Chemistry Review Sheet – Final Exam

Chapters One – Three:

1. Review density.
   a. A stone has a mass of 454 grams and occupies a volume of 0.290 liters. Calculate the density of the stone in g/cm^3 (remember, 1.00 mL = 1.00 cm^3).
   b. Explain why ice floats in water but sinks in liquid ethanol.
2. Review the rules for significant figures.
   a. Which has the least number of significant figures?
      0.04000, 1.04, 1.0000001, 0.000006
3. Review accuracy and precision.
   a. The actual number for a measurement is 178. Which of the following sets of numbers is more precise? Which is more accurate?
      177, 200, 157
      199, 200, 202
4. Review the difference between observations and interpretations.
   a. Which of the following is an interpretation? Observation?
      The candle is pink.
      The candle is long.
      When the candle burns, carbon dioxide is released.
5. Review Chemical and Physical Changes
   a. Identify the following as either a chemical or physical change.
      * melting
      * combustion
      * freezing
      * distillation
      * oxidation
      * neutralization
      * condensation
      * sublimation
      * precipitation
      * boiling
      * dissolving
      * decomposition
      * evaporation
      * dissociation
6. Review lab safety procedures.
   a. When should goggles be worn?
   b. How should one dilute an acid (especially sulfuric acid)?
   c. What should one use to neutralize a spilled acid?
7. Review the metric system.
   a. Convert 55 cm to m.
   b. Convert 34 L to mL.
   c. Add 1.050 L and 5 mL (express your answer in liters).

Chapters Eight and Nine:

8. Review nomenclature and formula writing (IUPAC system).
   a. Name K_2SO_4, K_2S, and $(NH_4)_2SO_4$.
   b. How many total atoms are in one formula unit (particle) of $(NH_4)_2SO_4$?
   c. Give the name of the anion and cation in the last compound.
   d. Give the formula of tin(II) chloride and copper(I) oxide.
   e. Give the formula of magnesium oxide.

9. Review chemical composition.
    a. Distinguish between an element, a compound, and a mixture.
10. Review molecular (formula) weight.
    a. What is the formula and molecular weight of carbon tetrachloride?
    b. What is the formula and formula weight of sodium nitrate?
    c. Give the formula and charge of the anion and cation in the last question.

Chapter Ten:

11. Review the different types of reactions (synthesis, decomposition, single replacement, double replacement, combustion, and neutralization (a type of double replacement).
    a. What are the most active (reactive) metals?
    b. What are the products in a complete combustion reaction?
    c. What are the reactants in a neutralization reaction?
12. Review equation balancing.
    a. Supply the missing part: $H_2SO_4 + CaCO_3 \rightarrow CO_2 + H_2O + ?$
    b. Supply the coefficients: $Al_2(SO_4)_3 + KOH \rightarrow Al(OH)_3 + K_2SO_4$

Chapter Eleven:

13. Review mole calculations.
    a. Calculate the number of argon atoms in a 12.0-gram sample of the element.
    b. How many moles of hydrogen atoms are in 10.0 moles of water?
    c. Which has the greater quantity of molecules:
       10.0 grams of water or 10.0 moles of water?
    d. What is the molar mass of nickel(II) nitrate?
    e. How many molecules of water are in 1.25 moles?
    f. How many grams of carbon in a 34.5-gram sample of carbon dioxide?
14. Review mass percentage (percentage composition).
    a. What is the mass percentage of chromium in K_2CrO_4?
15. Review empirical (simplest) formula.
    a. An unknown substance is analyzed and found to contain 18.0 grams of carbon, 24.0 grams of oxygen, and 3.00 grams of hydrogen. What is its empirical formula?
    b. An unknown sample is analyzed and found to contain 50. % oxygen and 50. % sulfur. Determine its simplest formula.
    c. A compound was analyzed and found to have an empirical formula of CH. Its molecular weight was determined to be 78.0 amu. What is the molecular formula of the compound?
16. Review molarity.
    a. A solution contains 12.0 grams of sodium chloride in 2.00 liters of water. What is the molarity?
    b. How many grams of sodium chloride are needed to prepare 3.00 liters of a .200 M solution?

