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. Be-easiest question on the exam. If you get it

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

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**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. Ca\(^+\) + Ca\(^+\) + I\(^-\) : + + I\(^-\) : + + I\(^-\) : → Ca\(^2+\) + Ca\(^2+\) + I\(^-\) : → Ca\(_2\)I
   Ca\(^3+\)+Ca\(^3+\)+: I\(^-\) : + : I\(^-\) : + : I\(^-\) : → Ca\(_2\)I

3. Ca\(^+\) + I\(^-\) : → Ca\(^2+\) + I\(^-\) : → CaI

4. Ca\(^+\) : + I\(^-\) : + I\(^-\) : → Ca\(^2+\) + I\(^-\) : + I\(^-\) : → CaI\(_2\) correct

5. Ca\(^+\) + Ca\(^+\) + I\(^-\) : → Ca\(^+\) + Ca\(^+\) + I\(^-\) : → CaI

6. Ca\(^+\) + + I\(^-\) : + + I\(^-\) : → Ca\(^3+\) + : I\(^-\) : + : I\(^-\) : → Ca\(_2\)I

7. Ca\(^+\) + I\(^-\) : → Ca\(^+\) + : I\(^-\) : → CaI

8. Ca\(^2+\) + I\(^-\) : → Ca\(^2+\) + : I\(^-\) : → CaI

9. Ca\(^+\) + Ca\(^+\) + Ca\(^+\) + I\(^-\) : + + I\(^-\) : → Ca\(^2+\) + Ca\(^2+\) + Ca\(^2+\) + : I\(^-\) : + : I\(^-\) : → Ca\(_3\)I

10. Ca\(^+\) + Ca\(^+\) + Ca\(^+\) + I\(^-\) : → Ca\(^2+\) + Ca\(^+\) + Ca\(^+\) + : I\(^-\) : → Ca\(_3\)I

**Explanation:**

Ca\(^+\) gives up two electrons to form the Ca\(^2+\) cation, and I\(^-\) : acquires an electron to form the I\(^-\) anion.

Two iodine anions combine with one calcium cation to form an electrically neutral compound:

Ca\(^+\) + I\(^-\) : + I\(^-\) : → Ca\(^2+\) + I\(^-\) : + : I\(^-\) : → CaI\(_2\)

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**002 10.0 points**

Choose the pair of names and formulas that
do NOT match.

1. \( \text{As}_4\text{O}_6 : \) tetraarsenic oxide \textbf{correct}
2. \( \text{SO}_3 : \) sulfur trioxide
3. \( \text{NO} : \) nitrogen monoxide
4. \( \text{Cl}_2\text{O}_7 : \) dichlorine heptoxide
5. \( \text{N}_2\text{O}_5 : \) dinitrogen pentoxide

\textbf{Explanation:}

\( \text{Cl}_2\text{O}_7, \text{N}_2\text{O}_5, \text{SO}_3, \text{and NO} \) are all covalent molecules and are named correctly using the appropriate prefixes.

\( \text{As}_4\text{O}_6 \) 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.

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\textbf{LDE Rank Lattice Energy 004}

\textbf{003 10.0 points}

Rank the crystal lattice energy of the salts \( \text{Al}_2\text{O}_3, \text{CaCl}_2, \text{CaO}, \text{NaF}, \text{Mg}_3(\text{PO}_4)_2 \) from least to greatest:

1. \( \text{NaF} < \text{CaO} < \text{CaCl}_2 < \text{Mg}_3(\text{PO}_4)_2 < \text{Al}_2\text{O}_3 \)

2. \( \text{NaF} < \text{CaCl}_2 < \text{CaO} < \text{Mg}_3(\text{PO}_4)_2 < \text{Al}_2\text{O}_3 \) \textbf{correct}

3. \( \text{CaO} < \text{NaF} < \text{CaCl}_2 < \text{Al}_2\text{O}_3 < \text{Mg}_3(\text{PO}_4)_2 \)

4. \( \text{Mg}_3(\text{PO}_4)_2 < \text{NaF} < \text{CaO} \text{ CaCl}_2 < \text{Al}_2\text{O}_3 \)

5. \( \text{Mg}_3(\text{PO}_4)_2 < \text{NaF} < \text{CaCl}_2 < \text{CaO} < \text{Al}_2\text{O}_3 \)

\textbf{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.

