2017-2018 AP Chemistry Summer Assignment

Complete all work (Yes, you have to show your work!!) on your own paper. I have provided some resources to help you with the topics below but your notes from Honors will be a big help, too. Don't be afraid to search for further resources so you can get a good grasp on the concepts! You are expected to be comfortable with these concepts upon arriving back to school. DO NOT wait until the last minute to complete this assignment!!

Tables to Memorize

Memorize the information in the tables provided on the attached Tables in Zumdahl to Memorize (page 2) paper.

Significant Figures (https://goo.gl/2nQw6h)

1. What is the number of significant figures in each of the following measured quantities? (A) 358 kg, (B) 0.054 s, (C) 6.3050 cm, (D) 0.0105 L, (E) 7.0500×10^{-3} m^3
2. Indicate the number of significant figures in each of the following measured quantities: (A) 3.774 km, (B) 205 m^2, (C) 1.700 cm, (D) 350.00 K, (E) 307.080 g
3. Round each of the following numbers to four significant figures and express the result in scientific notation: (A) 102.53070, (B) 656,980, (C) 0.008543210, (D) 0.000257870, (E) -0.0357202
4. (A) The diameter of Earth at the equator is 7926.381 mi. Round this number to three significant figures, and express it in scientific notation. (B) The circumference of Earth through the poles is 40,008 km. Round this number to four significant figures, and express it in scientific notation.
5. Carry out the following operations and express the answers with the appropriate number of significant figures. (A) 12.0550 + 9.05; (B) 257.2 – 19.789; (C) $(6.21 \times 10^3)(0.1050)$; (D) 0.0577/0.753

Dimensional Analysis (https://goo.gl/6FrrtA)

1. Perform the following conversions: (A) 0.076 L to mL, (B) 5.0×10^{-8} m to nm, (C) 6.88×10^5 ns to s, (D) 0.50 lb to g
2. (A) The speed of light in a vacuum is 2.998×10^8 m/s. Calculate its speed in km/hr. (B) The Vehicle Assembly Building at the Kennedy Space Center in Florida has a volume of 3,666,500 m^3. Convert this volume to liters and express the result in scientific notation. (C) An individual suffering from a high cholesterol level in her blood has 232 mg of cholesterol per 100 mL of blood. If the total blood volume of the individual is 5.2 L, how many grams of total blood cholesterol does the individual's body contain?
3. Perform the following conversions: (A) 5.00 days to s, (B) 0.0550 mi to m, (C) $1.89/gal to dollars per liter, (D) 0.510 in./ms to km/hr, (E) 22.50 gal/min to L/s, (F) 0.02500 ft^3 to cm^3
4. (A) How many liters of wine can be held in a wine barrel whose capacity is 31 gal? (B) If an automobile is able to travel 254 mi on 11.2 gal of gasoline, what is the gas mileage in km/L? (C) A pound of coffee beans yields 50 cups of coffee (4 cups = 1 qt). How many milliliters of coffee can be obtained from 1 g of coffee beans?

5. (A) If an electric car is capable of going 225 km on a single charge, how many charges will it need to travel from Boston, Massachusetts, to Miami, Florida, a distance of 1486 mi, assuming that the trip begins with a full charge? (B) If a migrating loon flies at an average speed of 14 m/s, what is its average speed in mi/hr?

Nomenclature (https://goo.gl/1RbO1f)

1. Locate each of the following elements in the periodic table; indicate whether it is a metal, nonmetal, or metalloid and give the name of the element. (A) Ca, (B) Ti, (C) Ga, (D) Th, (E) Pt, (F) Se, (G) Kr

2. Give the names and charges of the cation and anion in each of the following compounds: (A) CaO, (B) Na_2SO_4, (C) $KClO_4$, (D) $Fe(NO_3)_2$, (E) $Cr(OH)_3$

3. Complete the attached Nomenclature Tables by naming the compound from the chemical formula or writing the formula from the compound name. If the compound is ionic, put a check under the "Ionic?" column.

