Chemistry 1A03 Fall 2010
Test 1 Version 1 Solutions

1. What is the name of the element with **atomic number seven**?

A) nickel
B) neodymium
C) nitrogen
D) neptunium
E) neon

2. A student titrates NaOH (buret) against HCl (Erlenmeyer flask), with phenolphthalein indicator, as in experiment # 1. At the end of the titration, the solution in the Erlenmeyer flask is **dark pink**. If the actual concentration of NaOH is 0.1029 M, what is the **most likely** result obtained by the student?

A) 0.07458 M
B) 0.1029 M
C) 0.1358 M
D) 0.1008 M
E) 0.1035 M

Because the solution is dark pink, there is an excess of unreacted hydroxide left in the flask. The student's result does not account for all of the sodium hydroxide – i.e. the sodium hydroxide concentration is underestimated. Two of the above choices (D and E) correspond to less than the true sodium hydroxide concentration. It is unlikely that the student is as far off as answer E suggests. Therefore, D is the best answer – it is typical of results actually obtained by some students in this experiment.

3. In a lab experiment, a student adds aqueous sodium hydroxide (10.0 mL, 2.0 M) to aqueous cobalt(II) nitrate (50.0 mL, 0.30 M). When the reaction reaches completion, what **mass (in grams)** of cobalt(II) hydroxide is formed? Assume that the only reaction occurring is:

$$Co(NO_3)_2(aq) + 2\ NaOH(aq) \rightarrow Co(OH)_2(s) + 2\ NaNO_3(aq).$$

A) 0.74
B) 3.9
C) 2.9
D) 0.93
E) 1.9

Number of moles of NaOH $= 0.010\ L \times 2.0\ mol\ L^{-1} = 0.020$ mol
Number of moles of $Co(NO_3)_2$ $= 0.050\ L \times 0.30\ mol\ L^{-1} = 0.015$ mol

To consume 0.015 mol of $Co(NO_3)_2$ requires 0.030 mol of NaOH which is more than we have. NaOH is the limiting reagent – i.e., all 0.020 mol of NaOH is consumed to produce 0.010 mol of $Co(OH)_2(s)$ (2:1 in accord with balanced equation stoichiometric coefficients).
Mass of $Co(OH)_2(s)$ produced = 0.010 mol × (58.933 + 2 (15.999+1.0079) g mol^{-1}) = 0.93 g

4. Temperatures from the distant past can be determined by measuring the ratio of ^{18}O to ^{16}O in the layers of ice cores taken from Antarctic or Greenland ice sheets. Neutral atoms of ^{16}O, ^{17}O and ^{18}O **all** have

A) 16 protons
B) 8 electrons and 8 protons
C) 8 neutrons and 8 protons
D) 16 protons and 16 electrons
E) 8 neutrons

All O atoms have 8 protons. If the atom is neutral, there are also 8 electrons. The number of neutrons varies for different isotopes.

5. Determine the **FALSE** statements regarding the **general** trends of the periodic table.

A) The second ionization energy for Na will be larger than the first ionization energy of Ne.
B) An atom of element X will have a larger radius than X^+.
C) An atom of element X will have a smaller radius than X^-.
D) The element with the smallest first ionization energy in a row will have the largest atomic radius of that period.
E) Z increases across a period, and Z_{eff} decreases across a period.

Z_{eff} **increases** across a period. In simple terms it equals the group (A) number. The other statements are true.

6. **Identify** the element based on the following information:
i) It has a lower melting point than Al.
ii) It has a higher magnitude of electron affinity than Ge.
iii) It has a smaller atomic radius than P.
iv) It has a lower first ionization energy than F.

A) He
B) O
C) W
D) Rb
E) Si

Of these elements, only Si and O have higher magnitude of electron affinity than Ge. Si has a higher melting point than Al (it has the same covalent network structure as diamond). Si is also bigger than P, while O is smaller.

7. Which of the following elements would have the **least** metallic character?

A) As
B) S
C) Mg
D) Ga
E) Cr

Metallic character decreases up and to the right in the periodic table.

8. The **order of increasing electronegativity** (from the smallest to the largest value) for oxygen, silicon and sulfur is

A) O < Si < S
B) S < Si < O
C) S < O < Si
D) Si < O < S
E) Si < S < O

Electronegativity increases up and to the right in the periodic table.
... C N O F Ne
... Si P S ... And so: S > Si, O > S.