Chapter Twelve:

17. Review how to do mole calculations with balanced equations.
    a. Look at the balanced equation in number twelve. How many grams of aluminum hydroxide could be produced if 15.0 grams of aluminum sulfate reacted?
    b. Calculate how many grams of oxygen gas required to burn 15.0 grams of CH_4?

c. How many moles of barium hydroxide are needed to neutralize (completely react with) 3.00 moles of hydrochloric acid?

18. Review limiting reactants (reagents).
    a. Given: $2CO + O_2 \rightarrow 2CO_2$
    If we have 12.0 g of CO and 12.0 g of O_2, how many grams of CO_2 would be produced?

Chapter Fourteen:

19. Review the mol-volume relationship at STP (22.4 L/mol for gases).
    a. How many liters would 21.0 grams of oxygen gas occupy at STP?
    b. How many moles of gas are in a 11.5-L sample of a gas at a pressure of 760. mmHg and temperature of 0 degrees Celsius (STP)?

Chapter Four:

20. Review atomic number and atomic weight (mass) and how each relates to the concept of isotopes.
    a. How many protons, electrons, and neutrons are there in a single molecule of CO_2?
21. Review atomic structure.
    a. A particle contains 9 protons, 10 neutrons, and 10 electrons. Give the correct formula for this particle.
    b. What is the mass (in amu) of the particle in letter a?
    c. How many electrons are in an Mg^{+2} ion?
22. Review how to calculate atomic weight (average atomic mass).
    a. Boron consists of two naturally occurring isotopes: B-10 with a mass of 10.0129 amu and percent abundance of 19.6% and B-11 with a mass of 11.0093 amu and a percent abundance of 80.4%. Calculate the average atomic mass (atomic weight) of elemental boron.
    b. X-25 makes up 25.0% of a particular element. X-33 makes up the remaining 75.0% of that element. Calculate the atomic weight of element X.
23. Review nuclear symbols.
    a. Give the nuclear symbol for a neutral atom that has 3 protons and 4 neutrons.

Chapter Twenty-Five:

24. Review nuclear equation writing.
    a. Write the nuclear equation for the alpha decay of Rn-222.
    b. Write the nuclear equation for the beta decay of Pb-214.
    c. Write a balanced equation for the neutron bombardment of N-14 in which a proton and an atom of a different element is produced.
25. Review the concept of half-life.
    a. Carbon-14 has a half-life of about 5,700 years. How many grams of C-14 would remain in a 22.0-gram sample after 22,800 years?
26. Review what an isotope is.
27. Review the types of radiation.
    a. What is the most dangerous type of radiation?
28. Review fission and fusion.
    a. Which process produces excess neutrons, which may lead to a chain reaction?
    b. Which process begins with isotopes of large mass?
    c. Which process occurs in the sun?

Chapter Five:

29. Review electron configurations.
    a. Give the electron configuration of titanium.
    b. Where are the *s*, *p*, *d*, and *f* "blocks" located on the periodic table?
30. Review the relationship of electrons and light.
    a. Does an electron give off light when it moves away from the nucleus or towards it?

Chapters Six and Seven:

31. Review these terms: family, group, and period.
    a. Which group of metals is the most reactive?
    b. Which side of the periodic table are most of the elemental gases on?
32. Review the relationship of valence and groups in the periodic table.
    a. Name the group where all the elements form +2 ions.
    b. Based on their positions in the periodic table, write the formula for the compound that would form between Tl and Te.
33. Review the four major trends in the periodic table (along with their definitions).
    a. Which group of elements has the highest ionization energy?
    b. Which group of elements has the highest electronegativity?
    c. Which have larger atomic radii, cations or anions?
    d. Describe where the metals, nonmetals, and metalloids (semi-metals) are relative to the zig-zag line.
34. Who is the inventor of the periodic table?
35. Review the names of the different groups on the periodic table (alkali metals, alkali earth metals, halogens, noble gases, transition metals, lanthanides, and actinides).

