Ch. 9 Practice Questions

1. The hybridization of the carbon atom in the cation \( \text{CH}_3^+ \) is:
   A) \( sp^2 \)
   B) \( sp^3 \)
   C) \( dsp \)
   D) \( sp \)
   E) none of these

2. In the molecule \( \text{C}_2\text{H}_4 \) the valence orbitals of the carbon atoms are assumed to be
   A) not hybridized
   B) \( sp \) hybridized
   C) \( sp^2 \) hybridized
   D) \( sp^3 \) hybridized
   E) \( dsp \) hybridized

3. The hybridization of the central atom in \( \text{I}_3^- \) is:
   A) \( sp \)
   B) \( sp^2 \)
   C) \( sp^3 \)
   D) \( dsp^3 \)
   E) \( d^2sp^3 \)

4. Consider the following Lewis structure:
   \[
   \begin{array}{c}
   \text{H} \quad \text{O} \quad \text{H} \quad \text{H} \\
   \text{H} \quad \text{C} \quad \text{C} \quad \text{C} \quad \text{H} \\
   \text{H} \\
   \end{array}
   \]
   Which statement about the molecule is false?
   A) There are 10 sigma and 2 pi bonds.
   B) \( \text{C}-2 \) is \( sp^2 \) hybridized with bond angles of 120°.
   C) Oxygen is \( sp^3 \) hybridized.
   D) This molecule contains 28 valence electrons.
   E) There are some \( \text{H}–\text{C}–\text{H} \) bond angles of about 109° in the molecule.

5. The hybridization of the central atom, Al, in \( \text{AlBr}_3 \) is
   A) \( sp \)
   B) \( sp^2 \)
   C) \( sp^3 \)
   D) \( dsp^3 \)
   E) \( d^2sp^3 \)
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. \( H – C \equiv C – H \)

A) II and III  
B) II only  
C) III and IV  
D) I, II, and III  
E) III only

7. A \( \pi \) (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:
\[ \text{N—C—C—N} \]

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) \( \text{C}_2\text{H}_2 \)  
B) \( \text{C}_2\text{H}_4 \)  
C) \( \text{C}_2\text{H}_6 \)  
D) \( \text{C}_2\text{Cl}_4 \)  
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?
   A) 3.5
   B) 3
   C) 2.5
   D) 2
   E) 1.5

12. What is the bond order of \(\text{He}_2^+\)?
   A) 0
   B) \(\frac{1}{2}\)
   C) 1
   D) \(1\frac{1}{2}\)
   E) 2

13. If four orbitals on one atom overlap four orbitals on a second atom, how many molecular orbitals will form?
   A) 1
   B) 4
   C) 8
   D) 16
   E) none of these

14. When comparing \(\text{Be}_2\) and \(\text{H}_2\):
   I. \(\text{Be}_2\) is more stable because it contains both bonding and antibonding valence electrons.
   II. \(\text{H}_2\) has a higher bond order than \(\text{Be}_2\).
   III. \(\text{H}_2\) is more stable because it only contains \(\sigma_{1s}\) electrons.
   IV. \(\text{H}_2\) is more stable because it is diamagnetic, whereas \(\text{Be}_2\) is paramagnetic.
   A) I, II
   B) III only
   C) II, III
   D) II, III, IV
   E) III, IV

15. The configuration \((\sigma_{2s})^2(\sigma_{2s}^*)^2(\pi_{2p_y})^1(\pi_{2p_x})^1\) is the molecular orbital description for the ground state of
   A) \(\text{Li}_2^+\)
   B) \(\text{Be}_2\)
   C) \(\text{B}_2\)
   D) \(\text{B}_2^{2-}\)
   E) \(\text{C}_2\)
16. Which of the following species is paramagnetic?
   A) C₂
   B) O₂
   C) F₂
   D) Li₂
   E) none of these