Balancing, Stoichiometry, & Empirical Formula

1. Balance the following equations:

   (A) $Li(s) + N_2(g) \rightarrow Li_3N(s)$

   (B) $La_2O_3(s) + H_2O(l) \rightarrow La(OH)_3(aq)$

   (C) $NH_4NO_3(s) \rightarrow N_2(g) + O_2(g) + H_2O(g)$

   (D) $Ca_3P_2(s) + H_2O(l) \rightarrow Ca(OH)_2(aq) + PH_3(g)$

   (E) $Ca(OH)_2(aq) + H_3PO_4(aq) \rightarrow Ca_3(PO_4)_2(s) + H_2O(l)$

(https://goo.gl/3tjh7q)

2. Calculate the following quantities: (A) mass, in grams, of 0.105 moles sucrose ($C_{12}H_{22}O_{11}$), (B) moles of $Zn(NO_3)_2$ in 143.50 g of this substance, (C) number of molecules in 1.0×10^{-6} mol CH_3CH_2OH, (D) number of N atoms in 0.410 mol NH_3

3. (A) What is the mass, in grams, of 0.0714 mol of iron(III) sulfate? (B) How many moles of ammonium ions are in 8.776 g of ammonium carbonate? (C) What is the mass, in grams, of 6.52×10^{21} molecules of aspirin, $C_9H_8O_4$, (D) What is the molar mass of diazepam (Valium) if 0.05570 mol weighs 15.86 g?

(https://goo.gl/n1bvRm)

4. Determine the empirical formula of each of the following compounds if a sample contains (A) 0.104 mol K, 0.052 mol C, and 0.156 mol O, (B) 5.28 g Sn and 3.37 g F, (C) 87.5% N and 12.5% H by mass

5. What is the molecular formula of each of the following compounds? (A) empirical formula HCO_2, molar mass = 90.0 g/mol, (B) empirical formula C_2H_4O, molar mass = 88 g/mol

6. The fermentation of glucose ($C_6H_{12}O_6$) produces ethyl alcohol (C_2H_5OH) and CO_2:

$$C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(aq) + 2CO_2(g)$$

(A) How many moles of CO_2 are produced when 0.400 mol of $C_6H_{12}O_6$ reacts in this fashion? (B) How many grams of $C_6H_{12}O_6$ are needed to form 7.50 g of C_2H_5OH? (C) How many grams of CO_2 form when 7.50 g of C_2H_5OH are produced?

7. An iron ore sample contains Fe_2O_3 together with other substances. Reaction of the ore with CO produces iron metal:

$$Fe_2O_3(s) + CO(g) \rightarrow Fe(s) + CO_2(g)$$

(A) Balance this equation. (B) Calculate the number of grams of CO that can react with 0.150 kg of Fe_2O_3. (C) Calculate the number of grams of Fe and the number of grams of CO_2 formed when 0.150 kg of Fe_2O_3 reacts.

Solubility

1. Review and memorize the <u>Solubility Rules</u> on the attached paper.

2. Using solubility guidelines, predict whether each of the following compounds is soluble or insoluble in water: (A) $NiCl_2$, (B) Ag_2S, (C) $CsPO_4$, (D) $SrCO_3$, (E) $PbSO_4$

3. Predict whether each of the following compounds is soluble in water: (A) $Ni(OH)_2$, (B) $PbBr_2$, (C) $Ba(NO_3)_2$, (D) $AlPO_4$, (E) $AgCH_3COO$

Oxidation-Reduction

1. Review and memorize <u>Rules for Assigning Oxidation Numbers to Elements</u> on the attached paper.

2. Define oxidation and reduction in terms of (A) electron transfer and (B) oxidation numbers.

3. Can oxidation occur without accompanying reduction? Explain.

4. Determine the oxidation number for the indicated element in each of the following compounds: (A) Ti in TiO_2, (B) Sn in $SnCl_3^-$, (C) C in $C_2O_4^{2-}$, (D) N in N_2H_4, (E) N in HNO_2, (F) Cr in $Cr_2O_7^{2-}$

5. Which element is oxidized and which is reduced in the following reactions?
   (A) $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
   (B) $3Fe(NO_3)_2(aq) + 2Al(s) \rightarrow 3Fe(s) + 2Al(NO_3)_3(aq)$
   (C) $Cl_2(aq) + 2NaI(aq) \rightarrow I_2(aq) + 2NaCl(aq)$
   (D) $PbS(s) + 4H_2O_2(aq) \rightarrow PbSO_4(s) + 4H_2O(l)$

Molarity

1. (A) Calculate the molarity of a solution made by dissolving 0.750 grams of Na_2SO_4 in enough water to form exactly 850 mL of solution. (B) How many moles of $KMnO_4$ are present in 250 mL of a 0.0475 M solution? (C) How many milliliters of 11.6 M HCl solution are needed to obtain 0.250 mol of HCl?

