This print-out should have 25 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering.
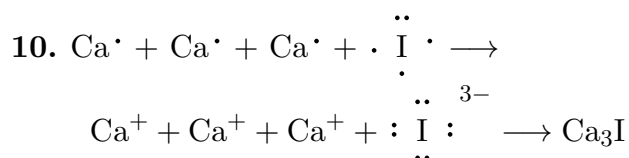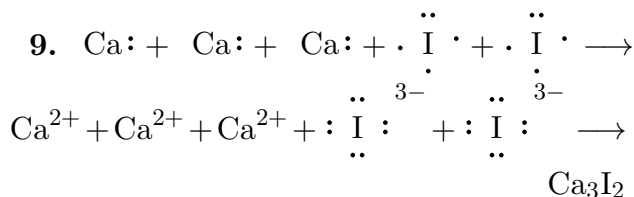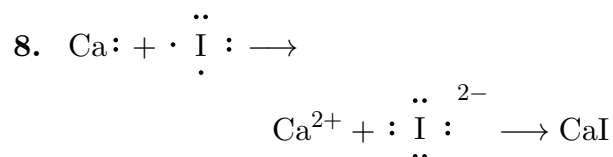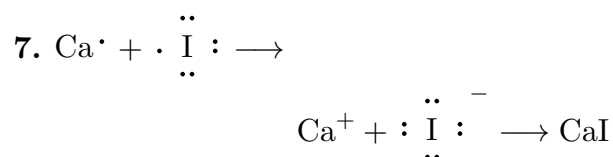
Important Message: Please please please remember to CORRECTLY bubble in your name, uteid, and version number for your exam. We are averaging about 20 students who fail to do this. Think of it as the easiest question on the exam. If you get it wrong, you'll get a zero when we grade it. Because Quest is not that user friendly, it will take at least 24 hours to get your avoidable mistake fixed. It is a huge pain in the ass - know that. Please take the time to bubble in your information carefully. Lets have NO mistakes on this information please. Thank you so much. I'm only putting this plea on this practice exam and not on the real exam. So just remember to do it.
  - Dr. McCord

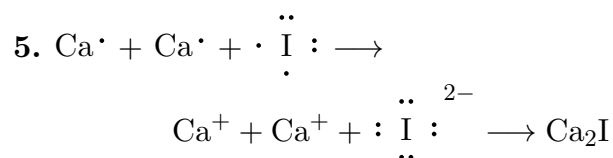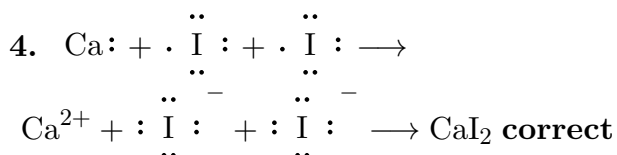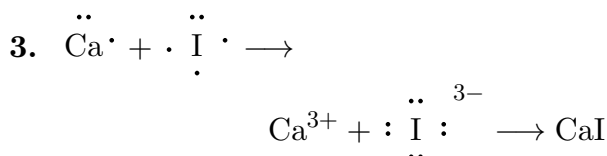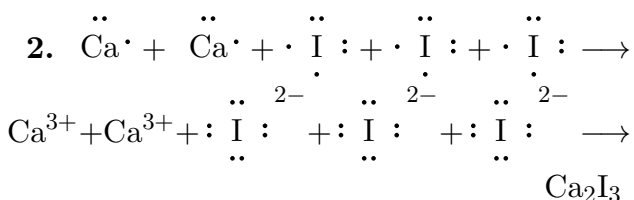**Holt da 6 3 rev 2b**
001   10.0 points
Use electron-dot notation to demonstrate the formation of ionic compounds involving the elements Ca and I.

