Midterm 1 Worksheet Solutions

The problems in this review are designed to help prepare you for your upcoming exam. Questions pertain to material covered in the course and are intended to reflect the topics likely to appear in the exam. Keep in mind that this worksheet was created by CARE tutors, and while it is thorough, it is not comprehensive. In addition to exam review sessions, CARE also hosts regularly scheduled tutoring hours.

Tutors are available to answer questions, review problems, and help you feel prepared for your exam during these times:

Session 1: Sept 19, 4-6pm John, Jesse, and Ryan  
Session 2: Sept 20, 8-10pm John, Jesse, and Sofia

Can’t make it to a session? Here’s our schedule by course:

https://care.engineering.illinois.edu/tutoring-resources/tutoring-schedule-by-course/

Solutions will be available on our website after the last review session that we host, as well as posted in the zoom chat 30 minutes prior to the end of the session.

Step-by-step login for exam review session:

1. Log into Queue @ Illinois
2. Click “New Question”
3. Add your NetID and Name
4. Press “Add to Queue”
5. Join the zoom link in the staff message

Please do not log into the zoom call without adding yourself to the queue

Good luck with your exam!
**Naming:**
Write the corresponding name/chemical formula for the following compounds:

1. $\text{H}_2\text{SO}_4$
   - Sulfuric Acid (Di-hydrogen Sulfate)

2. $\text{Ca}_3\text{P}_2$
   - Calcium Phosphide

3. Ammonium Chlorate
   - $\text{NH}_3\text{ClO}_3$

4. Copper (II) Acetate
   - $\text{Cu(\text{CH}_3\text{COO})}_2$

5. Tetraphosphorus Hexoxide
   - $\text{P}_4\text{O}_6$
Lewis Structures:

1. Draw the corresponding Lewis structures for the following compounds:

I. HCN

\[
\begin{align*}
\text{H} & \quad \text{C} \equiv \text{N}: \\
\end{align*}
\]

II. CO\(_2\)

\[
\begin{align*}
\ddot{\text{O}} & \equiv \text{C} \equiv \ddot{\text{O}} \\
\end{align*}
\]

III. AsF\(_5\)

\[
\begin{align*}
\text{F} & \quad \text{F} \\
\text{F} & \quad \text{F} \\
\text{F} & \quad \text{F} \\
\end{align*}
\]
2. Which of the following statements regarding the Lewis structure below are FALSE?

\[
\begin{array}{c}
\text{O} \\
\text{H} - \\
\text{C} - \\
\text{C} \equiv \\
\text{C} - \\
\text{C} - \\
\text{N} - \\
\text{H} \\
1 \\
2 \\
3 \\
\text{H} \\
\text{H}
\end{array}
\]

A) An sp\textsuperscript{2} hybrid orbital from C-1 overlaps with an sp hybrid orbital from C-2 to form the sigma bond between C-1 and C-2.

B) This molecule has three π bonds.

C) Two of the atoms in this compound are sp\textsuperscript{3} hybridized.

D) The π bonds between C-2 and C-3 are formed from overlap of sp hybrid orbitals. \[\text{Correct Answer: D} \]

E) There are 10 sigma bonds in this molecule.

The answer is D because the Pi-bonds that are formed between C-2 and C-3 are formed because of the overlap between the unhybridized P-orbitals. The hybridized sp-orbitals are used to make a sigma bond, not a pi bond. Pi-bonds are always made because of an overlap of an unhybridized p-orbital, I don’t know if there are any exceptions to this. An option that might trip many people off might be option B. In this molecule there are 3 pi bonds because there is one pi-bond between the 2 unhybridized p-orbitals of C-1 and O. There are 2 pi-bonds between C-2 and C-3. This is because each of the carbon C-2 and C-3 have 2 unhybridized p-orbitals, due to their sp-hybridization. The 2 unhybridized p-orbitals on C-2 will interact with the 2 unhybridized p-orbitals on C-3 to form 2 pi-bonds, thus giving you a triple bond.
Orbitals/Electron Configuration:
Write the electron configuration for the following elements in shorthand notation:

1. O
   
   \[ \text{[He]}2s^22p^6 \]

2. Pb
   
   \[ \text{[Xe]}4f^{14}5d^{10}6s^26p^2 \]

3. K
   
   \[ \text{[Ar]}4s^1 \]

4. Ce
   
   \[ \text{[Xe]} 6s^24f^{15}d^1 \]

5. Rn
   
   \[ \text{[Xe]}4f^{14}5d^{10}6s^26p^6 \]

6. Cu
   
   \[ \text{[Ar]} 4s^13d^{10} \text{ (*EXCEPTION*)} \]

7. Mn
   
   \[ \text{[Ar]}3d^54s^2 \]

Periodic Trends:

1. Rank from smallest to largest atomic radii: (Li, O, Lu, Rf, He)

   \[ \text{He < O < Li < Rf < Lu} \]

2. Which has a higher ionization energy, Boron or Beryllium?

   Beryllium. *Trick Question* because it goes against the trend. Look at the electron configurations of each. Notice that Boron has one unpaired electron whereas Beryllium does not.