2. Calculate (A) the number of grams of solute in 0.250 L of 0.175 M KBr, (B) the molar concentration of a solution containing 14.75 g of $Ca(NO_3)_2$ in 1.375 L, (C) the volume of 1.50 M Na_3PO_4 in milliliters that contains 2.50 g of solute.

3. (A) How many grams of solute are present in 50.0 mL of 0.488 M $K_2Cr_2O_7$? (B) If 4.00 g of $(NH_4)_2SO_4$ is dissolved in enough water to form 400 mL of solution, what is the molarity of the solution? (C) How many milliliters of 0.0250 M $CuSO_4$ contain 1.75 g of solute?

4. (A) How many milliliters of a stock solution of 10.0 M HNO_3 would you have to use to prepare 0.450 L of 0.500 M HNO_3? (B) If you dilute 25.0 mL of the stock solution to a final volume of 0.500 L, what will be the concentration of the diluted solution?

5. (A) Starting with solid sucrose, $C_{12}H_{22}O_{11}$, describe how you would prepare 250 mL of a 0.250 M sucrose solution. (B) Describe how you would prepare 350.0 mL of 0.100 M $C_{12}H_{22}O_{11}$ starting with 3.00 L of 1.50 M $C_{12}H_{22}O_{11}$.

Bonding (https://goo.gl/BozuVv)

1. (A) What are valence electrons? (B) How many valence electrons does a nitrogen atom possess? (C) An atom has the electron configuration $1s^22s^22p^63s^23p^2$. How many valence electrons does the atom have?

2. (A) What is the octet rule? (B) How many electrons must a sulfur atom gain to achieve an octet in its valence shell? (C) If an atom has the electron configuration $1s^22s^22p^3$, how many electrons must it gain to achieve an octet?

3. Write the electron configuration for the element titanium, Ti. How many valence electrons does this atom possess?

4. Write the electron configurations for the following ions, and determine which have noble-gas configurations: (A) Zn^{2+}, (B) Te^{2-}, (C) Sc^{3+}, (D) Rh^{3+}

5. (A) Construct a Lewis structure for hydrogen peroxide, H_2O_2, in which each atom achieves an octet of electrons. (B) Do you expect the O-O bond in H_2O_2 to be longer or shorter than the O-O bong in O_2?

6. By referring only to the periodic table, select (A) the most electronegative element in group 6A, (B) the least electronegative element in the group Al, Si, P, (C) the most electronegative element in the group Ga, P, Cl, Na

7. Which of the following bonds are polar: (A) B-F, (B) Cl-Cl, (C) Se-O, (D) H-I? Which is the more electronegative atom in each polar bond?

8. Write the Lewis structures for the following: (A) SiH_4, (B) CO, (C) SF_2, (D) H_2SO_4 (H is bonded to O), (E) ClO_2^-, (F) NH_2OH

TABLES IN ZUMDAHL TO MEMORIZE – page 2

Table 2.5 – Common Polyatomic Ions

Ion	Name
Hg_2^{2+}	Mercury (I)
NH_4^+	Ammonium
NO_2^-	Nitrite
NO_3^-	Nitrate
SO_3^{2-}	Sulfite
SO_4^{2-}	Sulfate
HSO_4^-	Hydrogen sulfate *or* **bisulfate**
OH^-	Hydroxide
CN^-	Cyanide
PO_4^{3-}	Phosphate
HPO_4^{2-}	Hydrogen phosphate
$H_2PO_4^-$	Dihydrogen phosphate