Chapter Nine:

36. Review ionic and covalent (molecular) bonding.
    a. What kind of combination or atoms do you need for an ionic bond? For a molecular bond?
37. Review structural formulas.
    a. Give the structural formula of ethane (C_2H_6).
38. Review the relationship between electronegativity and ionic/molecular bonding.
    a. Which type of bond results when there is a large difference in electronegativity between the two atoms in the bond?
    b. Which has greater ionic character: FeS or CaS?
39. Review Lewis structures and the octet rule.
    a. Give the Lewis structure for carbon dioxide (CO_2).
40. Review molecular geometry. What is the geometric shape of:
    a. CH_4
    b. H_2O
    c. NH_3
    d. SO_3
    e. CO_2
41. Review the three types of intermolecular forces (bonding).
    a. What minor type of bonding occurs between water molecules?
    b. Which intermolecular force is the weakest?
    c. When may a hydrogen bond form?

42. Review the concepts of bond and molecule polarity.

a. Is water a polar or nonpolar molecule?
b. Are the bonds in water polar or nonpolar?
c. Which molecules contain the only truly nonpolar bonds?
d. Which of the following molecules are polar?
   CH_4, Cl_2, HF, H_2O, CO_2

Chapter Fifteen:

43. Review writing net-ionic equations (with solubility rules).

a. Write the net-ionic equations for the dissolving of:
   - $BaCl_2$ in water.
   - Na_2SO_4 in water.
b. Write the net-ionic equation for the reaction that occurs when the previous two solutions are mixed.
c. Which anion is always soluble?

Chapter Sixteen:

44. Review calculation of heat changes.

a. Given: $2H_2 + O_2 \rightarrow 2H_2O + 571$ kJ
   Calculate the amount of heat released when 4.00 grams of hydrogen reacts.

Chapter Nineteen:

45. Review the colors of litmus around acids and bases.

a. What color is litmus paper around an acid?

46. Review titrations.

a. How many milliliters of 1.0 M acid are needed to neutralize 32.0 mL of 1.2 M base?

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47. How many moles of hydrogen gas will be produced if 6.50 g of magnesium are placed in 50.0 mL of 1.20 M hydrochloric acid solution?

48. Suppose 5.60 mL of .220 M lead (II) nitrate solution are mixed with 4.75 mL of .319 M potassium iodide solution and allowed to react. How many grams of precipitate should form?

Key - Chemistry Final Review Sheet

1. a. $\frac{454\,g}{290.\,cm^3} = 1.57\,g/cm^3$

   b. Substances will float on substances that are more dense than they are. Ice is more dense than ethanol.

2. a. .000006 has the least number of significant figures with only one.

3. The first set of numbers has an average of 178 which is very accurate when compared to the second set which has an average of 200. The second set, however, represents more precise measurement because all the values are close together.

4. The first is a qualitative observation, the second is a quantitative observation, and the third is an interpretation.

5.

Physical Changes	Chemical Changes
Melting	Combustion
Freezing	Oxidation
Distillation	Neutralization
Condensation	Precipitation
Sublimation	Decomposition
Boiling	
Dissolving	
Evaporation	
Dissociation	

6. a. Goggles should always be worn in the lab.
   b. Add acid to water, not the other way around (especially sulfuric).
   c. Sodium bicarbonate or baking soda (weak base) is used to neutralize spilled acids.

7. a. 55 cm = .55 m
   b. 34 L = 34 000 mL
   c. 1.050 L + .005 L = 1.055 L

8. a. K_2SO_4 - potassium sulfate
   K_2S - potassium sulfide
   $(NH_4)_2SO_4$ - ammonium sulfate

   b. There are 15 atoms in one particle of $(NH_4)_2SO_4$.

   c. anion - sulfate (SO_4^{-2})
   cation - ammonium (NH_4^+)

d. tin (II) chloride - $SnCl_2$
copper (I) oxide - Cu_2O

e. magnesium oxide - MgO

9. An element is a pure substance that consists of only one kind of atom. A compound is a pure substance that consists of two or more kinds of atoms that chemically combined (e.g., bonded). A mixture is two or more substances that are not chemically combined.