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\textbf{Campion 05 Ex 01 07}

\textbf{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 \textbf{correct}
3. 1520 kJ/mol
4. 1140 kJ/mol
5. 760 kJ/mol

\textbf{Explanation:}

\[ q_{\text{Zn}} = +2 \text{ C} \quad q_{\text{S}} = -2 \text{ C} \]
\[ q_{\text{Na}} = +1 \text{ C} \quad q_{\text{Cl}} = -1 \text{ C} \]

\[ V_{\text{NaCl}} = 760 \text{ kJ/mol} \]

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:

\[ V = \frac{q_1 q_2}{4 \pi \epsilon_0 r} \]
\[ r = \frac{q_1 q_2}{4 \pi \epsilon_0 V} \]

\[ \frac{r_{\text{NaCl}}}{4 \pi \epsilon_0 V_{\text{NaCl}}} = \frac{q_{\text{Na}} q_{\text{Cl}}}{q_{\text{Zn}} q_{\text{S}}} \]
\[ V_{\text{ZnS}} = \frac{q_{\text{Zn}} q_{\text{S}}}{q_{\text{Na}} q_{\text{Cl}}} \]

\[ = \frac{(+2 \text{ C})(-2 \text{ C})(-760 \text{ kJ/mol})}{(+1 \text{ C})(-1 \text{ C})} \]

\[ = -3040 \text{ kJ/mol} \]

the energy released by the interaction.

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\textbf{Internuclear Distance 1}

\textbf{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 the 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.

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**Line Drawing to Formula 006 10.0 points**

Determine the molecular formula for the molecule:

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**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.

```
H | e |
---|---|
H c B a C d C N: |
```

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 electronegativity differences of the labeled bonds are 0.5, 1.0, 0.1, 0.0 and 0.4, respectively.

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

1. \( \cdot \cdot \cdot \cdot \) \( \cdot \cdot \cdot \cdot \) \( \text{correct} \)
2. \( \cdot \cdot \cdot \cdot \) \( \cdot \cdot \cdot \cdot \)
3. \( \cdot \cdot \cdot \cdot \) \( \cdot \cdot \cdot \cdot \)
4. \( \cdot \cdot \cdot \cdot \) \( \cdot \cdot \cdot \cdot \)
5. \( \cdot \cdot \cdot \cdot \) \( \cdot \cdot \cdot \cdot \)
6. \( \cdot \cdot \cdot \cdot \) \( \cdot \cdot \cdot \cdot \)
7. \( \cdot \cdot \cdot \cdot \) \( \cdot \cdot \cdot \cdot \)
8. \( \cdot \cdot \cdot \cdot \) \( \cdot \cdot \cdot \cdot \)
9. \( \cdot \cdot \cdot \cdot \) \( \cdot \cdot \cdot \cdot \)
10. \( \cdot \cdot \cdot \cdot \) \( \cdot \cdot \cdot \cdot \)

Explanation:
The Lewis formula for carbon monoxide (CO) is \( \cdot \cdot \cdot \cdot \) \( \cdot \cdot \cdot \cdot \) : correct

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:

```
   H
  /   \H
C === C
  \   /
   H
```

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,
\[
N = (8 \times 2) + (2 \times 4) = 24 \, e^- \text{ and } S = 24 - 12 = 12 \, e^-.
\]
This would indicate 6 bonds and correspond to the structure shown above.)

Lewis BF3 dash

Which of the following is the correct Lewis formula for boron trifluoride (BF\(_3\)):

1. \( \cdot \cdot \cdot \cdot \) \( \cdot \cdot \cdot \cdot \) \( \cdot \cdot \cdot \cdot \) \( \cdot \cdot \cdot \cdot \)
2. \( \cdot \cdot \cdot \cdot \) \( \cdot \cdot \cdot \cdot \) \( \cdot \cdot \cdot \cdot \) \( \cdot \cdot \cdot \cdot \)
3. \( \cdot \cdot \cdot \cdot \) \( \cdot \cdot \cdot \cdot \) \( \cdot \cdot \cdot \cdot \) \( \text{correct} \)
4. \( \cdot \cdot \cdot \cdot \) \( \cdot \cdot \cdot \cdot \) \( \cdot \cdot \cdot \cdot \) \( \cdot \cdot \cdot \cdot \)