1. None of these

2. $\overset{..}{Ca}\cdot + \overset{..}{Ca}\cdot + \cdot \overset{..}{I}: + \cdot \overset{..}{I}: + \cdot \overset{..}{I}: \longrightarrow$
$Ca^{3+} + Ca^{3+} + : \overset{..}{I} :^{2-} + : \overset{..}{I} :^{2-} + : \overset{..}{I} :^{2-} \longrightarrow$
Ca_2I_3

3. $\overset{..}{Ca}\cdot + \cdot \overset{..}{I} \cdot \longrightarrow$
$Ca^{3+} + : \overset{..}{I} :^{3-} \longrightarrow CaI$

4. $Ca: + \cdot \overset{..}{I}: + \cdot \overset{..}{I}: \longrightarrow$
$Ca^{2+} + : \overset{..}{I} :^- + : \overset{..}{I} :^- \longrightarrow CaI_2$ **correct**

5. $Ca\cdot + Ca\cdot + \cdot \overset{..}{I}: \longrightarrow$
$Ca^+ + Ca^+ + : \overset{..}{I} :^{2-} \longrightarrow Ca_2I$

6. $\overset{..}{Ca}\cdot + \cdot \overset{..}{I}: + \cdot \overset{..}{I}: + \cdot \overset{..}{I}: \longrightarrow$
$Ca^{3+} + : \overset{..}{I} :^- + : \overset{..}{I} :^- + : \overset{..}{I} :^- \longrightarrow Ca_2I_3$

7. $Ca\cdot + \cdot \overset{..}{I}: \longrightarrow$
$Ca^+ + : \overset{..}{I} :^- \longrightarrow CaI$

8. $Ca: + \cdot \overset{..}{I}: \longrightarrow$
$Ca^{2+} + : \overset{..}{I} :^{2-} \longrightarrow CaI$

9. $Ca: + Ca: + Ca: + \cdot \overset{..}{I} \cdot + \cdot \overset{..}{I} \cdot \longrightarrow$
$Ca^{2+} + Ca^{2+} + Ca^{2+} + : \overset{..}{I} :^{3-} + : \overset{..}{I} :^{3-} \longrightarrow Ca_3I_2$

10. $Ca\cdot + Ca\cdot + Ca\cdot + \cdot \overset{..}{I} \cdot \longrightarrow$
$Ca^+ + Ca^+ + Ca^+ + : \overset{..}{I} :^{3-} \longrightarrow Ca_3I$
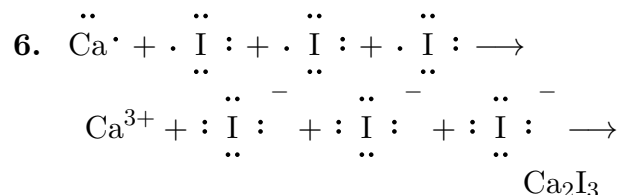
Explanation:
$Ca:$ gives up two electrons to form the Ca^{2+} cation, and $\cdot \overset{..}{I}:$ acquires an electron to form the $: \overset{..}{I} :^-$ anion.

Two iodine anions combine with one calcium cation to form an electrically neutral compound:
$Ca: + \cdot \overset{..}{I}: + \cdot \overset{..}{I}: \longrightarrow$
$Ca^{2+} + : \overset{..}{I} :^- + : \overset{..}{I} :^- \longrightarrow CaI_2$
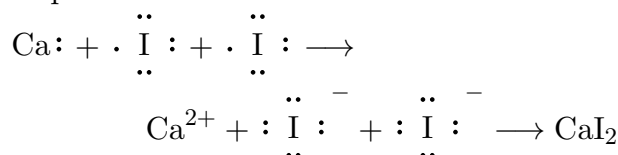
Mlib 01 2023
002   10.0 points
Choose the pair of names and formulas that

do NOT match.

1. As_4O_6 : tetraarsenic oxide **correct**

2. SO_3 : sulfur trioxide

3. NO : nitrogen monoxide

4. Cl_2O_7 : dichlorine heptoxide

5. N_2O_5 : dinitrogen pentoxide

Explanation:
Cl_2O_7, N_2O_5, SO_3, and NO are all covalent molecules and are named correctly using the appropriate prefixes.

As_4O_6 is also a covalent compound; appropriate prefixes are used in the name. It should be correctly named tetraarsenic hexoxide to indicate the presence of six oxygen atoms in each molecule.