17. Which of the following species has the largest dissociation energy?
   A) O₂
   B) O₂⁻
   C) O₂²⁻
   D) O₂⁺
   E) O₂²⁺

18. Which of the following diatomic molecules has a bond order of 2?
   A) B₂
   B) C₂
   C) P₂
   D) F₂
   E) Li₂

19. Consider the molecular orbital energy level diagrams for O₂ and NO. Which of the following is true?
   I. Both molecules are paramagnetic.
   II. The bond strength of O₂ 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 PbCl₄ is
   A) dsp²
   B) sp²
   C) d²sp³
   D) dsp³
   E) none of these
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Solutions to Ch. 9 Practice Questions

1) Lewis structure of $\text{CH}_3^+$ is

\[
\begin{array}{c}
\text{H} - \overset{\circ}{\text{C}} - \text{H} \\
\text{H}
\end{array}
\]

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:

\[
\text{# of lone pair } e^- = \left( \frac{\text{no. of valence electrons}}{2} \right) - \left( \frac{\text{no. of bonds}}{2} \right) - \left( \frac{\text{formal charge}}{2} \right)
\]

\[
= 4 - 3 - 1 = 0
\]

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

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

\[\uparrow \text{ from s} \quad 2 \text{ from p}\]

Note: C is violating the octet rule in this cation. It happens when organic molecules form what’s called a “carbocation”.

2) The only way to write the skeletal structure is $\text{H} - \overset{\circ}{\text{C}} - \overset{\circ}{\text{C}} - \text{H}$ 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:

\[
\begin{array}{c}
\text{H} - \overset{\circ}{\text{C}} = \overset{\circ}{\text{C}} - \text{H} \\
\text{H} \\
\text{H}
\end{array}
\]

The second bond in the double bond is a \( \pi \) bond, and does not affect hybridization, which is a property of only the \( \sigma \) bonds and lone pairs.

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

3 groups of $e^-$ \( \Rightarrow \) $sp^2$ hybridized
The skeletal structure is I – I – I.

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:

\[
\text{no. of lone pairs} = 7-2-(−1) = 6 \Rightarrow \frac{\text{no. of lone pairs}}{2} = \frac{6}{2} = 3
\]

Thus we have:

\[\text{I} – \text{I} – \text{I}\]

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

\[5 \text{ e}^− \text{groups} \Rightarrow \text{sp}^3 \text{d hybridization} \]

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.

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 \(\Rightarrow \text{sp}^2\) 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 \(\Rightarrow \text{sp}^2\) hybridization, not \(\text{sp}^3\).

The molecule contains \((4)(4) + (6)(1) + (1)(6) = 28\) valence e−.

The carbon in \(\text{H}_2\text{C}=\text{O}\) has \(\text{sp}^3\) 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 AlBr₃ does not need a lone pair:

\[ \text{Br} \quad \text{Al} \quad \text{Br} \]

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

\# of lone pair e⁻ = 3 - 3 - 0 = 0 \Rightarrow \text{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 \( \Rightarrow \text{sp}^2 \) hybridized

(6) I. \[ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} - \text{H} \]

Each C has 4 e⁻ groups \( \Rightarrow \text{only sp}^3 \)

II. \[ \text{H} - \text{C} - \text{C} = \text{C} - \text{C} - \text{H} \]

Two carbons with 4 e⁻ groups \( \Rightarrow \text{sp}^3 \) \( \{ \) 2 kinds

Two carbons with 3 e⁻ groups \( \Rightarrow \text{sp}^2 \) \( \} \)

III. \[ \text{H} - \text{C} = \text{C} - \text{C} = \text{C} - \text{H} \]

Each carbon has 3 e⁻ groups \( \Rightarrow \text{only sp}^2 \)

IV. \[ \text{H} - \text{C} = \text{C} - \text{H} \]

Each carbon has 2 e⁻ groups \( \Rightarrow \text{only sp} \)

(7) A \( \sigma \) bond is the result of the sidewise overlap of two parallel orbitals.