10. a. carbon tetrachloride - CCl_4 154 amu

b. sodium nitrate - $NaNO_3$ 85.0 amu

c. anion: NO_3^-
cation: Na^+

11. a. The most active metals are the alkali metals. The next most active metals are the alkali earth metals.

b. The products of complete combustion are CO_2 and H_2O.

c. An acid and a base are the reactants in a neutralization reaction.

12. a. $H_2SO_4 + CaCO_3 \longrightarrow CO_2 + H_2O + \boxed{CaSO_4}$

b. $Al_2(SO_4)_3 + 6\ KOH \longrightarrow 2\ Al(OH)_2 + 3\ K_2SO_4$

13. a. $12.0\text{g Ar}\left(\frac{\text{mol}}{40.0\text{g}}\right)\left(\frac{6.02\times10^{23}\text{ atoms}}{\text{mol}}\right) = 1.81\times10^{23}\text{ atoms Ar}$

b. $10.0\text{ mol } H_2O\left(\frac{2\text{ H}}{1\ H_2O}\right) = 20.0\text{ mol H atoms}$

c. One mole of H_2O has a mass of 18.0g. There are more molecules of H_2O in 10.0 mol H_2O than in 10.0g H_2O:

$$10.0\text{ mol } H_2O\left(\frac{6.02\times10^{23}\ H_2O\text{ molecules}}{\text{mol}}\right) = 6.02\times10^{24}\text{ molecules } H_2O$$

$$10.0\text{g } H_2O\left(\frac{\text{mol}}{18.0\text{g}}\right)\left(\frac{6.02\times10^{23}\ H_2O\text{ molecules}}{\text{mol}}\right) = 3.34\times10^{23}\text{ molecules } H_2O$$

d. nickel (II) nitrate - $Ni(NO_3)_2 \longrightarrow$ 183 g/mol

e. $1.25\text{ mol } H_2O\left(\frac{6.02\times10^{23}\text{ molecules}}{\text{mol}}\right) = 7.52\times10^{23}\text{ molecules}$

f. $34.5\text{g } CO_2\left(\frac{\text{mol}}{44.0\text{g}}\right)\left(\frac{1\text{ C}}{1\ CO_2}\right)\left(\frac{12.0\text{g}}{\text{mol}}\right) = 9.41\text{g C}$

OR

$\frac{12.0\text{ amu}}{44.0\text{ amu}} \times 100 = 27.3\%\text{ C in } CO_2 \longrightarrow 34.5\text{g}(.273) = 9.42\text{g C}$

14. a. $\frac{52.0\text{ amu}}{194\text{ amu}} \times 100 = 26.8\%\ Cr$

15. a. $18.0g\ C \left(\frac{mol}{12.0g}\right) = 1.50\ mol\ C$

$24.0g\ O \left(\frac{mol}{16.0g}\right) = 1.50\ mol\ O$

$3.00g\ H \left(\frac{mol}{1.01g}\right) = 2.97\ mol\ H$

$\frac{2.97\ mol\ H}{1.50\ mol\ C \sim O} = 2H / 1\ C \sim O$

$\therefore\ EF = \boxed{CH_2O}$

b. $50.\%\ O \rightarrow 50.g\ O \left(\frac{mol}{16.0g}\right) = 3.1\ mol\ O$

$50.\%\ S \rightarrow 50.g\ S \left(\frac{mol}{32.1g}\right) = 1.6\ mol\ S$

$\frac{3.1\ mol\ O}{1.6\ mol\ S} = 2\ O / 1\ S$

$\therefore\ EF = \boxed{SO_2}$

c. CH $\rightarrow$ empirical weight = 13.0 amu

$\frac{78\text{ amu}}{13.0\text{ amu}} = 6 \quad \therefore\ 6\ (CH) \longrightarrow MF = \boxed{C_6H_6}$