LDE Rank Lattice Energy 004
003   10.0 points
Rank the crystal lattice energy of the salts Al_2O_3, $CaCl_2$, CaO, NaF, $Mg_3(PO_4)_2$ from least to greatest:

1. NaF $<$ CaO $<$ $CaCl_2$ $<$ $Mg_3(PO_4)_2$ $<$ Al_2O_3

2. NaF $<$ $CaCl_2$ $<$ CaO $<$ $Mg_3(PO_4)_2$ $<$ Al_2O_3 **correct**

3. CaO $<$ NaF $<$ $CaCl_2$ $<$ Al_2O_3 $<$ $Mg_3(PO_4)_2$

4. $Mg_3(PO_4)_2$ $<$ NaF $<$ $<$ CaO $CaCl_2$ $<$ Al_2O_3

5. $Mg_3(PO_4)_2$ $<$ NaF $<$ $CaCl_2$ $<$ CaO $<$ Al_2O_3

Explanation:
Sodium fluoride has the least lattice energy because of its small charges, then calcium chloride and calcium oxide with incrementally increasing charges. Both magnesium phosphate and aluminum oxide have identical magnitudes for their charges, but both of the latter are smaller than than their counterparts in the former and consequently have a greater lattice energy.

Campion 05 Ex 01 07
004   10.0 points
If the interaction energy between a sodium ion and a chloride ion in table salt is 760 kJ/mol, what is the interaction energy between a zinc ion (Zn^{2+}) and a sulfide ion (S^{2-}) in a hypothetical structure in which the inter-ionic distances are the same as that of NaCl?

1. 1.140 kJ/mol

2. 3040 kJ/mol **correct**

3. 1520 kJ/mol

4. 1140 kJ/mol

5. 760 kJ/mol

Explanation:
$q_{Zn} = +2$ C $q_S = -2$ C
$q_{Na} = +1$ C $q_{Cl} = -1$ C
$V_{NaCl} = 760$ kJ/mol

If r is the distance between the two ions (the sum of the ionic radii), the energy interaction between ions is given by Coulomb's Law:

$$V = \frac{q_1\,q_2}{4\,\pi\,\epsilon_0\,r}$$
$$r = \frac{q_1\,q_2}{4\,\pi\,\epsilon_0\,V}$$

$$r_{NaCl} = r_{ZnS}$$
$$\frac{q_{Na}\,q_{Cl}}{4\,\pi\,\epsilon_0\,V_{NaCl}} = \frac{q_{Zn}\,q_S}{4\,\pi\,\epsilon_0\,V_{ZnS}}$$
$$V_{ZnS} = \frac{q_{Zn}\,q_S\,V_{NaCl}}{q_{Na}\,q_{Cl}}$$
$$= \frac{(+2\text{ C})(-2\text{ C})(-760\text{ kJ/mol})}{(+1\text{ C})(-1\text{ C})}$$
$$= -3040\text{ kJ/mol},$$

the energy released by the interaction.

Internuclear Distance 1
005   10.0 points

Consider a potential energy diagram for the interaction of a sodium ion (Na^+) with a chloride ion (Cl^-). Which of the following statements is/are true?

I) Repulsive forces predominate at very small internuclear distances.
II) The minimum potential energy occurs when attractive forces are greatest.
III) Attractive forces are linearly dependent on the internuclear distance.

1. I and II

2. III only

3. I and III

4. I only **correct**

5. II and III

6. I, II and III

7. II only

Explanation:
It is true that repulsive forces predominate at very small internuclear distances; this is the result of the positively charged nuclei electrostatically repelling each other, which occurs despite that fact that sodium ion and chloride ion have opposite charges overall. The attractive forces actually continue to increase as you force the ions closer together, but the repulsive forces increase more, thus offsetting the attractive forces and producing a net increase in potential energy. Thus, the minimum potential energy does not occur at maximum attractive force, but rather at a "sweet spot" where the sum of attractive and repulsive forces is minimized. Attractive forces are a rectangular hyperbola that asymptotically approach zero as the internuclear distance approaches infinity.

Line Drawing to Formula
006   10.0 points

Determine the molecular formula for the molecule:
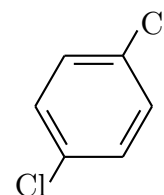
1. $C_6H_6Cl_2$

2. $C_4H_{12}Cl_2$

3. C_6H_6

4. $C_6H_4Cl_2$ **correct**

5. C_6H_4Cl

6. $C_4H_4Cl_2$

Explanation:

LDE Molecular Polarity 004
007   10.0 points

Consider the labeled bonds in the molecule below and rank them from least to most polar in terms of difference in electronegativity.
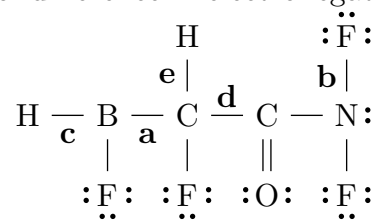
1. $a < b < c < d < e$

2. $d < c < e < a < b$ **correct**

3. $e < b < d < c < a$

4. $c < d < e < b < a$

5. $d < c < a < e < b$

Explanation:
The polarity of a bond is proportional to the difference in electronegativity of the two bonded atoms. The electronegatavity differences of the labeled bonds are 0.5, 1.0, 0.1, 0.0 and 0.4, respectively.

