HOLY CHILD

EDUCATING YOUNG WOMEN OF CONSCIENCE AND ACTION

AP Chemistry Summer Homework

Welcome to AP Chemistry - one of the most challenging courses you will take in high school! To ensure that you come fully prepared and ready for the rigor of AP Chem you will need to complete the following summer assignment. Getting familiar with the AP Chemistry textbook and style of questioning is going to be a key to your success. This assignment is meant to be a little review as well as some new material. You will be quizzed on this information within the first week of school and you will earn points for completing this packet/setting up your notes. Pace yourself and DO NOT wait until the last week of break to complete... it will take you longer than you think!

<u>On-line Resources</u>: Bozeman Science and Tyler DeWitt videos (on YouTube) are very helpful for concepts you may not understand or need a quick refresher. The Bozeman videos closely follow the textbook so they will be a very good resource for you. The Tyler DeWitt videos cover basic chemistry skills and concepts and he explains things in a very understandable manner with good visuals.

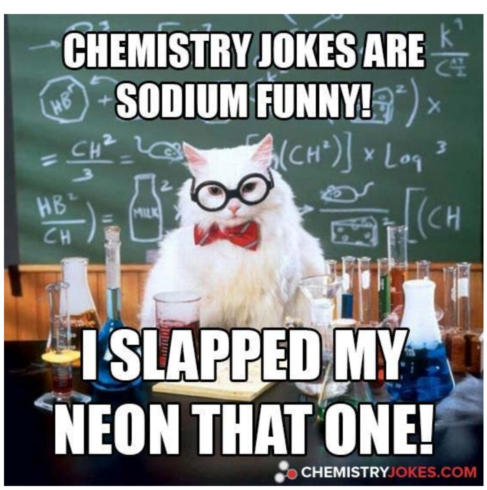
<u>Materials Needed</u>:

- 3-ring binder OR notebook (preferably a Mead 5-Star)
- Calculator (TI is preferred over Casio) - must have scientific notation capability (look for EE key)
- Periodic Table

<u>Structure of Homework</u>:

A. Concepts to take notes on *(do this on loose-leaf for your binder or your notebook -- develop a good note-taking and organization system)*
B. Things to memorize *(make flashcards!)*
C. Practice problems *(do in this packet)*

****If you cannot locate something in your textbook, use on-line resources!**



My email: j.tracey@holychildrye.org

Definitely make sure you do the summer homework

Don't procrastinate

Study so much. Do NOT underestimate this course. It is extremely difficult. Do not base your performance in regular Chemistry on AP Chem.

Read and reread notes until you are sure you completely understand it

Have fun and enjoy the time

Three things: Study, Study, Study

Be prepared to do a lot of work

Make sure you love chemistry and aren't afraid to do work outside of the classroom

If you focus on homework and make an effort to not only do what is asked, but understand what is being taught, the course will go very well.

Chapter 1: Chemical Foundations

1.1 Chemistry: An Overview – Read this short section as an introduction.

1.2 The Scientific Method

A. Take notes:
   - Quantitative and qualitative observations/data (examples are helpful)
   - Theory/model, natural law, compare and contrast these two

B. Memorize/Flashcards - n/a

C. Practice Problems - n/a

1.3 Units of Measurement

A. Take notes:
   - 2 required parts of a measurement
   - SI System - history, what is it
   - Fundamental SI Units (leave out luminous intensity)
   - Define volume, mass, and weight

B. Memorize/Flashcards - Table 1.2 -- Prefix, symbol, meaning, exponential notation of those in BLUE

C. Practice Problems

1. Mary's height is measured and recorded at 1.65. What is wrong with this measurement? Is it a quantitative or qualitative observation?

2. If you traveled to the moon...of volume, mass, and weight - which of these values would change? Which wouldn't? Explain.

1.4 Uncertainty in Measurement:

A. Take notes:

- Define: accuracy, precision, random error, systematic error (use examples and/or diagrams to help your understanding)

B. Memorize/Flashcards - n/a

C. Practice Problems

3. A student performed an analysis of a sample of limestone for its calcium content and got the following results: Trial 1: 14.92% Ca; Trial 2: 14.91% Ca; Trial 3: 14.88% Ca; Trial 4: 14.91% Ca. The actual amount of calcium in the sample is 15.70%.

a. Find the average experimental value and the range of values.

Average _______________ Range _____________________________

b. Would you say this data is accurate? Precise? Explain.

c. What type of error seems to have occurred? Explain how you know.