Explanation:
To draw the dot structure for \( \text{C}_2\text{H}_4 \) we must first calculate the number of valence electrons available from the atoms:
4. \[ \begin{array}{c}
\text{F} \\
\text{B} \\
\text{F}
\end{array} \]

5. \[ \begin{array}{c}
\text{F} \\
\text{B} \\
\text{F}
\end{array} \]

6. \[ \begin{array}{c}
\text{F} \\
\text{B} \\
\text{F}
\end{array} \]

7. \[ \begin{array}{c}
\text{F} \\
\text{B} \\
\text{F}
\end{array} \]

8. \[ \begin{array}{c}
\text{F} \\
\text{B} \\
\text{F}
\end{array} \]

9. \[ \begin{array}{c}
\text{F} \\
\text{B} \\
\text{F}
\end{array} \]

10. \[ \begin{array}{c}
\text{F} \\
\text{B} \\
\text{F}
\end{array} \]

Explanation:
The Lewis formula for boron trifluoride (BF₃) is:
\[ \text{F} - \text{B} - \text{F} \]

LDE Resonant Species 002
012 10.0 points
Which of the following species exhibit resonance/delocalization?
I) HCN
II) O₃
III) CO₃²⁻

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
A₁) \[ \begin{array}{c}
\text{N} \\
\text{C} \\
\text{N}
\end{array} \]
A₂) \[ \begin{array}{c}
\text{N} \\
\text{C} \\
\text{N}
\end{array} \]

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
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) \[ \text{\n} \begin{array}{c}
\text{N} \equiv \text{C} \equiv \text{N} \\
-1 & 0 & -1
\end{array} \]

A₂) \[ \text{\n} \begin{array}{c}
\text{N} \equiv \text{C} \equiv \text{N} \\
0 & 0 & -2
\end{array} \]

B₁) \[ \text{\n} \begin{array}{c}
\text{O} \equiv \text{As} \equiv \text{O} \\
-1 & +1 & -1
\end{array} \]

C₁) \[ \text{\n} \begin{array}{c}
\text{O} \equiv \text{I} \equiv \text{O} \\
0 & 0 & 0
\end{array} \]

B₂) \[ \text{\n} \begin{array}{c}
\text{O} \equiv \text{As} \equiv \text{O} \\
-1 & -1 & 0
\end{array} \]

C₂) \[ \text{\n} \begin{array}{c}
\text{O} \equiv \text{I} \equiv \text{O} \\
-1 & 0 & 0
\end{array} \]

B₃) \[ \text{\n} \begin{array}{c}
\text{O} \equiv \text{As} \equiv \text{O} \\
0 & 0 & 0
\end{array} \]

C₃) \[ \text{\n} \begin{array}{c}
\text{O} \equiv \text{I} \equiv \text{O} \\
-1 & 0 & 1
\end{array} \]

ChemPrin3e T03 26

014 10.0 points

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

1. I⁻
2. O₃
3. CH₃⁺
4. HOCl correct
5. NO₂⁻

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. NH₃
2. SO₃²⁻
3. NO₃⁻ correct
4. H₃O⁺
5. PF₃

Explanation:

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

$$\begin{array}{c}
\text{\textbullet F} \\
\text{\textbullet I} \\
\text{\textbullet F} \\
\text{\textbullet I} \\
\text{\textbullet F}
\end{array}$$

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.

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.

How many different types of sigma ($\sigma$) bonds are found in ethanoic acid, CH$_3$COOH? In other words, how many different combinations of atomic orbitals are used when forming the $\sigma$ bonds in ethanoic acid.

1. 5
2. 4
3. 7
4. 3 correct
5. none of the above

Explanation:
All of the $\sigma$ 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 ($\sigma$) and pi ($\pi$) 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:
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.

LDE MO Diagram 001
022 10.0 points
Consider the following molecular orbital diagram:

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:
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.

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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.

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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**