Lewis CO dash

008   10.0 points

Which of the following is the correct Lewis formula for carbon monoxide (CO)?

1. $: O \equiv C :$ **correct**

2. $: \overset{..}{\underset{..}{O}} - C :$

3. $\overset{..}{\underset{..}{O}} - \overset{..}{\underset{..}{C}}$

4. $\overset{.}{\underset{..}{O}} - \overset{..}{C} :$

5. $: \overset{..}{\underset{..}{O}} - \overset{..}{\underset{..}{C}} :$

6. $O - C$

7. $: O - \overset{..}{\underset{..}{C}} :$

8. $O - \overset{..}{\underset{..}{C}} :$

9. $\overset{..}{\underset{..}{O}} = \overset{..}{\underset{..}{C}}$

10. $: \overset{..}{\underset{..}{O}} - \overset{.}{\underset{.}{C}}$

Explanation:

The Lewis formula for carbon monoxide (CO) is $: O \equiv C :$

Mlib 03 1055
009   10.0 points

What kind of carbon-carbon bond is in a molecule of ethylene (C_2H_4)?

1. ionic

2. single

3. double **correct**

4. triple

Explanation:

To draw the dot structure for C_2H_4 we must first calculate the number of valence electrons available from the atoms:

$$A = 4 \times 2 \,(\text{C atom}) + 1 \times 4 \,(\text{H atoms})$$
$$= 12 \; e^-$$

Hydrogen can form only one bond, so carbon atoms must serve as the central atoms. We place the most symmetrical arrangement possible.

The correct dot structure for the molecule should show a complete octet (8 electrons) around the carbon atoms, two electrons around each hydrogen atom, and a total of 12 valence electrons for the entire structure:
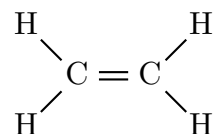
As can be seen above, the carbon-carbon bond is a double bond.

(If using the $S = N - A$ rule to determine the dot structure,

$N = (8 \times 2) + (2 \times 4) = 24 \, e^-$ and
$S = 24 - 12 = 12 \, e^-$.

This would indicate 6 bonds and correspond to the structure shown above.)

Lewis BF3 dash
010   10.0 points

Which of the following is the correct Lewis formula for boron trifluoride (BF_3)?

1. $: \overset{..}{\underset{..}{F}} - B \overset{\overset{..}{F} :}{\underset{\overset{..}{\underset{..}{F}}}{}}$

2. $: \overset{..}{\underset{..}{F}} - B - \overset{..}{\underset{..}{F}} : \quad | \quad : \overset{}{\underset{..}{F}} :$

3. $: \overset{..}{\underset{..}{F}} - B - \overset{..}{\underset{..}{F}} : \quad | \quad : \overset{}{\underset{..}{F}} :$ **correct**

4.

$$\text{F}\!=\!\text{B}$$
with F and F branches

5.

$$\text{F}-\text{B}-\text{F}$$
$$|$$
$$:\text{F}:$$

6.

$$\text{F}-\text{B}-\text{F}:$$
$$|$$
$$:\text{F}:$$

7.

$$\text{F}-\text{B}-\text{F}:$$
$$|$$
$$:\text{F}:$$

8.

$$\text{F}-\text{B}$$
with F: (double bond) and F branches

9. $$\text{F}-\text{F}\!\equiv\!\text{B}-\text{F}$$

10.