***bold** = commonly used name*

Ion	Name
NCS^-	Thiocyanate
CO_3^{2-}	Carbonate
HCO_3^-	Hydrogen carbonate *or* **bicarbonate**
ClO^-	Hypochlorite
ClO_2^-	Chlorite
ClO_3^-	Chlorate
ClO_4^-	Perchlorate
$C_2H_3O_2^-$	Acetate
MnO_4^-	Permanganate
$Cr_2O_7^{2-}$	Dichromate
CrO_4^{2-}	Chromate
O_2^{2-}	Peroxide
$C_2O_4^{2-}$	Oxalate

Table 2.6 – Prefixes Used to Indicate Number in Chemical Names

Prefix	Number Indicated
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5

Prefix	Number Indicated
hexa-	6
hepta-	7
octa-	8
nona-	9
deca-	10

Table 2.7 – Names of Acids* That Do Not Contain Oxygen

Acid	Name
HF	Hydrofluoric acid
HCl	Hydrochloric acid
HBr	Hydrobromic acid

Acid	Name
HI	Hydroiodic acid
HCN	Hydrocyanic acid
H_2S	Hydrosulfuric acid

**Note that these acids are aqueous solutions containing these substances.*

Table 2.8 – Names of Some Oxygen-Containing Acids

Acid	Name
HNO_3	Nitric acid
HNO_2	Nitrous acid
H_2SO_4	Sulfuric acid

Acid	Name
H_2SO_3	Sulfurous acid
H_3PO_4	Phosphoric acid
$HC_2H_3O_2$	Acetic acid

<u>Nomenclature Tables</u>

Compound Name	Chemical Formula	Ionic?
Ammonium chloride		
Aluminum oxide		
Potassium bromide		
Tin(II) iodide		
Tin(IV) dichromate		
Copper(II) sulfate		
Silver sulfide		
Cesium iodide		
Magnesium hydrogen carbonate		
Titanium(IV) oxide		
Nitrogen dioxide		
Sulfur trioxide		
Potassium hypochlorate		
Hydrogen sulfide		
Copper(I) oxide		
Barium hydroxide		
Radium nitrate		
Ammonium chromate		
Carbon tetrachloride		
Lithium cyanide		

Chemical Formula	Compound Name	Ionic?
NH_4Cl		
Al_2O_3		
KBr		
K_2S		
SnI_2		
Cu_2O		
Ag_2S		
CsI		
$Ni(NO_3)_2$		
Cl_2O		
$AlPO_4$		
$MnSO_3$		
$(NH_4)_2HPO_4$		
$ZnCr_2O_7$		
$(NH_4)HSO_4$		
IF_7		
Br_2O_3		
$Ba(OH)_2$		
$Mg(NO_2)_2$		
$Al(C_2H_3O_2)_3$		

Solubility Rules

1. The following are the solubility rules for common ionic solids. If there two rules appear to contradict each other, the preceding rule takes precedence.
2. Salts containing Group I elements (Li^+, Na^+, K^+, Cs^+, Rb^+) are soluble . There are few exceptions to this rule. Salts containing the ammonium ion (NH_4^+) are also soluble.
3. Salts containing nitrate ion (NO_3^-) are generally soluble.
4. Salts containing Cl^-, Br^-, or I^- are generally soluble. Important exceptions to this rule are halide salts of Ag^+, Pb^{2+}, and $(Hg_2)^{2+}$. Thus, $AgCl$, $PbBr_2$, and Hg_2Cl_2 are insoluble.
5. Most silver salts are insoluble. $AgNO_3$ and $Ag(C_2H_3O_2)$ are common soluble salts of silver; virtually all others are insoluble.
6. Most sulfate salts are soluble. Important exceptions to this rule include $CaSO_4$, $BaSO_4$, $PbSO_4$, Ag_2SO_4 and $SrSO_4$.
7. Most hydroxide salts are only slightly soluble. Hydroxide salts of Group I elements are soluble. Hydroxide salts of Group II elements (Ca, Sr, and Ba) are slightly soluble. Hydroxide salts of transition metals and Al^{3+} are insoluble. Thus, $Fe(OH)_3$, $Al(OH)_3$, $Co(OH)_2$ are not soluble.
8. Most sulfides of transition metals are highly insoluble, including CdS, FeS, ZnS, and Ag_2S. Arsenic, antimony, bismuth, and lead sulfides are also insoluble.
9. Carbonates are frequently insoluble. Group II carbonates ($CaCO_3$, $SrCO_3$, and $BaCO_3$) are insoluble, as are $FeCO_3$ and $PbCO_3$.
10. Chromates are frequently insoluble. Examples include $PbCrO_4$ and $BaCrO_4$.
11. Phosphates such as $Ca_3(PO_4)_2$ and Ag_3PO_4 are frequently insoluble.
12. Fluorides such as BaF_2, MgF_2, and PbF_2 are frequently insoluble.

