General Chemistry - Review Questions

1. A motorcycle emits 9.5 g of carbon monoxide per kilometer driven. How many pounds (lb) of carbon monoxide does the motorcycle generate over 5.0 years if the motorcycle is driven 15 000 miles per year? (1 km = 0.6214 mi; 1lb = 0.4536 kg).

2. Pyrite is called fool's gold because it looks like real gold. However, pyrite has a density of 4.5 g/mL while gold has a density of 19.3 g/mL. Use this information to determine which of the following statements is true. Support your choice.
   a. 25 grams of gold will occupy a greater volume than 25 grams of pyrite.
   b. 25 grams of gold will occupy the same volume as 25 grams of pyrite.
   c. 25 mL of gold will have less mass than 25 mL of pyrite.
   d. 25 mL of gold will have a greater mass than 25 mL of pyrite.

3. Either give the formula or the name of the following species:
   a. Nickel (II) phosphate
   b. hydrogen perchlorate
   c. magnesium thiosulfate
   d. NH_3
   e. Ca_3P_2
   f. $Fe(NO_3)_2$

4. For each of the following reactions:
   a. Balance (if needed).
   b. Name each compound, or give the chemical formulas where needed.

   a. $NaCl_{(aq)} + AgNO_{3(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$
   b. $CH_{4(g)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(l)}$
   c. $Zn_{(s)} + O_{2(g)} \rightarrow ZnO_{(s)}$
   d. Magnesium solid + hydrogen chloride solution → Magnesium chloride solution + hydrogen gas
   e. Sodium hydroxide solution + hydrogen sulfate solution → sodium sulfate solution + water

5. A compound contains 25.94% N and 74.6% O (by mass). What is the empirical formula?

6. When 12.00 g of calcium metal is reacted with water, 12.00 g of calcium hydroxide is produced. Use this balanced equation to calculate the percent yield for the reaction.
$$Ca(s) + 2\ H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$$

7. According to the chemical equation shown below, 4.0 g of iron is reacted with 4.0 g of water.
$$3\ Fe(s) + 4\ H_2O(l) \rightarrow Fe_3O_4(s) + 4\ H_2(g)$$
   a. What is the limiting reactant?
   b. How many grams of Fe_3O_4 are produced?
   c. What is the mass (in grams) of the excess reactant left over?

8. Write a balanced net ionic equation for the reaction of $AgNO_3(aq)$ with $Cu(s)$.

9. What is the oxidation number of the chromium atom in $K_2Cr_2O_4$?

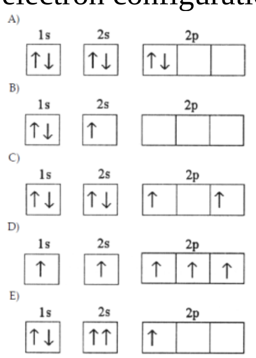
10. Which species functions as the oxidizing agent in the following reduction-oxidation reaction?

$$5\ Fe^{2+}(aq) + MnO_4^-(aq) + 8\ H^+(aq) \rightarrow Mn^{2+}(aq) + 5\ Fe^{3+}(aq) + 4\ H_2O(l)$$

11. Using the Balmer-Rydberg equation, an electron will undergo which of the following transitions when electromagnetic radiation with a 486.1 nm wavelength (λ) is absorbed?
    a. $m = 2 \rightarrow n = 3$
    b. $m = 2 \rightarrow n = 4$
    c. $n = 3 \rightarrow m = 2$
    d. $n = 4 \rightarrow m = 2$

12. True or False: The uncertainty principle states that it is impossible to know the exact position of an electron.

13. Which electron configuration represents a violation of the Pauli Exclusion Principle?



14. (a) What are the shapes of s, p, and d orbitals respectively?
    (b) How many 1s orbitals are there in an atom? How many 4p orbitals?
    (c) What is the maximum number of orbitals with n = 3 and l = 2?

15. An element in a ground state electron configuration has 4 electrons in the 4p orbitals. Which of the following statements **cannot** describe the electron configurations in this atom?
    a. At least one electron has an orbital angular momentum (*l)* of 2.
    b. Six electrons are in the n=4 shell.
    c. The valence electron configuration is identical to carbon.