$$\text{F}-\text{B}-\text{F}$$
$$|$$
$$:\text{F}:$$

Explanation:
The Lewis formula for boron trifluoride

$$(\text{BF}_3) \text{ is } \quad :\text{F}-\text{B}-\text{F}:$$
$$|$$
$$:\text{F}:$$
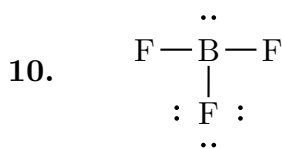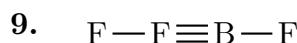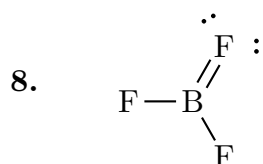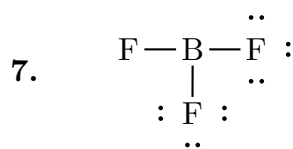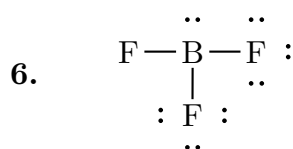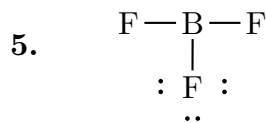
ChemPrin3e T02 23
011   10.0 points

Draw the Lewis structure of xenon difluoride and give the number of lone pairs of electrons around the central atom.

1. 4

2. 3 **correct**

3. 1

4. 2

5. 5

Explanation:
The Lewis structure is $:\text{F}-\text{Xe}-\text{F}:$

LDE Resonant Species 002
012   10.0 points

Which of the following species exhibit resonance/delocalization?
I) HCN
II) O_3
III) CO_3^{2-}

1. I, III

2. III only

3. I only

4. II, III **correct**

5. I, II, III

6. II only

7. I, II

Explanation:
Both ozone and carbonate have a single pair of resonant electrons and are famous examples of resonant molecules. Cyanide cannot have resonance since hydrogen can only form a single bond.

ChemPrin3e 02 46
013   10.0 points

Consider the following sets A, B, and C

$$\text{A}_1) \left[\ddot{\text{N}}\!=\!\text{C}\!=\!\ddot{\text{N}} \right]^{2-}$$

$$\text{A}_2) \left[:\text{N}\!\equiv\!\text{C}-\ddot{\text{N}}: \right]^{2-}$$

B₁) $\left[\begin{array}{c} \ddot{\text{O}} \\ | \\ \ddot{\text{O}}-\text{As}-\ddot{\text{O}} \\ | \\ \ddot{\text{O}} \end{array}\right]^{3-}$ C₁) $\left[\begin{array}{c} \ddot{\text{O}} \\ \| \\ \ddot{\text{O}}=\text{I}=\ddot{\text{O}} \\ | \\ \ddot{\text{O}} \end{array}\right]^{-}$

B₂) $\left[\begin{array}{c} \ddot{\text{O}} \\ \| \\ \ddot{\text{O}}-\text{As}-\ddot{\text{O}} \\ | \\ \ddot{\text{O}} \end{array}\right]^{3-}$ C₂) $\left[\begin{array}{c} \ddot{\text{O}} \\ | \\ \ddot{\text{O}}=\text{I}=\ddot{\text{O}} \\ | \\ \ddot{\text{O}} \end{array}\right]^{-}$

B₃) $\left[\begin{array}{c} \ddot{\text{O}} \\ \| \\ \ddot{\text{O}}=\text{As}=\ddot{\text{O}} \\ \| \\ \ddot{\text{O}} \end{array}\right]^{3-}$ C₃) $\left[\begin{array}{c} \ddot{\text{O}} \\ | \\ \ddot{\text{O}}-\text{I}-\ddot{\text{O}} \\ | \\ \ddot{\text{O}} \end{array}\right]^{-}$
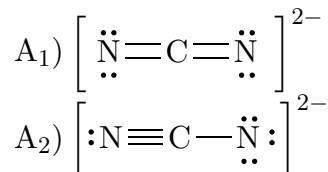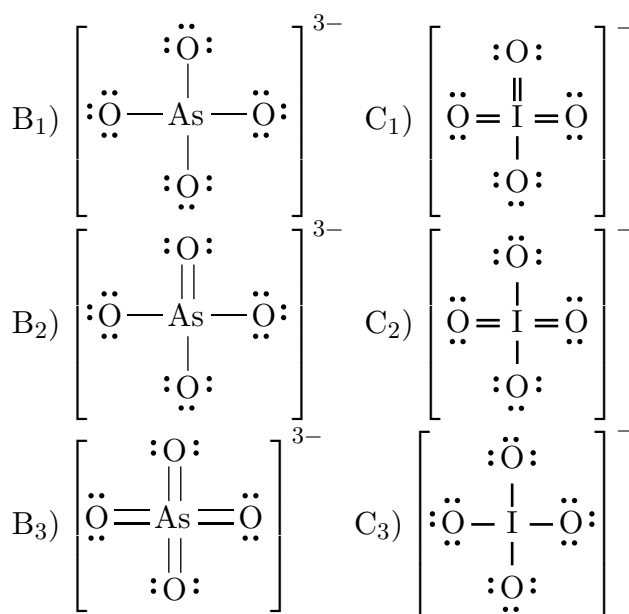
of ions.