9. Phosphorus has a **higher** first ionization energy than sulfur because

A) There is additional stability in phosphorus because of a half-filled subshell.
B) Phosphorus is a noble gas.
C) Phosphorus is in a different period than sulfur.
D) The statement is not true; sulfur has the higher first ionization energy.
E) Phosphorus exists as P_4 in the elemental state, whereas sulfur is S_8, meaning more electrons need to be removed, requiring a larger energy.

Phosphorus does have a higher first ionization energy than sulfur. This is because of the special stability of its half-filled p subshell valence electron configuration. Ionizing P spoils this configuration, whereas ionizing S actually achieves (in S^+) this special configuration. The latter makes ionizing S more favourable – i.e., lowering its ionization energy.

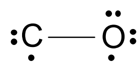
10. Which of the following atoms/cations would be the **hardest** to further ionize?

A) Ca^{2+}
B) Na
C) P^+
D) Al^{2+}
E) K^+

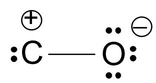
Ca^{2+} has the noble gas electron configuration (i.e. that of Ar). Ionizing Ca^{2+} requires removing an electron from another shell which is much lower in energy. Furthermore, Ca^{2+} is already a 2+ cation which is hard to ionize in any case; in contrast for Al^{2+} there is still 1 valence electron available.

11. Carbon monoxide is a gas which binds to hemoglobin in competition with oxygen. Which one of the following pictures represents the **correct charge-minimized Lewis structure** for **CO**?

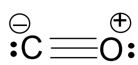
A)



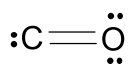
B)



C)



D)

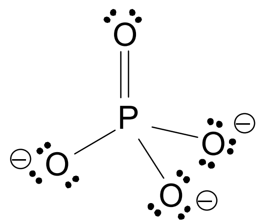


E)

In structures A, B and D, carbon has less than an octet. In structure E, carbon has 10 electrons in its valence shell. As carbon is in period two, it can only have up to 8 valence electrons.

12. Choose the **TRUE** statements regarding the charge-minimized structure of the phosphate ion, PO_4^{3-}.
(i) The average formal charge on O is –1/4.
(ii) The average P–O bond order is 5/4.
(iii) There are 4 equivalent resonance structures.
(iv) The O–P–O bond angles will be < 109.5°.

A) i, ii
B) ii, iii
C) i, iii
D) i, iv
E) ii, iv



The average formal charge on O is –3/4. The four P-O bonds are equivalent – there are four resonance structures correspond to the four possible positions for the double bond. The true structure is a hybrid of all resonance forms. All bond angles are the same, and equal to 109.5°.

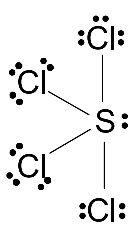
13. Choose the **FALSE** statement about the charge-minimized Lewis structures for the molecules $\mathbf{SO_2}$ and $\mathbf{SCl_2}$.

A) There are two resonance structures for SO_2.
B) The bond angle in SCl_2 is < 109.5°.
C) The VSEPR class for SO_2 is AX_2 E.
D) The formal charge on S in SCl_2 is zero.
E) The molecular geometry for both SO_2 and SCl_2 is bent.

There is only one structure for SO_2. All other statements are true.

14. What is the **name** of the **molecular shape** of $\mathbf{SCl_4}$?

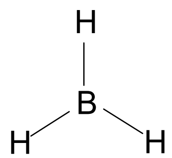
A) Seesaw
B) Square pyramidal
C) Square planar
D) T-shaped
E) Tetrahedral



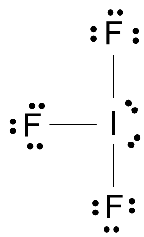
AX_4E VSEPR class = trigonal bipyramidal electron pair geometry with one lone pair
= seesaw molecular shape

15. Choose the **FALSE** statement regarding the following molecules and ion: BH_3, IF_3, NI_3, NO_3^-.

A) The central atom of only one molecule has an expanded octet.
B) Only one species will have non-ideal bond angles.
C) Only one of the molecules has a trigonal pyramidal molecular shape.
D) Two of the molecules have a permanent dipole.
E) Exactly two of the species shown have the same molecular shape.



AX_3 trigonal planar

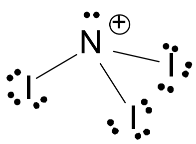


AX_3E_2 T-shaped

non-ideal bond angles

expanded octet on I

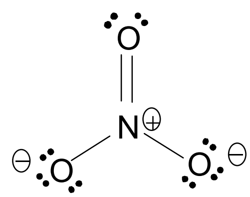
permanent dipole



AX_3E trigonal pyramidal

non-ideal bond angles

permanent dipole



AX_3 trigonal planar

16. A detector receives a signal consisting of green light, with a wavelength of 540 nm. The total energy of the signal is 2.50×10^{-14} J. **How many** photons reach the detector?