(8) The skeletal structure contains all the \( \sigma \) 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.

\( :\text{N} = \text{C} - \text{C} = \text{N} : \) Indeed the octets are satisfied, and all the atoms have 2 electron groups. So 4 atoms have sp hybridization.
9) \(\text{N} = \text{C} = \text{C} = \text{N}\): contains 2 triple bonds. Each triple bond contains two \(\pi\)-bonds. Number of \(\pi\) bonds = \(2 \times 2 = 4\)

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

\[ X - X < X = X < X = X \text{ triple bond strongest} \]

and the length of the bond goes in the reverse order:

\[ X - X > X = X > X = X \text{ triple bond shortest} \]

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).

\[
\begin{align*}
\text{C}_2\text{H}_2 & \quad \text{H} - \text{C} = \text{C} - \text{H} \quad \text{has the shortest C-C bond} \\
\text{C}_2\text{H}_4 & \quad \text{H} - \text{C} = \text{C} = \text{H} \\
\text{C}_2\text{H}_6 & \quad \text{H} - \text{C} - \text{C} - \text{H} \\
\text{C}_2\text{Cl}_4 & \quad \text{Cl} - \text{C} = \text{C} - \text{Cl} \\
\end{align*}
\]

11) \((\sigma_{2s})^2(\sigma_{2s}^*)^2(\sigma_{2p})^2(\pi_{2p})^4(\pi_{2p}^*)^2\)

\# of \(e^-\) in bonding orbitals = \(2 + 2 + 4 = 8\)

\# of \(e^-\) in anti-bonding orbitals = \(2 + 2 = 4\)

Bond order = \(\frac{8 - 4}{2} = 2\)
He\textsuperscript{+} has 3 valence electrons. Two from each He, minus one e\textsuperscript{-} lost.

He\textsubscript{2}\superscript{+} has a bond order of \( \frac{1}{2} \), which is simply given by
\[
\frac{1}{2} (\text{no. of bonding e}^- - \text{no. of antibonding e}^-)
\]

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

4 + 4 = 8 molecular orbitals

We need the molecular orbital diagrams for Be\textsubscript{2} and H\textsubscript{2}, and consider only the valence orbitals.

**Be\textsubscript{2}**

\[
\sigma\textsubscript{2s} \quad \sigma\textsubscript{2s}^* \quad \Pi\textsubscript{2p}\textsuperscript{y}
\]

bond order = \( \frac{2-2}{2} = 0 \)

unstable

Because it has a bond order of zero, Be\textsubscript{2} is unstable. H\textsubscript{2} is more stable than Be\textsubscript{2} because it has a bond order of 1. Its stability has nothing to do with being diamagnetic. If H\textsubscript{2} had an e\textsuperscript{-} in \( \sigma\textsubscript{1s}^* \) instead of having both its electrons in \( \sigma\textsubscript{1s} \), it would not be stable.

**H\textsubscript{2}**

\[
\sigma\textsubscript{1s} \quad \sigma\textsubscript{1s}^* \quad \Pi\textsubscript{1s}\textsuperscript{1s}
\]

bond order = \( \frac{2-0}{2} = 1 \)

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\textsuperscript{-} in the given MO configuration and see if we can eliminate some or all of the wrong choices.

\[(\sigma\textsubscript{2s})^2(\sigma\textsubscript{2s}^*)^2(\Pi\textsubscript{2p}\textsuperscript{y})^1(\Pi\textsubscript{2p}\textsuperscript{x})^1 \text{ has } 2 + 2 + 1 + 1 = 6 \text{ valence e}^-\]