Determine the formal charge on each atom and identify the structure of lowest energy for each set A, B, and C.

1. A₂, B₂, C₂

2. A₂, B₃, C₃

3. A₁, B₁, C₁

4. A₁, B₂, C₃

5. A₂, B₃, C₂

6. A₁, B₂, C₁ **correct**

Explanation:

A₁) $\left[\ddot{\text{N}}=\text{C}=\ddot{\text{N}}\right]^{2-}$
     $\;\;\;{-1}\quad 0 \quad {-1}$

A₂) $\left[\text{:N}\equiv\text{C}-\ddot{\text{N}}\text{:}\right]^{2-}$
     $\;\;\;\;\; 0 \quad\; 0 \quad\; {-2}$

B₁) $\left[\begin{array}{c} \overset{-1}{\ddot{\text{O}}} \\ \overset{+1}{|} \\ \ddot{\text{O}}-\text{As}-\ddot{\text{O}} \\ \overset{|}{\underset{-1}{}} \\ {-1}\ddot{\text{O}} \end{array}\right]^{3-}$ C₁) $\left[\begin{array}{c} \overset{0}{\text{O}} \\ \| \\ \ddot{\text{O}}=\text{I}=\ddot{\text{O}} \\ | \\ {-1}\ddot{\text{O}} \end{array}\right]^{-}$
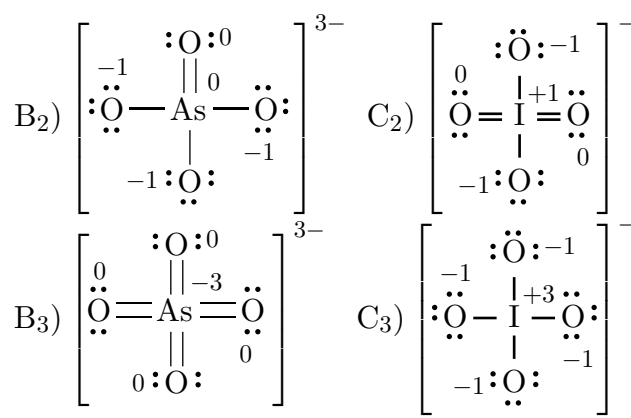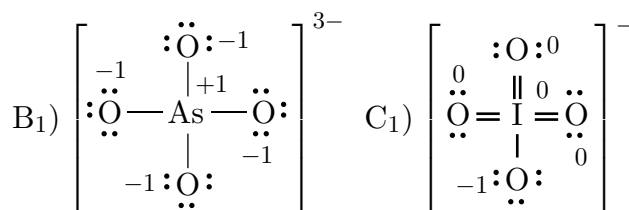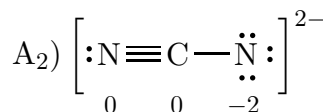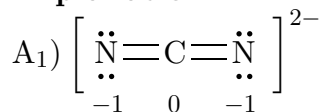
B₂) $\left[\begin{array}{c} \overset{0}{\text{O}} \\ \overset{0}{\|} \\ \overset{-1}{\ddot{\text{O}}}-\text{As}-\overset{}{\ddot{\text{O}}} \\ \overset{|}{\underset{-1}{}} \\ {-1}\ddot{\text{O}} \end{array}\right]^{3-}$ C₂) $\left[\begin{array}{c} \overset{-1}{\ddot{\text{O}}} \\ \overset{0}{} \overset{+1}{|} \\ \ddot{\text{O}}=\text{I}=\ddot{\text{O}} \\ | \quad 0 \\ {-1}\ddot{\text{O}} \end{array}\right]^{-}$

