Welcome to AP Chemistry!

According to the College Board, "*The AP Chemistry course is designed to be the equivalent of the general chemistry course usually taken during the first college year. The college course in general chemistry differs qualitatively from the usual first secondary school course in chemistry with respect to the kind of textbook used, the topics covered, the emphasis on chemical calculations and the mathematical formulation of principles, and the kind of laboratory work done by students. Quantitative differences appear in the number of topics treated, the time spent on the course by students, and the nature and the variety of experiments done in the laboratory. Students in an AP Chemistry course should spend at least five hours a week in individual study outside of the classroom.*"

I am excited that you have accepted the challenge that an AP Chemistry course has to offer. To ensure that all students in the AP Chemistry class are ready to partake in this high-pace, rigorous journey on the first day of school, the following summer assignment must be completed. The purposes of the assignment are to revisit chemical concepts learned in your 1st year chemistry class and expose you to the level of rigor demanded by the AP curriculum. This will allow us to focus our attention on the advanced chemistry topics that will be tested on the AP exam on May 7, 2020 Your summer assignment consists of the following:

1. Read Chapters 1 – 4 of your text and complete the Practice Questions (Parts I-IV) on the following pages. These problems, worked out in their entirety, are **due the first full day of classes**.
2. Memorize the polyatomic ions and solubility rules listed.

Please take the assignment seriously and start in early August or sooner — there's a lot to do and you won't be able to complete it all on the night before!

If at any time you would like to ask me a question, please email me at boconnell@achs.net.

I look forward to a great year,

Ms. O'Connell

* eBook = Chemistry: The Central Science, 14e – Pearson

authors = Theodore Brown, H. Eugene LeMay

Read Chapters 1 – 4 and answer the following practice questions. These problems, worked out in their entirety, are **due the first full day of classes**. All work must be shown for calculations in order to receive credit for the problem.

I. Nomenclature

Name each of the following compounds and state whether they are ionic or covalent:

a. CuI _____

b. CuI_2 _____

c. CoI_2 _____

d. $NaHCO_3$ _____

e. S_4N_4 _____

f. SF_6 _____

g. NaOCl _____

h. $BaCrO_4$ _____

Write formulas for each of the following compounds and state whether they are ionic or covalent:

a. potassium cyanide _____

b. copper (II) nitrate _____

c. selenium tetrabromide _____

d. iodous acid _____

e. lead (IV) sulfide _____

f. copper (I) chloride _____

g. gallium arsenide _____

h. cadmium selenide _____

Each of the following compounds is incorrectly named. What is wrong with each name and what is the correct name for each compound?

a. $FeCl_3$, iron chloride _____

b. NO_2, nitrogen (IV) oxide _____

c. CaO, calcium (II) monoxide _____

d. Al_2S_3, dialuminum trisulfide _____

e. $Mg(C_2H_3O_2)_2$, manganese diacetate _____

f. $FePO_4$, iron (II) phosphide _____

g. P_2S_5, phosphorous sulfide _____

h. Na_2O_2, sodium oxide _____

i. HNO_3, nitrate acid _____

j. H_2S, sulfuric acid _____

II. Chemical Equations

For each equation below:

- identify the type (synthesis, decomposition, single replacement, metathesis/double replacement, or combustion)
- predict the products, and then
- write the balanced equation and net ionic equation.
- Remember to use the solubility rules for double replacement reactions and the activity series for single replacement reactions. Reminder: when writing these reactions, ignore all of the information about excess, or bubbling, or mixing. These are just excess words used to make complete sentences. Simply pull out the chemical formulas.
- For example:
  Solutions of silver nitrate and magnesium iodide are combined.
  This is a double replacement reaction.
  $2AgNO_3(aq) + MgI_2(aq) \rightarrow 2AgI(s) + Mg(NO_3)_2(aq)$
  $2Ag^+(aq) + 2I^-(aq) \rightarrow 2AgI(s)$