1.5 Significant Figures and Calculations:

A. Take notes:

- Rules for counting sig figs (shorten these up!)
- Rules for sig figs in mathematical operations
- Rules for rounding

B. Memorize/Flashcards - n/a

C. Practice Problems

4. Indicate the number of sig figs in each of the following:

a. A mile is about 5300 ft. ____________ b. 463.8052 ___________

c. A liter is equivalent to 1.059 qt. ______ d. 0.003840 ___________

e. 17.00 ________ f. 4.0050×10^6 ________

5. Perform the following mathematical operations and express each result to the correct number of sig figs:

a. 52.331 + 26.01 + 0.9981 = _______________

b. 7.25 – 6.8350 = ________________

c. (6.404)(2.91) = ________________

d. (6.6262)(2.998)/2.54 = _____________

e. $(6.02 \times 10^{23})(24.3)/184.5$ = ____________

<u>1.6 Learning to Solve Problems Systematically</u> – Read this short section.

<u>1.7 Dimensional Analysis</u>:

A. Take notes:

- Study examples 1.5-1.10. Put examples in notes as needed

B. Memorize/Flashcards - n/a

C. Practice Problems **(use Table 1.4 on pg. 18 for English-Metric equivalents)**

6. The circumference of the earth is 25,000 miles at the equator. What is the circumference in kilometers?

7. The typical speed limit on the Interstate system is 70 miles per hour. A. Convert this to kilometers per hour. B. Convert it to meters per second.

8. Congratulations! You and your spouse are the proud parents of a new baby, born while you are vacationing in Spain, a country that uses the metric system. The nurse has informed you that the baby weights 3.91 kg and is 51.4 cm long. Convert your baby's weight to pounds AND her length to inches.

1.8 Temperature

A. Take notes:

- 3 temperature scales – uses of each, boiling point, freezing point of each (Figure 1.9 would be a helpful diagram to put in your notes)
- Formulas for converting between temperature scales

B. Memorize/Flashcards – How to convert between Celsius and Kelvin

C. Practice Problems

9. Convert to Kelvin: a. 39.2°C ___________ b. -25°C ______________

10. Convert to Celsius:

a. -459°F ___________ b. 68°F ______________

11. Convert to Fahrenheit:

a. 801°C ____________ b. 100°C ____________

1.9 Density

A. Take notes:

- Definition of density, formula to calculate density, units of density, uses of density

B. Memorize/Flashcards - n/a

C. Practice Problems

12. A rectangular block has dimensions 2.9 cm x 3.5 cm x 10.0 cm. The mass of the block is 615.0 grams. What is the density of the block?

13. A material will float on the surface of a liquid if the material has a density less than that of the liquid. Given that the density of water is approximately 1.0 g/mL, will a block of material having a volume of 125 cm^3 and having a mass of 1.5 pounds sink or float when placed in the water?

14. Diamonds are measured in carats; 1 carat = 0.200g. The density of diamond is 3.51 g/cm^3. What is the volume of a 5.0 carat diamond?

1.10 Classification of Matter

A. Take notes:

- 3 states of matter – shape, volume – definite or indefinite
- Make a flow chart to define, give examples, and organize: matter, mixture, homogeneous mixure, heterogeneous mixture, solution, pure substance, element, compound
- Physical change – define, give examples
- Ways to separate mixtures
- Chemical change – define, give examples

B. Memorize/Flashcards - n/a

C. Practice Problems

15. Classify as a mixture or a pure substance:

a. water ____________________
b. blood ____________________
c. uranium __________________

d. brass ____________________

e. iron ____________________

f. Of the pure substances, which are elements? ______________________

e. What would be an easy way to decide if something is an element?

16. Suppose a teaspoon of magnesium filings and a teaspoon of powdered sulfur are placed together in a beaker. Would this be considered a mixture or a pure substance?

Now suppose the beaker is heated so they react with each other forming magnesium sulfide. Would this be considered a mixture of a pure substance?

17. During a very cold winter, the temperature may remain below freezing for extended periods. However, fallen snow can still disappear, even though it is too cold to melt. This is possible because a solid can vaporize directly, without passing through the liquid state. Is this process (called sublimation) a physical or a chemical change?

18. Which of the following describes a chemical property?

a. The density of iron is 7.87 g/cm^3
b. A platinum wire glows red when heated.
c. An iron bar rusts.
d. Aluminum is a silver-colored, non-magnetic metal

Chapter 2: Atoms, Molecules, & Ions

2.1 The Early History of Chemistry

A. Take notes: Outline the early history of chemistry from 400 BC through Priestly. Include important people, contributions, and vocabulary

B. Memorize/Flashcards - n/a

C. Practice Problems

1. Describe the contributions of Joseph Priestly, his relationship to Benjamin Franklin, and why he left England.

2.2 Fundamental Chemical Laws

A. Take notes: Define and give an example of each law:

- Law of Conservation of Mass
- Law of Definite Proportion (Proust's Law)
- Law of Multiple Proportions

B. Memorize/Flashcards - n/a

C. Practice Problems

2. In the illustration at right, two substances massed prior to a chemical reaction. After the two substances are combined, a white precipitate is formed in an aqueous solution. If the substances are again massed, what will the new mass be? Which chemical law is demonstrated in this example?