B₃) $\left[\begin{array}{c} \overset{0}{\text{O}} \\ \overset{0}{\|}\overset{-3}{} \\ \overset{0}{\text{O}}=\text{As}=\text{O} \\ \overset{\|}{\underset{0}{}} \\ 0\,\text{O} \end{array}\right]^{3-}$ C₃) $\left[\begin{array}{c} \overset{-1}{\ddot{\text{O}}} \\ \overset{-1}{} \overset{+3}{|} \\ \ddot{\text{O}}-\text{I}-\ddot{\text{O}} \\ | \quad {-1} \\ {-1}\ddot{\text{O}} \end{array}\right]^{-}$

ChemPrin3e T03 26
014   10.0 points

Which of the following has bond angles slightly less than 109.5°?

1. I_3^-

2. O_3

3. CH_3^+

4. HOCl **correct**

5. NO_2^-

Explanation:

Only HOCl has four regions of electron density around the central atom; the two lone pairs repel the bonds more, making the bond angle less than 109.5°.

Mlib 03 2087
015   10.0 points

Which of the following is planar?

1. NH_3

2. SO_3^{2-}

3. NO_3^- **correct**

4. H_3O^+

5. PF_3

Explanation:

All except the nitrate ion have trigonal pyramidal molecular geometries; the nitrate ion has trigonal planer geometry.

ChemPrin3e T03 16
016   10.0 points

What is the shape (molecular geometry) of IF_4^+?

1. square planar

2. tetrahedral

3. seesaw **correct**

4. trigonal bipyramidal

5. T-shaped

Explanation:

There are five regions of electron density (including one lone pair) around the central atom:

$$\left[:\ddot{\text{F}} - \text{I} - \ddot{\text{F}}: \atop \diagup \diagdown \atop \ddot{\text{F}}: \quad :\ddot{\text{F}}: \right]^+$$
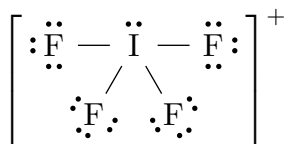
ChemPrin3e T03 35
017   10.0 points

Which of the following is polar?

1. IF_5 **correct**

2. SF_6

3. ICl_4^-

4. XeF_4

5. PCl_5

Explanation:

XeF_4 and ICl_4^- have 6 RHED with two RHED being lone pairs situated opposite each other so their effects cancel. Only IF_5 has an unbalanced number of lone pairs which places one of the polar I—F bonds opposite a lone pair.

This unopposed dipole and the lone pair itself make it polar. The others have either 5 or 6 RHED and no lone pairs on the central atom.

LDE VB Hybridization 001
018   10.0 points

Which of the following statements concerning hybrid orbitals is/are true?

I) Hybrid orbitals are energetically degenerate.

II) Any element can form sp^3d^2 hybrid orbitals.

III) Hybridizing a $2s$ and a $2p$ orbital would produce one single sp hybrid orbital.

1. III only

2. II, III

3. I only **correct**

4. II only

5. I, II

6. I, III

7. I, II, III

Explanation:

Statement I is true; hybridization was developed as a theoretical framework to explain the energetic degeneracy of bonds in molecules. Statement II is false; hybridization involving d orbitals requires access to empty d orbitals, and thus begins in period 3. Statement III is false; the number of orbitals used to hybridize is always equal to the number of hybridized orbitals, so using a $2s$ and a $2p$ orbital would result in two sp hybrid orbitals.

LDE VB Sigma Pi Bonds 005
019   10.0 points

How many different types of sigma (σ) bonds are found in ethanoic acid, CH_3COOH? In other words, how many different combinations of atomic orbitals are used when forming the σ bonds in ethanoic acid.

1. 5

2. 4

3. 7

4. 3 **correct**

5. none of the above

Explanation:
All of the σ bonds found in ethanoic acid are $sp^3 - 1s$, $sp^3 - sp^2$, or $sp^2 - sp^2$.
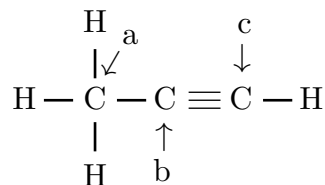

LDE VB Sigma Pi Bonds 004
020   10.0 points
How many sigma (σ) and pi (π) bonds are in the Lewis structure for $C(COOH)_4$?