16. a. $\frac{12.0g\ NaCl \left(\frac{mol}{58.5g}\right)}{2.00L} = .103\ M\ NaCl$

b. $(3.00\ L)(2.00M)\left(\frac{58.5g}{mol}\right) = 35.1g\ NaCl$

17. a. $15.0g\ Al_2(SO_4)_3 \left(\frac{mol}{342g}\right)\left(\frac{2\ Al(OH)_3}{1\ Al_2(SO_4)_3}\right)\left(\frac{78.0g}{mol}\right) = 6.84g\ Al(OH)_3$

b. $CH_4(g) + 2\ O_2(g) \longrightarrow CO_2(g) + 2\ H_2O\ (g)$

$15.0g\ CH_4 \left(\frac{mol}{16.0g}\right)\left(\frac{2\ O_2}{1\ CH_4}\right)\left(\frac{32.0g}{mol}\right) = 60.0g\ O_2$

c. $Ba(OH)_2\ (aq) + 2\ HCl\ (aq) \longrightarrow BaCl_2\ (aq) + 2H_2O\ (l)$

$3.00\ mol\ HCl \left(\frac{1\ Ba(OH)_2}{2\ HCl}\right) = 1.50\ mol\ Ba(OH)_2$

18. a. $2CO + O_2 \longrightarrow 2CO_2$

$12.0g\ CO \left(\frac{mol}{28.0g}\right) = \overset{LR}{.429\ mol\ CO} \left(\frac{2\ CO_2}{2\ CO}\right)\left(\frac{44.0g}{mol}\right) = 18.9g\ CO_2$

$12.0g\ O_2 \left(\frac{mol}{32.0g}\right) = .375\ mol\ O_2 \left(\frac{2\ CO_2}{1\ O_2}\right)\left(\frac{44.0g}{mol}\right) = 33.0g\ CO_2$

19. a. $21.0g\ O_2 \left(\frac{mol}{32.0g}\right)\left(\frac{22.4L}{mol}\right) = 14.7L$

b. $11.5\ L \left(\frac{mol}{22.4\ L}\right) = .510\ mol$

20. a. CO_2 $C \rightarrow 6p^+$ $O \rightarrow 8p^+ \times 2$ 22 protons + 22 electrons

44.0 amu ← molecular wt (CO_2)
$-\ 22$ amu ← wt of protons
22 neutrons

21. a. F^-

b. 19.0 amu

c. 10 electrons

22. a. $(10.0129\ amu)(.196) + (11.0093\ amu)(.804) = 10.81\ amu$

b. $(25\ amu)(.250) + (33\ amu)(.750) = 31\ amu$

23. a. ${}^{7}_{3}Li$

24. a. ${}^{222}_{86}Rn \longrightarrow {}^{4}_{2}He + {}^{218}_{84}Po$

b. ${}^{214}_{82}Pb \longrightarrow {}^{0}_{-1}e + {}^{214}_{83}Bi$

c. ${}^{14}_{7}N + {}^{1}_{0}n \longrightarrow {}^{1}_{1}P + {}^{14}_{6}C$

25. a. $\frac{22{,}800\ yrs}{5700\ yrs/half\text{-}life} = 4$ half-lives $\left(\frac{1}{2}\right)^4 = \frac{1}{16}$ left $\therefore \frac{1}{16}(22.0g) = 1.38g$

26. refer to your notes

27. a. Gamma radiation

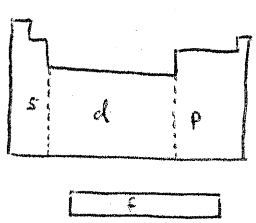
28. a. Fission

b. Fusion

c. Fusion

29. a. Ti: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^2 4s^2$

b.



30. a. Light is given off when electrons fall towards the nucleus.

31. a. The alkali metals are the most active.

b. Room-temperature gases are nonmetals located in the upper-right region of the periodic table (hydrogen is an exception).