In summary…

Soluble

Always Soluble	*Soluble with Exceptions*
All compounds of the alkali metals: Li^+, Na^+, K^+, Rb^+, Cs^+	All compounds of Cl^-, Br^-, I^- except with: Ag^+, Pb^{2+}, Hg_2^{2+}
All compounds containing: NH_4^+, NO_3^-, ClO_4^-, ClO_3^-, $CH_3CO_2^-$	All compounds of SO_4^{2-} except with: Pb^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, Hg_2^{2+}

Insoluble

Insoluble with Exceptions

All compounds of PO_4^{3-}, CO_3^{2-}, S^{2-}, SO_3^{2-}
except with: alkali metals or NH_4^+

All compounds of OH^- or O^{2-}
except with: alkali metals, NH_4^+, Ca^{2+}, Sr^{2+}, or Ba^{2+}

Rules for Assigning Oxidation Numbers to Elements

Oxidation numbers are bookkeeping numbers. They allow chemists to do things such as balance redox (*red*uction/*ox*idation) equations. Oxidation numbers are positive or negative numbers, but don't confuse them with positive or negative charges on ions or valences.

Oxidation numbers are assigned to elements using these rules:

- **Rule 1:** The oxidation number of an element in its free (uncombined) state is zero — for example, Al(s) or Zn(s). This is also true for elements found in nature as *diatomic* elements H_2, O_2, N_2, F_2, Cl_2, Br_2, or I_2 and for sulfur, found as: S_8
- **Rule 2:** The oxidation number of a *monatomic* (one-atom) ion is the same as the charge on the ion, for example: $Na^+ = +1$ $S^{2-} = -2$
- **Rule 3:** The sum of all oxidation numbers in a neutral compound is zero. The sum of all oxidation numbers in a *polyatomic* (many-atom) ion is equal to the charge on the ion. This rule often allows chemists to calculate the oxidation number of an atom that may have multiple oxidation states, if the other atoms in the ion have known oxidation numbers.
- **Rule 4:** The oxidation number of an alkali metal (IA family) in a compound is +1; the oxidation number of an alkaline earth metal (IIA family) in a compound is +2.
- **Rule 5:** The oxidation number of oxygen in a compound is usually –2. If, however, the oxygen is in a class of compounds called *peroxides* (for example, hydrogen peroxide), then the oxygen has an oxidation number of –1. If the oxygen is bonded to fluorine, the number is +1.
- **Rule 6:** The oxidation state of hydrogen in a compound is usually +1. If the hydrogen is part of a *binary metal hydride* (compound of hydrogen and some metal), then the oxidation state of hydrogen is –1.
- **Rule 7:** The oxidation number of fluorine is always –1. Chlorine, bromine, and iodine usually have an oxidation number of –1, unless they're in combination with an oxygen or fluorine.

These rules give you another way to define oxidation and reduction — in terms of oxidation numbers. For example, consider this reaction, which shows oxidation by the loss of electrons: $Zn(s) \rightarrow Zn^{2+} + 2e^-$

Notice that the zinc metal (the reactant) has an oxidation number of zero (rule 1), and the zinc cation (the product) has an oxidation number of +2 (rule 2). In general, you can say that a substance is oxidized when there's an increase in its oxidation number.

Reduction works the same way. Consider this reaction: $Cu^{2+} + 2e^- \rightarrow Cu(s)$

The copper is going from an oxidation number of +2 to zero. A substance is reduced if there's a decrease in its oxidation number.