Li\textsuperscript{2+} has (2)(1) = 1 e\textsuperscript{-}, Be\textsubscript{2} has (2)(2) = 4 e\textsuperscript{-}, B\textsubscript{2} has (2)(3) = 6 e\textsuperscript{-}, B\textsubscript{2} has (2)(3) = 8 e\textsuperscript{-}, C\textsubscript{2} has (2)(4) = 8 e\textsuperscript{-}
For demonstration purposes, let's construct the $B_2$ molecular orbitals.

\[ \sigma_{2p}^* \quad \Pi_{2p}^{x} \quad \Pi_{2p}^{y} \]

\[ \begin{array}{c}
2p & 1 & - & - \\
1 & \Pi_{2p}^{x} & \Pi_{2p}^{y} & 1 & - & 2p
\end{array} \]

\[ \begin{array}{c}
2s & 1l & \downarrow & \Pi_{2p}^{x} & \Pi_{2p}^{y} & 1l & 2s
\end{array} \]

\[ \sigma_{2s} \]

\[ (\sigma_{2s})^2 (\sigma_{2s})^2 (\Pi_{2p}^{x})^1 (\Pi_{2p}^{y})^1 \]

we could also lump all $\Pi_{2p}$ orbitals into one term, $\Pi_{2p}$, in which case it would be $\Pi_{2p}^2$.

Since the orbitals formed from orbitals in filled subshells will have paired electrons, we will focus on the MO formed from the highest unfilled (if any) subshells of the atoms.

$C_2$

\[ \begin{array}{c}
\sigma_{2p}^* \\
\Pi_{2p} \\
1 & 1 & - \\
2p & 1 & - & 2p
\end{array} \]

all $e^-$ in MO paired

$\Rightarrow$ diamagnetic

$O_2$

\[ \begin{array}{c}
\sigma_{2p}^* \\
\Pi_{2p} \\
1 & 1 & - \\
2p & 1 & - & 2p
\end{array} \]

2 unpaired $e^-$ in $\Pi_{2p}^*$

$\Rightarrow$ paramagnetic
Let's use a more compact MO diagram for $F_2$ (which has the same ordering as $O$). Two $F$ atoms bring 5 $p$ electrons each (for a total of 10).

$\sigma_{2p}^* \quad 1\nu \quad 1\nu$

$\Pi_{2p}^* \quad 1\nu \quad 1\nu$

$\Pi_{2p} \quad 1\nu \quad 1\nu$

$\sigma_{2p} \quad 1\nu$

No unpaired $e^- \Rightarrow$ diamagnetic

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

\[
\begin{array}{c|c|c|c}
\text{Bond order} & \frac{6-2}{2} = 2 & \frac{6-3}{2} = 1.5 & \frac{6-1}{2} = 2.5 \\
\sigma_{2p}^* & 1 \nu \quad 1\nu & 1\nu \quad 1\nu & 1\nu \\
\Pi_{2p}^* & 1\nu \quad 1\nu & 1\nu \quad 1\nu & 1\nu \\
\Pi_{2p} & 1\nu \quad 1\nu & 1\nu \quad 1\nu & 1\nu \\
\sigma_{2p} & 1\nu & 1\nu & 1\nu \\
\end{array}
\]

\[
\begin{align*}
O_2^+ \\
4 + 4 - 2 = 6 \; p \; e^- \\
\sigma_{2p}^* & - \\
\Pi_{2p}^* & - \\
\Pi_{2p} & 1\nu \quad 1\nu \\
\sigma_{2p} & 1\nu \\
\text{Bond order} & = \frac{6-0}{2} = 3 \quad \leftarrow \text{largest bond order}
\end{align*}
\]