2.3 Dalton's Atomic Theory

A. Take notes: Four parts of Dalton's Atomic Theory

B. Memorize/Flashcards – Seven elements that occur as diatomic molecules (pg. 50 in margin). ***Hint: Use a mnemonic device like 'I Have No Bright Or Clever Friends'***

C. Practice Problems - n/a

2.4 Early Experiments to Characterize the Atom

A. Take notes:
- J.J. Thomson, cathode-ray tubes, and discovery of electron
- Henri Becquerel, radiation, gamma rays, alpha and beta particles
- Ernest Rutherford, gold foil experiment, discovery of nucleus

B. Memorize/Flashcards - n/a

C. Practice Problems

3. What evidence led to the conclusion that cathode rays had a negative charge?

4. Consider Ernest Rutherford's gold foil experiment illustrated in Fig. 2.12. How did the results of this experiment lead Rutherford away from the plum pudding model of the atom to propose the nuclear model of the atom?

2.5 The Modern View of Atomic Structure: An Introduction

A. Take notes
- Neutron, proton, electron – mass, charge, location in atom
- Define atomic number, mass number, isotope
- Symbol used to show atoms and their mass and atomic numbers

B. Memorize/Flashcards - n/a

C. Practice Problems

5. T or F The chemical properties of element arise from their neutrons

6. T or F In nature, most elements exist as mixtures of their isotopes

7. Write the symbol for an element with 85 protons, 125 neutrons, and 85 protons

8. Use your periodic table to determine the number of protons, neutrons, and electrons in the following elements:

Mg	U	Fe	Rn

2.6 Molecules and Ions

A. Take notes:

- Define chemical bond
- Define, relate, and give examples of: covalent bonds, molecule, chemical formula, structural formula
- Define and relate: ionic bond, ion, anion, cation

B. Memorize/Flashcards - n/a

C. Practice Problems

9. Which of the following explain how an ion is formed? Explain your choice
    a. adding or subtracting protons to/from an atom
    b. adding or subtracting neutrons to/from an atom
    c. adding or subtracting electrons to/from an atom

10. Circle which of the following elements would form a cation:

I	Ca	Rb	Ne	Ni	O

2.7 An Introduction to the Periodic Table

A. Take notes:

- Physical and chemical properties of metals
- Physical and chemical properties of nonmetals
- Define group (family) and period

B. Memorize/Flashcards - n/a

C. Practice Problems

11. Cut out the periodic table on the last page of the notes, tape it into your notebook, and color-code and/or label the following special groups or periods of the periodic table: alkali metals, alkaline earth metals, halogens, noble gases, transition metals, lanthanides, actinides, line separating metals from non-metals

2.8 Naming Simple Compounds

A. Take notes:

- Rules for naming Binary Ionic Compounds (Type I)
- How to name Binary Ionic Compounds (Type II)
- Naming Ionic Compounds with Polyatomic Ions

- Naming Binary Covalent Compounds (Type III)
- Naming acids

B. Memorize/Flashcards

- Table 2.6 – Prefixes Used to Indicate Number in Chemical Names
- Common Polyatomic Ions – know formula AND charge for: hydroxide, nitrate, acetate, cyanide, permanganate, carbonate, sulfate, dichromate, phosphate, ammonium

C. Practice Problems

12. Name the following compounds:

a. CuI ______________________________

b. CuI_2 ______________________________

c. CoI_2 ______________________________

d. Na_2CO_3 ____________________________

e. $NaHCO_3$ ___________________________

f. S_4N_4 ______________________________

g. $SeCl_4$ _____________________________

h. $NaOCl$ ____________________________

i. $BaCrO_4$ ______________________________

j. NH_4NO_3 ____________________________

13. Write formulas for the following:

a. sulfur difluoride _______________

b. sulfur hexafluoride _____________

c. sodium dihydrogen phosphate _____________

d. lithium nitride _____________

e. chromium (III) carbonate ___________

f. tin (II) fluoride __________

g. ammonium acetate _____________

h. cobalt (III) nitrate _____________

i. potassium chlorate ____________

j. sodium hydride ___________

k. calcium phosphate _____________

Chapter 3: Stoichiometry

3.1 Counting by Weighing – Read this section

3.2 Atomic Masses

A. Take notes:
   - Modern system of atomic mass and the use of the mass spectrometer
   - Define isotope
   - Average atomic mass – define and how to calculate (I called this a weighted average in Unit 3 of Chemistry)

