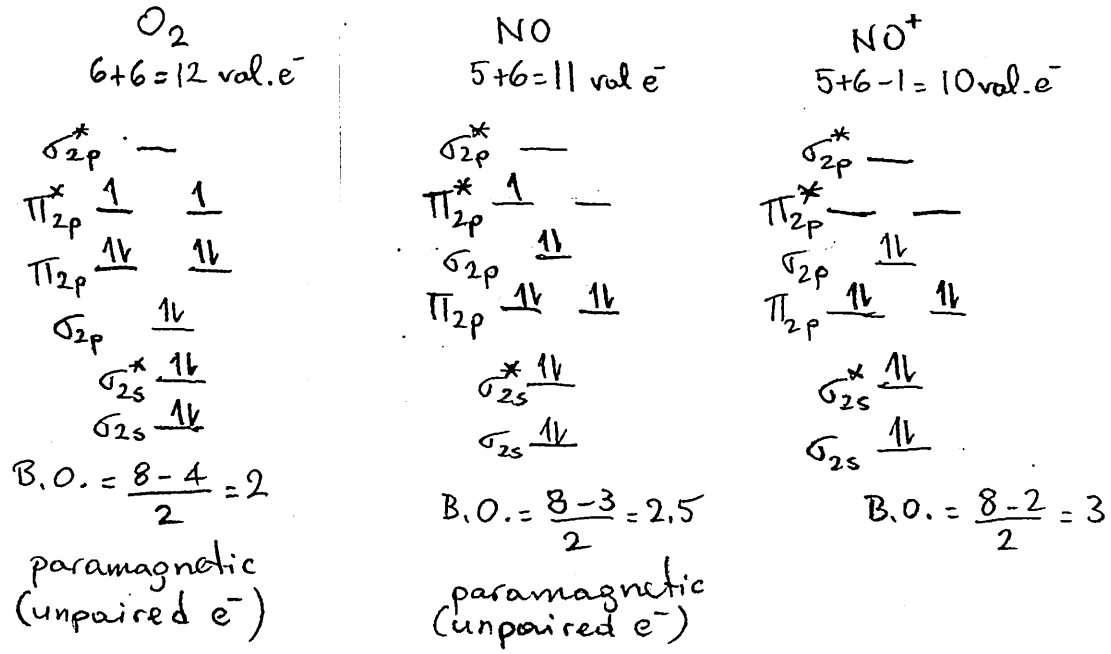
$$\text{Bond order} = \frac{8-2}{2} = 3 \quad \leftarrow \text{largest bond order}$$

O_2^{2+} has the largest bond dissociation energy

18) We construct the M.O. energy diagrams for all the choices



19) O_2 has the "expected" M.O. diagram, while NO has the modified diagram like N_2



- * Both O_2 and NO are paramagnetic
- * Bond strength of O_2 is not greater than that of NO
 - While it's not totally safe to use B.O. to compare the bond strengths of diatomic species made of different atoms, O_2 and NO are close enough
- * NO is not an example of a homonuclear diatomic molecule
 - It's heteronuclear
- * The ionization energy of NO is smaller than that of NO^+
 - The higher the energy of the highest-energy e^- , the smaller the ionization energy. The highest energy e^- in NO is in π_{2p}^* , whereas that of NO^+ is in the lower-energy σ_{2p} M.O.

(20) Although PbCl_4 is a metal-nonmetal compound, it actually has covalent bonds, as you must surmise from the context (although "none of these" would be appropriate if it were an ionic compound). A statement like " PbCl_4 is a yellow, oily liquid which is stable below 0°C and decomposes at 50°C " at the beginning of a question would give you enough hint to conclude that the compound is not ionic and you can use the covalent bond concepts like hybridization. (Simple binary ionic compounds have high melting points and don't decompose easily).

Since the peripheral atoms are halogens (expected to have single bonds and zero formal charges, making the formal charge of Pb also zero, because PbCl_4 is a neutral molecule), we can use our shortcut for finding the no. of lone pairs on the central atom.

$$\left(\begin{array}{l} \text{no. of} \\ \text{lone pair } e^- \end{array}\right) = \left(\begin{array}{l} \text{no. of} \\ \text{valence } e^- \end{array}\right) - \left(\begin{array}{l} \text{no. of} \\ \text{bonds} \end{array}\right) - \left(\begin{array}{l} \text{formal} \\ \text{charge} \end{array}\right) = 4 - 4 - 0 = 0$$

No lone pairs, and 4 bonds \Rightarrow 4 e^- groups \Rightarrow sp^3 hybridization

So the correct answer indeed is "none of these".

However, I don't know if the writers of the question meant that because they (understandably) thought that it had ionic bonds and hybridization didn't apply, or actually understood that it had sp^3 hybridization. Lucky for the person who has to answer the question, both trains of thought lead to "none of these".