A) 5.50×10^{2}
B) 1.50×10^{4}
C) 8.40×10^{7}
D) 6.80×10^{4}
E) 2.50×10^{-5}

The energy of one photon with wavelength 540 nm is
$E = h\,c\,/\,\lambda\ =\ [\ 6.6256 \times 10^{-34}\ \text{Js} \times 2.9979 \times 10^{8}\ \text{m/s}\ ]\,/\,540 \times 10^{-9}\ \text{m}$
$=\ 3.678 \times 10^{-19}$ J (one extra digit is kept to avoid rounding error)

Number of photons = Total energy of photons / Energy of one photon
$=\ 2.50 \times 10^{-14}\ \text{J}\,/\,3.678 \times 10^{-19}\ \text{J}\ =\ 6.80 \times 10^{4}$ photons

17. When electromagnetic waves interact with matter, **they may**

A) cause heating
B) cause ionization
C) cause electronic transitions
D) break bonds
E) all of the above

When electromagnetic waves (namely light) are absorbed by matter, the energy of the light is converted to additional energy in the matter. This energy can end up in the form of heat. Absorption of a photon can cause an electronic transition wherein an electron is excited to a higher energy level. If the photon energy exceeds the ionization threshold, then ionization can result. If the photon energy absorbed by a molecule exceeds a bond energy, then a bond can be broken.

18. The energy difference between the n = 1 state and the n = 2 state in the He^+ ion is 6.534×10^{-18} J, which corresponds to a photon wavelength of 30.40 nm. Which of the following **photon wavelengths**, when **absorbed**, will cause excitation from n = 1 to n = 2 in He^+?

A) Any wavelength less than or equal to 30.40 nm
B) Only 30.40 nm
C) Any wavelength greater than or equal to 30.40 nm
D) Any wavelength that is an integer multiple of 30.40 nm
E) Any integer multiple of 6.534×10^{-18} J, converted to a wavelength

To cause a specific electronic transition (here, from n = 1 to n = 2 in He^+), the absorbed photon energy must exactly match the energy of the transition. Correspondingly, the photon wavelength must be exactly the right value for the transition (here, 30.40 nm).

19. An arsenic atom (Z = 33) is in its ground state. Which one of the following sets of quantum numbers (n, ℓ, m_ℓ, m_s) could **not** possibly describe one of its electrons?

A) 4, 2, 2, –½
B) 3, 2, –2, ½
C) 2, 1, –1, ½
D) 3, 0, 0, –½
E) 4, 1, 0, ½

Ground state electron configuration of As:
$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^3$
(n, ℓ, m_ℓ, m_s) = (4, 2, 2, –½) corresponds to a 4d electron that does not appear in the ground state electron configuration of As.

20. Which of the following statements is **FALSE**?

- **A)** Nitrogen atoms in their ground state are paramagnetic.
- **B)** Calcium atoms in their ground state are paramagnetic.
- **C)** $[Ar]4s^1 3d^1$ is the electron configuration of an excited state of a Ca atom.
- **D)** $[He]2s^2\,2p^5$ is the electron configuration of the ground state of a F atom
- **E)** $[He]2s^2$ is the electron configuration of the ground state of a Be atom.

Ground state electron configurations:
Be: $[He]\,2s^2$
N: $[He]\,2s^2\,2p^3$ paramagnetic (the 3 2p electrons all unpaired – by Hund's rule)
F: $[He]\,2s^2\,2p^5$
Ca: $[Ar]\,4s^2$ **diamagnetic** (the 2 4s electrons are paired) – i.e., **not paramagnetic**

An excited state electron configuration:
Ca: $[Ar]\,4s^1 3d^1$

21. The percentage by mass of bicarbonate in an antacid tablet is 32.5 %. Calculate the volume of carbon dioxide gas (**in mL**) generated at 37 °C and 1.00 atm from a 3.29 g tablet according to the reaction:

$$HCO_3^-(aq) + H^+(aq) \rightarrow H_2O(l) + CO_2(g)$$

- **A)** 53
- **B)** 618
- **C)** 446
- **D)** 1370
- **E)** 27.2

32.5 % (by mass) bicarbonate in a 3.29 g antacid tablet means ...
Mass of HCO_3^- = 3.29 g × 0.325 = 1.069 g (keep an extra digit to avoid rounding error).
Moles of HCO_3^- = Mass of HCO_3^- / Molar mass of HCO_3^-
= 1.069 g / (1.0079 + 12.011 + 3×15.999 g mol^{-1}) = 0.01752 mol