16. Atoms of which element have the orbital-filling diagram shown below?

17. Which element has the ground-state configuration of $[Ar]4s^1 3d^5$?

18. Which ion has the same electron configuration as Kr?
    a. Rb^+ b. Br^- c. Se^{2-} d. All of these

19. Write the electron configurations for each of the following species:
    a. In and In^{3+} b. Rh and Rh^{3+}

20. Arrange these ions N^{3-}, O^{2-}, Mg^{2+}, Na^{+}, and F^{-} in order of increasing ionic radius, starting with the smallest first.

21. Of the following, which element has the highest first ionization energy? Support your answer.
    iodine bromine fluorine chlorine

22. Which of the following represents the change in electronic configuration that is associated with the second ionization energy of magnesium? Support your choice.
    a. $[Ne]3s^2 \rightarrow [Ne]3s^1 + e^-$
    b. $[Ne]3s^2 \rightarrow [Ne] + 2\ e^-$
    c. $[Ne]3s^1 \rightarrow [Ne] + e^-$
    d. $[Ne]3s^2 + e^- \rightarrow [Ne]3s^23p^1$

23. Is the following statement "True or False"? If false, correct the underlined part.
    "The general trend for ionization energy and electron affinity values is that they <u>both decrease as one goes across a period</u> from left to right and <u>decrease as one goes up a group</u>."

24. Which molecule contains the most easily broken carbon-carbon bond?
    a. $H_3C–CH_3$
    b. $H_2C=CH_2$
    c. $F_2C=CF_2$
    d. $HC\equiv CH$

25. Between CF_4 and CI_4, carbon tetrafluoride has stronger bonds. Why?

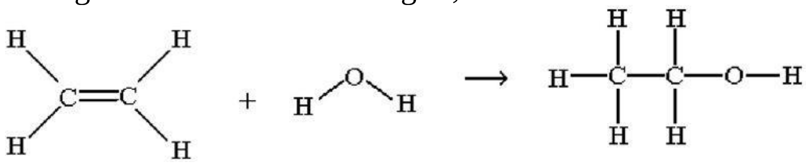
26. The electronegativity is 2.1 for H and 3.0 for N. Based on these electronegativities, NH_4^+ would be expected to:
    a. be ionic and contain H^- ions.
    b. be ionic and contain H^+ ions.
    c. have polar covalent bonds with partial negative charges on the H atoms.
    d. have polar covalent bonds with partial positive charges on the H atoms.

27. Draw all possible resonance structures for ClO_3^-. For each structure, calculate formal charges for each atom and identify which atom has the negative charge.

28. Draw all possible resonance structures for $SOCl_2$. For each resonance structure, assign formal charges to all the atoms. Which structure is the most stable?

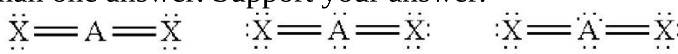
29. How many lone pairs are on the Br atom in IBr_2^-? Show your work.

30. One method for making ethanol, C_2H_5OH, involves the gas-phase hydration of ethylene (C_2H_4) as shown in below. Estimate ΔH for this reaction from the given average bond dissociation energies, D.



Bond	D, kJ/mol
C=C	615
C–H	410
C–C	350
C–O	350
O–H	460

31. Which of the following molecules would have a linear geometry? There could be more than one answer. Support your answer.



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32. What is the smallest bond angle in SF_6?

33. Which orbital hybridization is associated with a tetrahedral cloud arrangement?

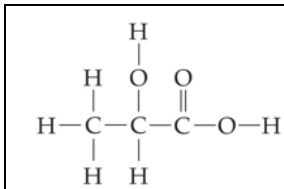
34. What is the hybridization on the N atom in NO_2^- and in NO_3^-?

35. What is the hybridization on the C atom in CO_2 and in HCN? What is the molecular shape?

36. When melting S_8, _________ forces must be overcome and S_8 is expected to have a _________ melting point than MgS.
   a. covalent bonding, higher
   b. covalent bonding, lower
   c. intermolecular, higher
   d. intermolecular, lower

37. Draw NH_3 and NF_3:
   a. Indicate the direction of the net dipole.
   b. Which of these molecules has polar bonds, but has no net dipole?
   c. What is the shape of each molecule?

38. Lactic acid is a metabolite formed in tired muscles. Using the Lewis structure of lactic acid shown on the right:
   a. Describe the hybridization of each carbon atom.
   b. State the kinds of orbitals on each atom that overlap to form:
      i. carbon–carbon bonds
      ii. carbon–oxygen bonds
      iii. carbon–hydrogen bonds.