$O_2^+$ has the largest bond dissociation energy.
18. We just construct the MO energy diagrams for all the choices

\begin{align*}
B_2 & \quad C_2 & \quad P_2 & \quad F_2 & \quad Li_2 \\
\begin{array}{ccc}
\sigma_{2p} & - \\
\Pi_{2p} & 1 \\
\end{array} & & & & & & & & &
\begin{array}{ccc}
\sigma_{2p}^* & - \\
\Pi_{2p} & 1 \\
\end{array} & & & & & & & & &
\begin{array}{ccc}
\sigma_{2p} & - \\
\Pi_{2p} & 1 \\
\end{array} & & & & & & & & &
\begin{array}{ccc}
\sigma_{2p} & 1 \\
\Pi_{2p} & 1 \\
\end{array} & & & & & & & & &
\begin{array}{ccc}
\sigma_{2p}^* & 1 \\
\Pi_{2p} & 1 \\
\end{array} & & & & & & & & &
\begin{array}{ccc}
\sigma_{2s} & 1 \\
\Pi_{2p} & 1 \\
\end{array} & & & & & & & & &
\begin{array}{ccc}
\sigma_{2s} & 1 \\
\Pi_{2p} & 1 \\
\end{array} & & & & & & & & &

\text{B.O.} = \frac{2-0}{2} = 1 & \quad \text{B.O.} = \frac{4-0}{2} = 2 & \quad \text{B.O.} = \frac{6-6}{2} = 3 \\
\text{B.O.} = \frac{6-4}{2} = 2 & \quad \text{B.O.} = \frac{6-2}{2} = 2
\end{align*}

\text{C}_2 \text{ has a bond order of 2}

19. \(O_2\) has the "expected" energy diagram, while \(NO\) has a modified diagram like \(N_2\)

\begin{align*}
O_2 & \quad NO \\
4+4 & = 8 \text{ p.e}^- & 3+4 & = 7 \text{ p.e}^- \\
\begin{array}{ccc}
\sigma_{2p} & - \\
\Pi^*_{2p} & 1 \\
\Pi_{2p} & 1 \\
\sigma_{2p} & 1 \\
\end{array} & & & & & & & & &
\begin{array}{ccc}
\sigma_{2p}^* & - \\
\Pi^*_{2p} & 1 \\
\Pi_{2p} & 1 \\
\sigma_{2p} & 1 \\
\end{array} & & & & & & & & &
\begin{array}{ccc}
\sigma_{3p} & 1 \\
\Pi^*_{2p} & 1 \\
\Pi_{2p} & 1 \\
\sigma_{2p} & 1 \\
\end{array} & & & & & & & & &
\begin{array}{ccc}
\sigma_{2p} & 1 \\
\Pi_{2p} & 1 \\
\end{array} & & & & & & & & &

\text{B.O.} = \frac{6-2}{2} = 2 & \quad \text{B.O.} = \frac{6-1}{2} = 2.5
\end{align*}

\(\text{paramagnetic (unpaired } \epsilon^-)\) \\
\(\text{paramagnetic (unpaired } \epsilon^-)\)

Ⅰ. Both molecules are paramagnetic \((\text{TRUE})\)

Ⅱ. Bond strength of \(O_2\) is greater than the bond strength of \(NO\) \((\text{FALSE})\)

Ⅲ. \(NO\) is an example of a \textit{homonuclear} diatomic molecule \((\text{FALSE})\)

Ⅳ. The ionization energy of \(NO\) is smaller than the ionization energy of \(NO^+\) \((\text{TRUE})\)
Although \( \text{PbCl}_4 \) 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 "\( \text{PbCl}_4 \) 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 \( \text{PbCl}_4 \) is a neutral molecule), we can use our shortcut for finding the no. of lone pairs on the central atom.

\[
\text{(no. of lone pair e\textsuperscript{-}) = (no. of valence e\textsuperscript{-}) - (no. of bonds) - (formal charge) = 4 - 4 - 0 = 0}
\]

No lone pairs, and 4 bonds \( \Rightarrow \) 4 e\textsuperscript{-} groups \( \Rightarrow \) \( \text{sp}^3 \) 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 \( \text{sp}^3 \) hybridization. Lucky for the person who has to answer the question, both trains of thought lead to "none of these".