1. Ammonium sulfate reacts with barium nitrate.

2. Zinc metal is added to a solution of copper (II) chloride.

3. Propane gas (C_3H_8) is burned in excess oxygen.

4. Solid calcium chlorate is heated strongly.

5. Magnesium and nitrogen gas are heated together.

6. Chlorine gas is bubbled through a solution of sodium bromide.

7. Sodium metal is added to distilled water.

8. Sulfuric acid is combined with sodium hydroxide.

9. Solid sodium carbonate is heated in a crucible.

III. Chemical Quantities (show all calculations)

1. How many **significant figures** are in each of the following?
   a. 1.9200 mm
   b. 0.0301001 kJ
   c. 6.022×10^{23} atoms
   d. 460.000 L
   e. 0.000036 cm^3
   f. 10000
   g. 1001
   h. 0.001345
   i. 0.0101
   j. 3.21×10^{-2}

2. Record the following in correct **scientific notation**:
   a. 4050,000,000 cal
   b. 0.000123 mol
   c. 0.00345 Å
   d. 700,000,000 atoms

3. Calculate the following to the **correct number** of significant figures.
   a. 1.270 g / 5.296 cm^3
   b. 12.235 g / 1.010 L
   c. 12 g + 0.38 g
   d. 170g + 2.785 g
   e. 2.1 x 3.2102
   f. 200.1 x 120
   g. 17.6 + 2.838 + 2.3 + 200

4. Assume silicon has three major isotopes in nature as shown in the table below. Calculate and fill in the missing information. Show all work.

Isotope	Mass (amu)	Abundance
Si-28	27.98	
Si-29		4.70%
Si-32	29.97	3.09%

5. Calcium carbonate decomposes upon heating. How many grams of calcium oxide will be produced after 12.25 g of calcium carbonate is completely decomposed?

6. When ammonia gas, oxygen gas and methane gas (CH_4) are combined, the products are hydrogen cyanide gas and water. Calculate the mass of each product produced when 2.25×10^2 g of oxygen gas is reacted with an excess of the other two reactants. If the actual yield of the experiment is 105 g of HCN, calculate the percent yield.

7. One type of electromagnetic radiation has a frequency of 107.1 MHz, another type has a wavelength of 2.12×10^{-10} m, and another type has photons with energy equal to 3.97×10^{-19} J/photon. Identify each type of electromagnetic radiation and place them in order of increasing photon energy and increasing frequency.

8. Determine the empirical formula of the compounds with the following compositions by mass:
   a. 10. 4 % C, 27. 8% S , 61. 7 % Cl

   b.21.7 % C, 9.6 % O, and 68.7 % F

9. Calculate the percentage by mass of the following compounds:

   a. SO_3

   b. CH_3COOCH_3

   c. Ammonium Nitrate.

IV. Atomic Structure, Periodicity, and Bonding Review

1. Answer the following questions based on the given electron configuration and identify the elements.

a. Arrange these atoms in order of increasing size: $[Kr]5s^2 4d^{10} 5p^5$; $[Kr]5s^2 4d^{10} 5p^1$; $[Kr]5s^2 4d^{10} 5p^3$.

b. Arrange these atoms in order of decreasing first ionization energy: $[Ne]3s^2 3p^5$; $[Ar]4s^2 3d^{10} 4p^3$; $[Ar]4s^2 3d^{10} 4p^5$.

2. Write the expected ground-state electron configuration for the following:

a. the element with one unpaired 5p electron that forms a covalent with compound fluorine

b. the first-row transition metal with the most unpaired electrons

b. the metalloid in period 3

c. the halogen in period 5

d. the element with atomic number 47

e. the sodium ion

Periodic Trends

1. Discuss the importance of Mendeleev's periodic law.

2. Identify each element as a metal, metalloid, or nonmetal.
   a) fluorine _____
   b) germanium _____
   c) zinc _____
   d) phosphorous _____
   e) lithium _____

3. Give two examples of elements for each category.
   a) noble gases _____
   b) halogens _____
   c) alkali metals _____
   d) alkaline earth metals _____

4. What trend in atomic radius do you see as you go down a group/family on the periodic table? What causes this trend?

5. What trend in atomic radius do you see as you go across a period/row on the periodic table? What causes this trend?

6. Circle the atom in each pair that has the largest atomic radius.

a)	Al	B	b)	S	O	c)	Br	Cl
d)	Na	Al	e)	O	F	f)	Mg	Ca

7. Define ionization energy.

8. Is it easier to form a positive ion with an element that has a high ionization energy or an element that has a low ionization energy? Explain.

9. Use the concept of ionization energy to explain why sodium form a 1+ ion (Na^+) but magnesium forms a 2+ ion (Mg^{2+}).