1. $8\ \sigma, 4\ \pi$

2. $12\ \sigma, 4\ \pi$

3. $16\ \sigma, 0\ \pi$

4. $12\ \sigma, 0\ \pi$

5. $16\ \sigma, 4\ \pi$ **correct**

Explanation:

ChemPrin3e 03 76
021   10.0 points

Carbon has a valence of four in nearly all of its compounds and can form chains and rings of C atoms. Consider the propyne structure.
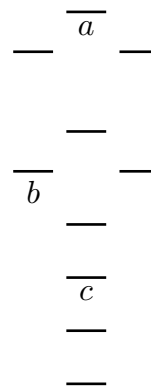
$$
\begin{array}{ccccc}
\text{H} & a & & c & \\
| \swarrow & & & \downarrow & \\
\text{H} - \text{C} - \text{C} & \equiv & \text{C} - \text{H} \\
| & \uparrow & & \\
\text{H} & b & &
\end{array}
$$

What hybridizations would you expect for the carbon atoms identified by a, b, and c, respectively?

1. $sp^3,\ sp^2,\ sp^3$

2. $sp^2,\ sp^2,\ sp^3$

3. $sp^2,\ sp,\ sp$

4. $sp^3,\ sp,\ sp^3$

5. $sp^3,\ sp,\ sp$ **correct**

6. $sp^3,\ sp^3,\ sp^2$

7. $sp^2,\ sp,\ sp^2$

Explanation:

$$
\begin{array}{ccccc}
\text{H}\ sp^3 & & sp & \\
| \swarrow & & \downarrow & \\
\text{H} - \text{C} - \text{C} & \equiv & \text{C} - \text{H} \\
| & \uparrow & & \\
\text{H} & sp &
\end{array}
$$
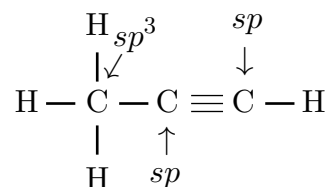
LDE MO Diagram 001
022   10.0 points
Consider the following molecular orbital diagram:

$$
\begin{array}{c}
\underline{}\ \ \overline{a}\ \ \underline{} \\[6pt]
\underline{} \\[4pt]
\underline{b}\quad \underline{} \\[4pt]
\underline{} \\[6pt]
\overline{c} \\[4pt]
\underline{}
\end{array}
$$

What are the names of the labeled orbitals, a, b, and c, respectively?

1. $\pi^*_{2p}, \sigma^*_{2p}, \sigma_{2s}$

2. $\pi^*_{2p}, \sigma_{2p}, \sigma^*_{2s}$

3. $\sigma^*_{2p}, \pi_{2p}, \sigma^*_{1s}$

4. $\sigma^*_{2p}, \pi_{2p}, \sigma_{2s}$ **correct**

5. $\sigma^*_{2p}, \pi_{2p}, \sigma^*_{2s}$

Explanation:

LDE Bond Order 009
023   10.0 points
All of the species below have the same bond order except for one of them. Which is it?

1. Ne_2^+

2. B_2^- **correct**

3. H_2^-

4. F_2^-

5. H_2^+

Explanation:
All of the species have a bond order of 0.5 except for B_2^-, which has a bond order of 1.5.

Msci 09 0116
024   10.0 points
An antibonding orbital is formed when

1. a p_x-orbital overlaps a p_z-orbital.

2. a free electron is present in the molecule.

3. an s-orbital overlaps a p-orbital.

4. None of these is correct.

5. the overlap of the corresponding atomic orbitals leads to destructive interference. **correct**

Explanation:
In-phase overlap of two atomic orbitals results in molecular bonding orbitals; out-of-phase overlap of two atomic orbitals results in molecular antibonding orbitals.

LDE Paramagnetism 004
025   10.0 points
Which of the following species is/are paramagnetic?
  I) Li_2^-
 II) O_2
III) H_2^+

1. II and III

2. I and III

3. I, II and III **correct**

4. II only

5. III only

6. I only

7. I and II

Explanation:
Li_2^- and H_2^+ both have an odd number of electrons and therefore must be paramagnetic. O_2 has 16 total electrons, the last two of which must go into separate degenerate π^* anti-bonding orbitals.