Chem 1A Midterm 2 Review Packet

Kim’s section:

1.) How many hydrogen atoms are there in 25.0 g of CH2Cl2?

\[
25.0 \text{ g CH}_2\text{Cl}_2 \times (1\text{ mol CH}_2\text{Cl}_2/84.93 \text{ g CH}_2\text{Cl}_2) \times (2\text{ mol H/1 mol CH}_2\text{Cl}_2) \times 6.022 \times 10^{23}
\]

atoms H = \(3.55 \times 10^{23}\) atoms H

2.) Calculate the formal charges of the oxygen and sulfur atoms in \(\text{H}_2\text{SO}_4\) and draw the best Lewis electron-dot structure.

\text{Formal charge} = \#\text{ of valence electrons} - \#\text{ of electrons in lone pairs} - \frac{1}{2}\text{ the number of bonding electrons}

\[
\begin{array}{l}
\text{O: 6-6-1} = (-1) \\
\text{S: 6-6-1} = (-1) \\
\text{O: 6-0-4} = (+2)
\end{array}
\]

\[
\begin{array}{l}
\text{O: 6-4-2} = 0 \\
\text{S: 6-6-0} = 0 \\
\text{O: 6-4-2} = 0
\end{array}
\]

Formal charge formula: expected valence electrons – # of electrons in lone pairs - # of bonds

*One on the right is the best Lewis structure because it is neutral.*
3.) Draw three resonance structures for the molecule nitrous oxide, N$_2$O (the atomic arrangement is NNO). Indicate formal charges. Rank the structures in their relative importance to the overall properties of the molecule.

Resonance structures that are happy will have fewer formal charges and the formal charges should be on atoms that “make sense.” For instance, oxygen is highly electronegative so having a formal charge of + on oxygen would not result in a happy molecule. Oxygen would rather have a – charge. In this case, the most stable resonance structure would be structure number 3 because one nitrogen has a formal charge of 0, one nitrogen has a charge of +1, and oxygen has -1. This makes sense, unlike the other structures that depict the charges incorrectly. It’s best to have the central atom have no formal charge, but if that is not possible, make sure to manipulate your resonance structure to have the least number of formal charges. **Ranking of structures in their relative importance to the overall properties of the molecule: I < II < III (III is the best resonance structure!!!)**

4.) For each of the following pairs of ions, write the formula of the corresponding compound.
   a. Fe$^{3+}$ and CN$^- = \text{Fe(CN)}_3$
   b. K$^+$ and SO$_4^{2-} = \text{K}_2\text{SO}_4$
   c. Li$^+$ and N$_3^{3-} = \text{Li}_3\text{N}$
   d. Ca$^{2+}$ and P$^{3-} = \text{Ca}_3\text{P}_2$

5.) In which of the following are the name and formula correctly paired?
   a. sodium sulfite: Na$_2$S
   b. calcium carbonate: Ca(CO$_3$)$_2$
c. magnesium hydroxide Mg(OH)$_2$
d. nitrite: NO$_2$
e. iron (III) oxide: FeO

*Memorize all polyatomic ions and know how to name compounds!

6.) For which element is the gaining of an electron most exothermic?
   a. Li
   b. N
c. F
d. B

Electron affinity: The energy associated with an element in its gaseous state gaining an electron – does not show a general trend as we move down a column in the periodic table, but it generally becomes more negative (more exothermic) to the right across a row.

The electron gain enthalpies of the halogens elements are most exothermic, this due to the reason that the valance shell electronic configuration of the halogens is $ns^2np^5$ and such they require one more electron to acquire the stable noble gas configuration, i.e. $ns^2np^6$. Therefore, fluorine (F) is the gaining of an electron most exothermic. Thus, the correct option is (c).

In summary: most exothermic follows the trend for electron affinity and ionization energy

7.) The species F$^-$, Ne, and Na$^+$ all have the same number of electrons. Which is the predicted order when they are arranged in order of decreasing size (largest first)?
   a. F$^-$ > Ne > Na$^+$
b. Ne > Na$^+$ > F$^-$
c. Na$^+$ > F$^-$ > Ne
d. F$^-$ > Na$^+$ > Ne

Looking at the trend for atomic radius on the periodic table, we can see that these ions all have similar “size.” They all have the same number of electrons. However, a negatively charged ion will always be “bigger” and a positively charged ion will always be “smaller.”