From the ideal gas law, we have
Volume = $V = nRT/P$ = 0.01752 mol × 0.08206 L atm K^{-1} mol^{-1} × 310.15 K / 1.00 atm
= 0.446 L = 446 mL

22. The density of a noble gas is measured to be 2.71 g L^{-1} at 3.00 atm and 0 °C. **What is the gas**?

A) Ar
B) He
C) Kr
D) Xe
E) Ne

The ideal gas law, $PV = n\,RT$, can be written as

$P/RT = n/V = (m/M)/V = d/M$, where *m* is the mass of the gas and *M* is its molar mass. $d = m/V$ is the density of the gas. Therefore,

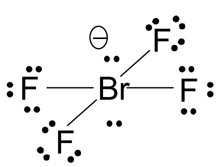
$M = d\,RT/P$ = 2.71 g L^{-1} × 0.08206 L atm K^{-1} mol^{-1} × 273.15 K / 3.00 atm
= 20.3 g mol^{-1} .

This is close to the molar mass of Ne, 20.2 g mol^{-1} , the difference attributable to experimental error.

23. Choose the **TRUE** statements regarding the **BrF_4^-** ion.

(i) The electron pair geometry is octahedral.
(ii) The formal charge on Br is zero.
(iii) The molecular shape is square planar.

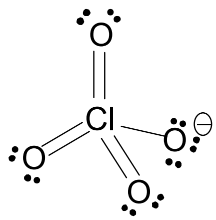
A) i, ii, iii
B) iii only ← 1.5 marks for this choice
C) i, ii
D) i, iii
E) ii, iii



The VSEPR class is AX_4E_2 . The molecule has an octahedral electron pair geometry. Its molecular shape is square planar. Bromine has a formal charge of −1.

24. Choose the **FALSE** statement about the charge-minimized Lewis structures of the **ClO_4^-** and **ClO_3^-** ions.

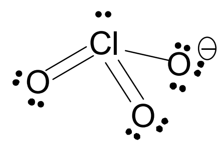
A) The average Cl-O bond order is lower for the ClO_3^- ion.
B) There are more resonance structures for the ClO_4^- ion than for the ClO_3^- ion.
C) The average Cl-O bond length is greater in ClO_3^-.
D) The average Cl-O bond energy is greater for the ClO_4^- ion.
E) The average formal charge on the O atoms in ClO_4^- is –1.



Average Cl-O bond order = (2+2+2+1)/4 = 7/4

Average formal charge on O = (0+0+0–1)/4 = –1/4

There are four equivalent O atoms. This means four equivalent resonance structures (the single bond could be any of the four Cl-O bonds).



Average Cl-O bond order = (2+2+1)/3 = 5/3 < 7/4

These Cl-O bonds are longer and weaker (lower bond energy) than those of ClO_4^-.

Average formal charge on O = (0+0–1)/3 = –1/3

There are three equivalent O atoms. This means three equivalent resonance structures (the single bond could be any of the three Cl-O bonds).

25. It takes 492 kJ of energy to remove one mole of electrons from the atoms on the surface of solid gold. What is the **speed of the ejected electrons (in m/s),** if the incoming light has a wavelength of 200.0 nm?

A) 2.33×10^5
B) 8.56×10^7
C) 6.22×10^5
D) 1.63×10^2
E) 5.80×10^6

Energy to remove one mole of electrons = 492 kJ = 492×10^3 J
Energy to remove one electron = 492×10^3 J / 6.022×10^{23} = 8.17×10^{-19} J
Energy of one photon with a wavelength of 200.0 nm = $h\,c\,/\,\lambda$
= [6.6256×10^{-34} Js × 2.9979×10^8 m/s] / 200.0×10^{-9} m = 9.931×10^{-19} J

Photon energy = Energy to remove electron + Kinetic energy of ejected electron
Therefore,
Kinetic energy of ejected electron = 9.931×10^{-19} J − 8.17×10^{-19} J = 1.76×10^{-19} J

$= m_e u^2 / 2$, where m_e is the mass of the electron and u is its speed

Therefore,

$$u = [2 \times 1.76 \times 10^{-19} \text{ J} / 9.10 \times 10^{-31} \text{ kg}]^{1/2} = 6.22 \times 10^5 \text{ m s}^{-1}$$