3. Rank from most to least ionization energy: (P, As, Te, O)

   \[ O > P > As > Te \]
Balanced Equations
Balance the following equation

\[ 2 \text{C}_8\text{H}_{18} + 25 \text{O}_2 \rightarrow 16 \text{CO}_2 + 18 \text{H}_2\text{O} \]

Electron Configuration and Ground State:
Circle the following elements that have 2 unpaired electrons in ground state.

\[
\begin{array}{cc}
\text{Cu}^+ & \text{Ni}^{2+} & \text{Zn}^{2+} & \text{Cr}^{2+} & \text{Ti}^{2+} \\
\text{Cu}^+: [\text{Ar}]4s^03d^{10} & \text{Ni}^{2+}: [\text{Ar}]4s^03d^8 & \text{Zn}^{2+}: [\text{Ar}]4s^03d^{10} & \text{Cr}^{2+}: [\text{Ar}]4s^03d^4 & \text{Ti}^{2+}: [\text{Ar}]4s^03d^2
\end{array}
\]

Bond Length:
Which bond has the shortest bond length and give an explanation as to why?
A) Single
B) Double Bond
C) Triple Bond

Triple bonds are stronger than double bonds, which are stronger than single bonds. The increased number of shared electrons results in more attraction between the atoms.

Chemical Changes/Polarity/Solubility:
How many of the following processes are examples of a chemical change?

(I) H\textsubscript{2}O (l) → H\textsubscript{2}O (g)
(II) I\textsubscript{2} (s) → I\textsubscript{2} (g)
(III) CH\textsubscript{4} (g) + 2O\textsubscript{2} (g) → CO\textsubscript{2} (g) + 2H\textsubscript{2}O (l)
(IV) C\textsubscript{6}H\textsubscript{12}O\textsubscript{6}(s) → C\textsubscript{6}H\textsubscript{12}O\textsubscript{6}(aq)
(V) 2H\textsubscript{2}O\textsubscript{2}(aq) → 2H\textsubscript{2}O(l) + O\textsubscript{2}(g)
A) 1 B) 2 C) 3 D) 4 E) 5
Electromagnetic Radiation:

1. For a hydrogen atom, how many of the following three electronic transitions are exothermic? Circle them.

A) \( n = 6 \) to \( n = 1 \)  
B) \( n = 2 \) to \( n = 3 \)  
C) \( n = 3 \) to \( n = 5 \)

2. When an electron in a 2p orbital of a lithium atom makes a transition to the 2s orbital, a photon of wavelength 670.8 nm is emitted. Calculate the energy difference between the 2p and 2s orbitals in lithium.

Calculate the frequency and then multiply by Plank’s Constant to solve for the energy difference:

\[
c = \nu \lambda
\]

\[
\nu = 4.437 \times 10^{14} \text{ Hz}
\]

\[
E = h \nu = \frac{c}{\lambda}
\]

\[
E = 2.96 \times 10^{-19} \text{ J}
\]

3. Calculate the change in energy for the \( n = 4 \) to the \( n = 2 \) transition in hydrogen.

\[
E = -R_1 Z^2 \left( \frac{1}{n_2^2} - \frac{1}{n_1^2} \right)
\]

where \( Z = 1, n_2 = 2, \) and \( n_1 = 4 \)

Plug in the values so \( E = -(2.178 \times 10^{-18})(1^2)(\frac{1}{2^2} - \frac{1}{4^2}) \)

\[
E = -4.084 \times 10^{-19} \text{ J}
\]

4. Does a visible light (\( \lambda = 400 - 700 \text{nm} \)) photon have enough energy to excite a H electron from \( n=1 \) energy to \( n = 6 \) energy state?

Need to find maximum amount of energy from visible light. Lowest wavelength provides highest energy.

\[
E = \frac{hc}{\lambda} \rightarrow E = \frac{hc}{400 \text{nm}} = 4.966 \times 10^{-19} \text{ J}
\]

Same equation as 3, but with \( Z = 1, n_2 = 6 \) and \( n_1 = 1 \)

Plug in values and solve to find \( E = 2.1775 \times 10^{-18} \text{ J} \) which is the energy required to move an electron from energy level 1 to energy level 6.

The energy from visible light (\( 4.966 \times 10^{-19} \text{ J} \)) is less than the required amount. Therefore, the answer is visible light does NOT have enough energy to excite a H electron from energy level 1 to 6.
Shape/Geometry of Compounds:

1. How many of the following compounds have a square planar shape?

   A) KrF₂
   B) PCl₅
   C) XeF₄
   D) TeF₄
   E) ICl₃

2. Which of the following compounds will be least soluble in water? Hint: Water is a polar solvent and “like dissolves like”

   A) KrF₂ (Only Non-Polar Compound Listed)
   B) PF₃
   C) IF₅
   D) COS
   E) SO₂