Jasmine’s section:

8.) Write an orbital diagram for the ground state of the potassium atom. Is the atomic substance diamagnetic or paramagnetic?

![Orbital diagram](image_url)

Paramagnetic because the 4s subshell has an unpaired electron.
9.) The ground-state electron configurations listed here are incorrect. Explain what mistakes have been made in each and write the correct electron configurations.

(I) Al: 1s²2s²2p⁴3s²3p³
Aluminum has 13 electrons, which matches with the number of electrons in the question, but the 2p subshell is not filled. The 2p subshell can have 6 electrons, but it only has 4 electrons. This violates the Aufbau Principle, which states that the lowest energy level is filled first before going to the next energy level.

Correct Electron configuration: 1s²2s²2p⁶3s²3p¹

(II) B: 1s²2s²2p⁵
Boron has 5 electrons, which does not match with the number of electrons in the question (9 electrons). Therefore, the 2p subshell should have 1 electron, not 5 electrons.

Correct Electron configuration: 1s²2s²2p¹

(III) F: 1s²2s²2p⁶
Fluorine has 9 electrons, which does not match with the number of electrons in the question (10 electrons). It has an extra electron. Therefore, the 2p subshell should have 5 electrons, not 6 electrons.

Correct Electron configuration: 1s²2s²2p⁵

10.) Zirconium is a Group IVB element in Period 5. What would you expect for the valence-shell configuration of zirconium?

4d⁵5s²
The valence electrons are the electrons involved in bonding. For transition metals, the valence electrons are in the (n-1)d subshell and ns subshell. # Valence electrons = sum of electrons in the (n-1)d subshell + ns subshell. This equation only applies to the transition metals. For representative elements (Group 1A-7A), the equation is # of valence electron = ns electrons + np electrons. Zirconium has four valence electrons.

11.) Write Lewis formulas for each of the following:

(I) BCl₃
(II) TiCl₂⁺
(III) BeBr₂
12.) Write Lewis electron-dot structures for the following:
   (I) GaCl₄⁻
   (II) BF₅

13.) One of the following compounds has a carbon–nitrogen bond length of 116 pm; the other has a carbon–nitrogen bond length of 147 pm. Match a bond length with each compound.
Methylamine has a bond length of 147 pm and acetonitrile has a bond length of 116 pm. The carbon-nitrogen bond in Methylamine is a single bond while acetonitrile has a triple bond. Triple bonds are shorter than single bonds.

14.) What is the bond order in $\text{O}_2^{2-}$?
   a) 2.0  
   b) 3.5  
   c) 1.0  
   d) 0  
   e) 1.5

\[
\text{Bond order} = \frac{\# \text{ of bonds}}{\# \text{ of regions}} = \frac{1 \text{ bond}}{1 \text{ region}} = 1.0
\]

* # of regions = # of resonance structures

Jonathan’s section:

15.) Which of the following is a covalent compound? (multiple choice)
   A) NaOH  
   Incorrect, because Na has a +1 charge and OH has a -1
charge, so they are able bond through their opposite charges, making them an ionic compound

B) \((\text{NH}_4)_2\text{SO}_4\) Incorrect, even though there are no metals in this compound, \(\text{NH}_4\) has a +1 charge, while \(\text{SO}_4\) has a -2 charge, which makes it an ionic compound

C) \(\text{N}_2\text{O}_4\) Correct, because it is the only answer choice where charges do not influence the bond

D) \(\text{K}_2\text{CO}_3\) Incorrect, because the K has a +1 charge, while the CO3 has a -2 charge, meaning the two are able to bond together through their opposite charges.

16.) Place the following in order of increasing electronegativity.

i) B, Rb, O, Ga \(\text{Rb} < \text{Ga} < \text{B} < \text{O}\)

ii) N, K, F, Be \(\text{K} < \text{Be} < \text{N} < \text{F}\)

iii) Ba, Si, Ca, S \(\text{Ba} < \text{Ca} < \text{Si} < \text{S}\)

iv) Te, Sn, F, Br \(\text{Sn} < \text{Te} < \text{Br} < \text{F}\)