39. Identify the intermolecular force that predominates in the following substances:
   a. HCl
   b. CH_3CH_3 (ethane)
   c. CH_3NH_2 (methylamine)
   d. Kr

40. Rank the following substances in order of increasing boiling point (lowest to highest temperature). Consider their structures, IMF, and molar weights.
    a. H_2S (mol. wt. = 34)
    b. CH_3OH (mol. wt. = 32)
    c. C_2H_6 (mol. wt. = 30)
    d. Ne (mol wt. = 20)

Review Questions - Answers

1. $9.5\ \cancel{g} \times \frac{1\ \cancel{kg}}{1000\ \cancel{g}} \times \frac{1\ lb}{0.4536\ \cancel{kg}} = 0.0209\ lb$

   $5.0\ \cancel{years} \times \frac{15\ 000\ \cancel{miles}}{1\ \cancel{year}} \times \frac{1\ km}{0.6214\ \cancel{mi}} = 120\ 695.2044\ km$

   $\frac{0.0209\ lb}{x} = \frac{1\ \cancel{km}}{120\ 695.2044\ \cancel{km}} = 2.5 \times 10^3\ lb\ CO$

2. d) 25 mL of gold will have a greater mass than 25 mL of pyrite.
   Pyrite $= 25\ g \times \frac{1\ mL}{4.5\ g} = 5.55\ mL$ Gold $= 25\ g \times \frac{1\ mL}{19.3\ g} = 1.30\ mL$
   $= 25\ mL \times \frac{4.5\ g}{1\ mL} = 112.5\ g$ $= 25\ mL \times \frac{19.3\ g}{1\ mL} = 482.5\ g$

3. Formula or the name
   a. $Ni_3(PO_4)_2$
   b. $HClO_4$
   c. MgS_2O_3
   d. nitrogen trihydride
   e. calcium phosphide
   f. iron (II) nitrate

4. Reactions:
   a. $NaCl_{(aq)} + AgNO_{3(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$
      already balanced
      Solutions of sodium chloride + Silver nitrate → Solid silver chloride + Sodium nitrate solution
   b. $CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}$
      Carbon tetrahydride (methane) + oxygen gas → carbon dioxide gas and liquid water
   c. $Zn_{(s)} + 1/2O_{2(g)} \rightarrow ZnO_{(s)}$ OR $2Zn_{(s)} + O_{2(g)} \rightarrow 2ZnO_{(s)}$
      Solid zinc + oxygen gas → Solid zinc oxide
   d. $Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$
   e. $2NaOH_{(aq)} + H_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + 2H_2O_{(l)}$

5. 25.94 g x 1 mol / 14.007 g
   = 1.85 / 1.85
   = 1
   74.6 g x 1 mol / 15.999 g
   = 4.66 / 1.85
   = 2.5
   $N_1O_{2.5}$
   $= N_{1 \times 2}O_{2.5 \times 2}$
   $= N_2O_5$

6. Percent yield: $Ca(s) + 2\ H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$
   12.00 g 12.00g (actual)
   MW = 40.078+ 2(15.999 + 1.008) = 74.092 g/mol
   12.00 g x 1 mol/40.078g = 0.2994 mol
   1: 1 mole ratio
   0.2994 mol x 74.092 g/mol = 22.183 g
   Percent yield = (actual yield / theoretical yield) x 100
   = (12.00 g / 22.183 g) x 100
   = 54.10%

7. $3\ Fe(s) + 4\ H_2O(l) \rightarrow Fe_3O_4(s) + 4\ H_2(g)$
   $4.0\text{ g Fe} \times \frac{1\text{ mol Fe}}{55.845\text{ g}} \times \frac{1\text{ mol }Fe_3O_4}{3\text{ mol Fe}} \times \frac{3(55.845)+4(15.999)\text{ g}}{1\text{ mol }Fe_3O_4} = 5.528\text{ g }Fe_3O_4$
   $4.0\text{ g }H_2O \times \frac{1\text{ mol }H_2O}{18.015\text{ g}} \times \frac{1\text{ mol }Fe_3O_4}{4\text{ mol }H_2O} \times \frac{3(55.845)+4(15.999)\text{ g}}{1\text{ mol }Fe_3O_4} = 12.908\text{ g }Fe_3O_4$
   - Fe limiting
   - 5.528 g Fe_3O_4 produced.