32. a. alkali earth metals

b. Tl_2Te_3 (Tl forms a +3 ion; Te forms a −2 ion)

33. a. Noble gases (Ionization energy is the energy required to remove an electron.)

b. Halogens (Electronegativity is the tendency to gain an electron.)

c. anions

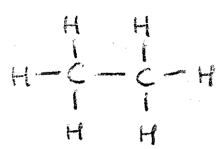
d. Metals are located to the left of the zig-zag (metal line); nonmetals to the right (including H); and semi-metals (metalloids) are located along the zig-zag (excluding Al, which is a metal).

34. Dmitri Mendeleev

35. Refer to your notes

36. a. Ionic bonds form when bonding atoms have an electronegativity difference of 1.7 or greater. They usually form between a metal and a nonmetal. Molecular (covalent) bonds form when electronegativity differences between bonding atoms are less than 1.7 – this usually occurs between nonmetals.

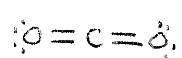
37. a.



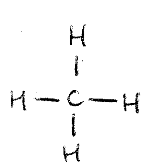
38. a. ionic bonds

b. Based on electronegativity differences alone, CaS would have greater ionic character.

39. a.

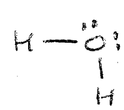


40. a.



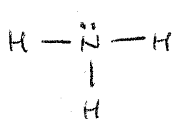
tetrahedral

b.



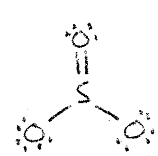
bent, nonlinear

c.



trigonal pyramidal

d.



trigonal planar (planar triangular)

e. $:\ddot{O}=C=\ddot{O}:$ linear

41. a. hydrogen bonding

b. Dispersion (London) Forces are the weakest intermolecular forces.

c. Hydrogen bonds Form when a hydrogen atom bonded to either a nitrogen, oxygen, or Fluorine atom is attracted to another N, O, or F atom on a separate molecule.

42. a. polar

b. polar

c. Molecules where the bonded atoms are identical (e.g., diatomic molecules) or have nearly identical electronegativities exhibit nonpolar bonding.

d. HF and H_2O

43. a. $BaCl_2(s) \longrightarrow Ba^{+2}(aq) + 2Cl^-(aq)$

$Na_2SO_4(s) \longrightarrow 2Na^+(aq) + SO_4^{-2}(aq)$

b. $Ba^{+2}(aq) + SO_4^{-2}(aq) \longrightarrow BaSO_4(s)$

c. nitrate, NO_3^-

44. a. $4.00\text{g } H_2 \left(\frac{\text{mol}}{2.02\text{g}}\right)\left(\frac{571\text{ KJ}}{2\text{ mol } H_2}\right) = 565\text{ KJ released}$

45. a. red

46. a. $(1.0\text{M})\,x = (1.2\text{M})(.0320\text{ L})$

$x = .0400\text{ L} = 40.0\text{ mL}$

47. $Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$

$6.50\text{g Mg}\left(\frac{\text{mol}}{24.3\text{g}}\right) = .267\text{ mol Mg}$

$1.20\text{ M }(.0500\text{ L}) = .0600\text{ mol HCl}$ ← LR

$.0600\text{ mol HCl}\left(\frac{1\ H_2}{2\text{ HCl}}\right) = .0300\text{ mol } H_2$

48. $Pb(NO_3)_2(aq) + 2KI(aq) \longrightarrow PbI_2(s) + 2KNO_3(aq)$

$(.0560\text{ L})(.220\text{ M}) = .0123\text{ mol } Pb(NO_3)_2$

$(.0475\text{ L})(.319\text{ M}) = .0152\text{ mol KI}$

$\frac{.0152\text{ KI}}{.0123\ Pb(NO_3)_2} = 1.2 < 2$

$\therefore$ LR = KI

$.0152\text{ mol KI}\left(\frac{1\ PbI_2}{2\text{ KI}}\right)\left(\frac{461\text{g}}{\text{mol}}\right) = 3.50\text{g } PbI_2$