Electronegativity Trend: Up and to the Right

17.) Place the following in order of increasing atomic or ionic radii:

i) \(\text{Ge}^-, \text{Si}^-, \text{Sn}^-, \text{C}^-\) \(\text{C}^- < \text{Si}^- < \text{Ge}^- < \text{Sn}^-\); The highest principal number of the valence electrons increases and therefore valence electrons occupy larger orbitals

ii) \(\text{O}^-, \text{O}, \text{O}^2-, \text{O}^+\) \(\text{O}^+ < \text{O} < \text{O}^- < \text{O}^{2-}\); Anions are larger than cations because there are more electrons, meaning there’s less pull for each electron

iii) \(\text{S}^{2-}, \text{Cl}^-, \text{Ca}^{2+}, \text{K}^+\) \(\text{Ca}^{2+} < \text{K}^+ < \text{Cl}^- < \text{S}^{2-}\); While these are all isoionic, anions are bigger than cations

iv) \(\text{Co}, \text{Co}^+, \text{Co}^{2+}, \text{Co}^{3+}\) \(\text{Co}^{3+} < \text{Co}^{2+} < \text{Co}^+ < \text{Co}\); Cations are smaller than their neutral counterparts

18.) What is the correct electron configuration for \(\text{Fe}^{2+}\)? (multiple choice)

A) \(1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 4d^6\)

B) \(1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6\) The 2 electrons lost are from the highest n-value
orbitals first, which in this case is the 4s orbital. Therefore, the 3d\(^6\) remains untouched.

C) \(1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^6 \ 4s^2 \ 3d^4\)

D) \(1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^6 \ 4s^2 \ 3d^{10} \ 4p^6 \ 5s^2 \ 4d^4\)

19.) Write the electron configurations for the following. The shorthand version can be utilized

i) Cr\(^{3+}\) \(1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^6 \ 3d^3\) or \([Ar\] 3d\(^3\)\)

ii) Fe\(^{3+}\) \(1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^6 \ 3d^6\) or \([Ar\] 3d\(^5\)\)

iii) Cu\(^{+}\) \(1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^6 \ 3d^{10}\) or \([Ar\] 3d\(^{10}\)\)

iv) Co\(^{2+}\) \(1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^6 \ 3d^7\) or \([Ar\] 3d\(^7\)\)

20.) Which neutral atom is isoelectronic with N\(^{3-}\)? (multiple choice)

A) P

B) Cl

C) Ne Has the same number of electrons as N\(^{3-}\) (10 electrons)

D) O

Isoelectronic: having the same numbers of electrons or the same electronic structure

21.) The +3 cation of an unknown element, X, has the following as its outermost electron configuration in its ground state:

What is element X?

Element X is Br because in its cation state, it only has 2 electrons in the 4p orbital. However, in its ground state, it has 5 electrons in the 4p orbital because it is a 3+
cation. Therefore, element X must be Br because that is the only element with 5 electrons in the 4p orbital.
## Fundamental Constants

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<td>Avogadro's Number ($N_A$)</td>
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<td>Electron Mass</td>
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<td>Faraday Constant ($F$)</td>
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<td>Gas Constant ($R$)</td>
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<td>Speed of light in vacuum ($c$)</td>
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## Some Prefixes Used with SI Units

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## Useful Conversion Factors and Relationships

\[
1 \text{ atm} = 760 \text{ mm Hg} = 760 \text{ torr} = 101,325 \frac{N}{m^2} = 101,325 \text{ Pa}
\]

\[
1 \text{ cal} = 4.184 \text{ J (exactly)}
\]

\[
1 \text{ L•atm} = 101.325 \text{ J}
\]

\[
1 \text{ J} = 1 \text{ C} \times 1 \text{ V} = 1 \text{ C} \cdot \text{V}
\]

\[
1 \text{ amu} = 1.66054 \times 10^{-24} \text{ g}
\]

\[
K = (°C + 273.15)
\]
Equations

\[ \nu = \frac{c}{\lambda} \]
\[ \Delta x \cdot \Delta p \geq \frac{\hbar}{4\pi} \]
\[ E = h\nu \]
\[ E_n = -R_H \left( \frac{Z^2}{n^2} \right) \]
\[ E = mc^2 \]
\[ E = \frac{hc}{\lambda} \]
\[ \Delta E = -R_H \left( \frac{Z^2}{n_f^2} - \frac{Z^2}{n_i^2} \right) \]
\[ \lambda = \frac{h}{mv} \]

PERIODIC TABLE OF THE ELEMENTS

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