   $4.0\text{ g Fe} \times \frac{1\text{ mol Fe}}{55.845\text{ g}} \times \frac{4\text{ mol }H_2O}{3\text{ mol Fe}} \times \frac{18.015\text{ g}}{1\text{ mol }H_2O} = 1.720\text{ g }H_2O\text{ used}$
   4.0 g - 1.720 g H_2O used = 2.28 g left over

8. $2Ag^+(aq) + Cu(s) \rightarrow 2\ Ag(s) + Cu^{2+}(aq)$

9. oxidation number of chromium:
   $K_2Cr_2O_4 = 0$
   $2(+1) + 2x + 4(-2) = 0$
   $2 + 2x - 8 = 0$
   $x = +3$

10. $5\ Fe^{2+}(aq) + MnO_4^-(aq) + 8\ H^+(aq) \rightarrow Mn^{2+}(aq) + 5\ Fe^{3+}(aq) + 4\ H_2O(l)$
    +2 +7 -2 +1 +2 +3 +1 -2
    $x + 4(-2) = -1$ $2(+1) + (-2) = 0$
    $x = +7$
    Fe^{2+} loses e^- and is oxidized; therefore, it's the reducing agent
    Mn gains e^- and is reduced; oxidizing agent = MnO_4^-

11. b. m = 2 → n = 4

$$\frac{1}{\lambda} = R\left(\frac{1}{m^2} - \frac{1}{n^2}\right)$$

1 / 486.1 nm = (1.097 x 10^{-2} nm^{-})($½^2$ - $1/4^2$)
0.002057 = 0.002057

12. True

13. e

14. Orbitals:
    a. spherical, dumbbell, cloverleaf
    b. 1; 3
    c. maximum number of orbitals with n = 3 and l = 2?
        i. n = 3
        ii. l = 0 → n-1, so 0, 1, and 2
        iii. if l = 2, then d orbitals, so total of 5

15. c) The valence electron configuration is identical to carbon.
    4 electrons in 4 p orbitals = valence shell
    4 = n
    l = 0 → n – 1 = 0 (s), 1 (p), 2 (d), 3 (f)
    p = l = 1

16. Neodynium, Nb, Atomic # 60

17. Chromium, Cr

18. d. All of these

19. electron configurations:

(a) In		[Kr] $5s^2$ $4d^{10}$ $5p^1$
	In^{3+}	[Kr] $4d^{10}$ (lost $5p^1$ and $5s^2$ electrons)
(b) Rh		[Kr] $5s^2$ $4d^7$
	Rh^{3+}	[Kr] $4d^6$ (lost the $4d^7$ and $5s^2$ electrons)

20. Increasing ionic radius: $Mg^{2+} < Na^+ < F^- < O^{2-} < N^{3-}$
    - Electron configuration of all these ions = [Ne]
    - As cations, Mg^{2+} and Na^+ are smaller because they both lose their valence shell, and both experience a higher Zeff because of decreased shielding. However, with one less proton, Na^+ experience a lower Zeff relatively.
    - As anions, F^-, O^{2-}, and N^{3-} all experience a decrease Zeff as their proton to electron ratio decrease. So N^{3-} experiences the lowest Zeff and is the largest ion.

21. Fluorine. Ei is the energy needed to remove an electron. Since fluorine is the smallest of these halogens, with only 2 shells, its electrons are strongly attracted to the nucleus.

22. Second ionization energy: c. $[Ne]3s^1 \rightarrow [Ne] + e^-$ Because it has already lost an electron.

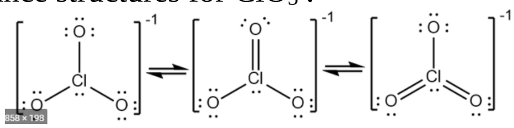
23. False, they both increase as one goes across a period from left to right and both decrease as one does down a group.

24. a.

25. Iodine's valence electrons are in the 5th shell, so the bonds are longer and easily broken.

26. d.

27. Resonance structures for ClO_3^-:

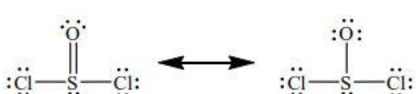


FC = group # - (# bonds) – (# nonbonding e)
Far left structure:
FC O: 6 – 1 – 6 = -1 x 3 = -3
Cl: 7 – 3 – 2 = + 2
Check overall charge: -3 + 2 = -1
Oxygen has the negative charge.