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V

Limiting Reactants

1

6. Aluminum reacts with bromine to produce aluminum bromide according to the following equation:

$$\_\_Al\ (s) + \_\_Br_2\ (l) \rightarrow \_\_AlBr_3\ (s) \qquad \text{[unbalanced]}$$

In the lab you perform this reaction using an excess of aluminum and 55.7 grams of bromine. You end up collecting 60.0 grams of product aluminum bromide. What is your percent yield?

2

7. Iron reacts with oxygen to give iron(III) oxide according to the following balanced equation:

$$4\ Fe\ (s) + 3\ O_2\ (g) \rightarrow 2\ Fe_2O_3\ (s)$$

Identify the limiting reactant in each of the following reactant mixtures:

a. 0.288 moles Fe and 0.240 moles O_2
b. 10.0 grams Fe and 10.0 grams O_2
c. 1.45×10^{24} atoms of Fe and 7.21×10^{23} molecules of O_2

3

8. Sucrose ($C_{12}H_{22}O_{11}$) reacts with potassium chlorate according to the following balanced equation:

$$C_{12}H_{22}O_{11}\ (s) + 8\ KClO_3\ (s) \rightarrow 12\ CO_2\ (g) + 11\ H_2O\ (g) + 8\ KCl\ (s)$$

Suppose 3.5 moles of $C_{12}H_{22}O_{11}$ are mixed with 21.0 moles of $KClO_3$. Identify the limiting reactant, then calculate how many moles of each product (CO_2, H_2O and KCl) will be formed in this reaction.

4

9. Boron trifluoride reacts with hydrogen according to the following balanced equation:

$$2\ BF_3\ (g) + 3\ H_2\ (g) \rightarrow 2\ B\ (s) + 6\ HF\ (g)$$

Suppose 40.00 grams of BF_3 are mixed with 5.00 grams of H_2.

a. Identify the limiting reactant.
b. Calculate the mass of HF generated.
c. Calculate the mass of excess reactant that remains when the reaction is complete.
d. When this reaction is performed in the laboratory, the percent yield of HF is 72.6%. What was the experimental yield of HF?

5

10. Pure carbon reacts with sulfur dioxide gas producing liquid carbon disulfide and carbon monoxide gas.

a. Write the balanced equation for this reaction.
b. What mass of carbon monoxide will be produced if 40.0 grams of carbon are allowed to react with 95.0 grams of sulfur dioxide? What mass of excess reactant is left over?

Gas laws

Boyles Law (6.3)

$P_1V_1 = P_2V_2$

Charles Law (6.4)

$\frac{V1}{T1} = \frac{V2}{T2}$

Guy-Lussac's Law (6.5)

$\frac{P_1}{T_1} = \frac{P_2}{T_2}$

Combined Gas Law (6.6)

$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

Avogadros Law and STP (6.7) **Standard T = 0 °C & Standard P =1 atm**

$\frac{V_1}{n_1} = \frac{V_2}{n_2}$ **Molar volume: 1 mole = 22.4 L at STP**

$\frac{22.4\ L}{1\ mol} = \frac{V_2}{n_2}$

Dalton Law of Partial Pressure (6.8)

$P_1 + P_2 + \ldots . P_n = P_{total}$

Notes:

1. Temperatures must be in K, where K = C + 273

2. Pressures and volumes must be in consistent units.

VI

Gas Law Problems.

Boyle's Law (6.3)

1. A gas has a volume of 300 mL at 300 mm Hg. What will its volume be if the pressure is changed to 500 mm Hg? (180 mL)

2. A gas has a volume of 460 mL at 500 mm Hg. What will be the volume at 1.2 atm? (252.2 mL)

3. A gas has a volume of 5 liters at 3 atm. To expand the volume to 7500 ml, what the new pressure (in atm) have to be? (2 atm)

Charles' Law (6.4)

4. A gas has a volume of 4 liters at 50 °C. What will its volume be (in liters) at 100°C? (4.6 L)

5. A gas has a volume of 350 ml at 45°C. If the volume changes to 400 ml, what is the new temperature? (answer in °C) (90.4 °C)

Guy-Lussac's Law (6.5)

6. The gases in a hair spray can are at a temperature of 27°C and a pressure of 30 lbs/in^2. If the gases in the can reach a pressure of 90 lbs/in^2, the can will explode. To what temperature must the gases be raised in order for the can to explode? Assume constant volume. (630 °C)