Middle structure:
FC O-: 6 – 1 – 6 = -1 x 2 = -2
O=: 6 -2 -4 = 0
Cl: 7 – 4 – 2 = + 1
Check overall charge: -2 + 1 = -1
Single bonded oxygen has the negative charge.

Far right structure:
FC O-: 6 – 1 – 6 = -1
O=: 6 -2 -4 = 0 x 2 = 0
Cl: 7 – 5 – 2 = 0
Check overall charge: -1 + 0 = -1
Single bonded oxygen has the negative charge.

28. Resonance structures for $SOCl_2$:



Formal charges = Group # - (# covalent bonds) – (# electrons in lone pairs)

0 for Cl, 0 for S, 0 for O	0 for Cl, +1 for S, -1 for O
Most stable because the central atom (S) is the least electronegative and it has no charge.	

29. Three lone pairs are on the Br atom in IBr_2^-

$7 + 2(7) + 1 = 22$ electrons – 4 = 18

[:Br – I – Br:]$^-$

30. $\Delta H° = D$(Reactant bonds) − D(Product bonds)

= {(C=C + C-H + 2(O-H)} – {(C-C + 5(C-H) + C-O + O-H)}

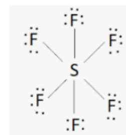
= {(615 kJ + 4(410 kJ) + 2(460 kJ)} – (350 kJ + 5(410kJ) + 350kJ + 460kJ)

= 3175 kJ – 3210 kJ

= -35 kJ

31. I is linear with 2 charge clouds.
II is nonlinear with 4 charge clouds and it is tetrahedral.
With 5 charge clouds, electron geometry of III is trigonal bipyramidal, with a linear molecular geometry.

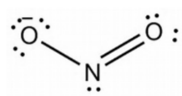
32. Octahedral molecule with smallest bond angle of 90º.

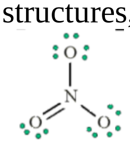


33. sp^3

34. Hybridization on the N atom in NO_2^- and in NO_3^-:

- Lewis structures (they both have resonance structures, but not needed for this question)





- # atoms bonded to nitrogen = 2 = 3
- # of lone pairs = 1 = 0
- 2 + 1 = 3, therefore sp^2 3 + 0 = 3, therefore also sp^2

(NOTE: *if total = 2 = sp*
if total = 4 = sp^3)

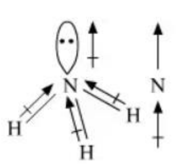
35. Hybridization and molecular shape:

- count # atoms bonded to carbon = 2 :O= C = O:
- Count # of lone pairs on carbon = 0
- 2 + 0 = 2, therefore both are sp and linear H−C≡N:

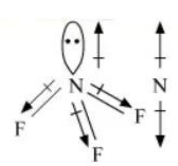
36. d) When melting S_8, <u>intermolecular</u> forces must be overcome and S_8 is expected to have a <u>lower</u> melting point than MgS.

37. NH_3 and NF_3:

a. EN = 3.0 – 2.1 =0.9   4.0 – 3.0 = 1.0



Dipole moment in NH_3



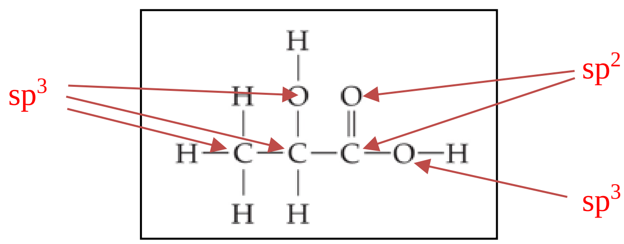
Dipole moment in NF_3

b. Both molecules have polar bonds.
   i. The lone pair of electrons and the dipoles in NH_3 all go in the same direction, thus the molecule is polar.
   ii. The lone pair and the dipoles in NF_3 go in opposite direction, thus it has a lower net dipole and is less polar.

c. Both have a trigonal pyramidal molecular shape.

38.

a. Identify the hybridization of each carbon atom and all 3 oxygen atoms.
b. State the kinds of orbitals for each atom that overlap to form:
   i. carbon–carbon bonds: sp^3 - sp^3 and sp^3 – sp^2
   ii. carbon–oxygen bond (in the chain): sp^2 – sp^3
   iii. carbon–hydrogen bonds: sp^3 - s



39. IMF:
   a. dipole-dipole
   b. LDF
   c. H-bonds
   d. LDF

40. boiling points: d < c < a < b