7. Maybelline Cousteau's backup oxygen tank reads 900 mmHg while on her boat, where the temperature is 27°C. When she dives down to the bottom of an unexplored methane lake on a recently-discovered moon of Neptune, the temperature will drop down to –183°C. What will the pressure in her backup tank be at that temperature? (270 mmHg)

Combined Gas Law (6.6)

8. A gas has a volume of 39 liters at STP. What will its volume be at 4 atm and 25°C? (10.6 L)

9. 400 ml of a gas is contained at 300 mm Hg and 0°C. What will its volume be in mL at 140 mm Hg and 10°C? (888.5 mL)

10. 500 ml of gas is contained at STP. The volume changes to 560 ml at 20°C and what pressure (in atm)? (0.96 atm)

Avogadros Law and Molar Volume at STP (6.7) (1 mole of any gas = 22.4 L at STP)

11. 50 g of nitrogen (N_2) has a volume of ___ liters at STP. (40 L)

12. 100 g of oxygen(O_2) is added to the gas in Question 16. What is the volume of the combined gases at STP. (110 L)

13. What is the density of carbon dioxide at STP? (2.0 g/L)

Dalton's Law Worksheet (6.8)

14) A metal tank contains three gases: oxygen, helium, and nitrogen. If the partial pressures of the three gases in the tank are 35 atm of O_2, 5 atm of N_2, and 25 atm of He, what is the total pressure inside of the tank? (**atm)**

15) Blast furnaces give off many unpleasant and unhealthy gases. If the total air pressure is 0.99 atm, the partial pressure of carbon dioxide is 0.05 atm, and the partial pressure of hydrogen sulfide is 0.02 atm, what is the partial pressure of the remaining air? (

16) If the air from problem 15 contains 22% oxygen, what is the partial pressure of oxygen near a blast furnace? **atm)**

SOLUBILITY RULES

1. Salts of ammonium (NH_4^+) and Group IA are always soluble.

2. a. All chlorides (Cl^-) are soluble except AgCl, Hg_2Cl_2, and $PbCl_2$ which are insoluble.

   b. All bromides (Br^-) are soluble except AgBr, Hg_2Br_2, $HgBr_2$, and $PbBr_2$ which are insoluble.

   c. All iodides (I^-) are soluble except AgI, Hg_2I_2, HgI_2, and PbI_2 which are insoluble.

3. Chlorates (ClO_3^-), nitrates (NO_3^-), and acetates (CH_3COO^-) are soluble.

4. Sulfates (SO_4^{-2}) are soluble except $CaSO_4$, $SrSO_4$, $BaSO_4$, Hg_2SO_4, $HgSO_4$, $PbSO_4$, and Ag_2SO_4 which are insoluble.

5. Phosphates (PO_4^{-3}), and carbonates (CO_3^{-2}) are insoluble except NH_4^+ and Group IA compounds.

6. All metallic oxides (O^{-2}) are insoluble except NH_4^+ and Group IA compounds.

7. All metallic hydroxides (OH^-) are insoluble except NH_4^+ and Group IA and Group IIA from calcium down.

8. All sulfides (S^{-2}) are insoluble except NH_4^+ and Groups IA and IIA.

Polyatomic Ion Names

1+
ammonium, NH_4^+
hydronium, H_3O^+

2+
mercury (I), Hg_2^{2+}

1-
acetate, $C_2H_3O_2^-$, or CH_3COO^-
bromate, BrO_3^-
perchlorate, ClO_4^-
chlorate, ClO_3^-
chlorite, ClO_2^-
hypochlorite, ClO^-
cyanide, CN^-
hydrogen carbonate, HCO_3^- (also called bicarbonate)
hydrogen sulfate, HSO_4^-
hydroxide, OH^-
iodate, IO_3^-
nitrate, NO_3^-
nitrite, NO_2^-
permanganate, MnO_4^-
thiocyanate, SCN^-

2-
carbonate, CO_3^{2-}
chromate, CrO_4^{2-}
dichromate, $Cr_2O_7^{2-}$
oxalate, $C_2O_4^{2-}$
peroxide, O_2^{2-}
sulfate, SO_4^{2-}
sulfite, SO_3^{2-}
thiosulfate, $S_2O_3^{2-}$

3-
phosphate, PO_4^{3-}
phosphite, PO_3^{3-}
arsenate, AsO_4^{3-}