B. Memorize/Flashcards - n/a

C. Practice Problems

   1. An element consists of 1.40% of an isotope with mass 203.973 u, 24.10% of an isotope with mass 205.9745 u, 22.10% of an isotope with mass

206.9759 u, and 52.40% of an isotope with mass 207.9766 u. Calculate the average atomic mass, AND identify the element.

3.3 The Mole

A. Take notes:
- Mole, definition and relate to Avogadro's number

B. Memorize/Flashcards - n/a

C. Practice Problems - n/a

You will do practice problems using this concept after section 3.4. You can use the examples in the book as a guide OR use the **Mole Road Map** from Unit 9 of Chemistry (I hope you still have your notebook!)

3.4 Molar Mass

A. Take notes:
- Define molar mass, unit of molar mass
- Work through Interactive Examples 3.6 – 3.8. Also refer to Unit 9 of Chemistry. Take notes as needed

B. Memorize/Flashcards - *****We will be rounding all molar masses to the nearest hundredth.***

C. Practice Problems

2. Calculate the mass 500. atoms of iron

3. A diamond contains 5.0×10^{21} atoms of carbon. How many moles of carbon is this?

4. Aluminum metal is produced by passing an electric current through a solution of aluminum oxide (Al_2O_3) dissolved in molten cryolite (Na_3AlF_6). Calculate the molar mass of each substance.

5. a. Calculate the molar mass of calcium phosphate, $Ca_3(PO_4)_2$

   b. How many moles would be in 1.00 gram of calcium phosphate?

6. Freon-12 (CCl_2F_2) is used as a refrigerant in air conditions. Calculate the number of molecules of Freon-12 in 5.56 mg of Freon-12.

3.5 Learning to Solve Problems – Read this section

3.6 Percent Composition of Compounds

A. Take notes:
   - Percent composition/mass percent
   - Example of how to find percent composition (if you refer to your Chemistry notes from Unit 9, the method I showed uses 1 less step)

B. Memorize/Flashcards - n/a

C. Practice Problems

7. Calculate the percent composition by mass of baking soda, aka sodium bicarbonate.

8. In 1987 the first substance to act as a superconductor at a temperature above that of liquid nitrogen (77K) was discovered. The formula of this substance is $YBa_2Cu_3O_7$. Calculate the percent composition by mass of Yttrium in this substance.

3.7 Determining the Formula of a Compound

A. Take notes:

- Define empirical formula; steps for determining empirical formula
- Define molecular formula; steps for determining molecular formula

(Again, it may be helpful to refer to notes and examples from Unit 9 Chemistry)

B. Memorize/Flashcards - n/a

C. Practice Problems

9. A compound containing only sulfur and nitrogen is 69.6% by mass. What is the empirical formula of the compound?

10. Determine the molecular formula of a compound that contains 26.7% P, 12.1% N, and 61.2% Cl, and has a molar mass of 580 g/mol.

3.8 Chemical Equations

A. Take notes:
    - Define reactants and products; use an example to identify reactants and products in an equation

B. Memorize/Flashcards – symbols (and their meanings) used to identify physical states (solid, liquid, gas, aqueous solution)

C. Practice Problems - n/a

3.9 Balancing Chemical Equations

A. Take notes:
    - If you can find your notes for 'Tips & Tricks for Balancing Equations', it would be a good idea to fill them in. If not, consider leaving a space and we will fill them in as a class.

B. Memorize/Flashcards - n/a

C. Practice Problems

11. Write out AND balance the equations in # 102 on pg. 131

    a.

    b.

    c.

    d.

12. Liquid silicon tetrachloride is reacted with very pure solid magnesium, producing solid silicon and solid magnesium chloride. Write a complete, balanced equation for this reaction using proper symbols:

<u>3.10 Stoichiometric Calculations: Amounts of Reactants and Products</u>

A. Take notes:

- Read the section and follow examples 3.15 and 3.16. Also refer to the Problem-Solving Strategy on pg. 111. You may also want to refer to the stoichiometry notes from Unit 10 of Chemistry. The Stoichiometry Road Map may also help:

B. Memorize/Flashcards - n/a

C. Practice Problems – refer to page 132 and complete the following:

105. What mass of iron (III) oxide with sufficient aluminum must be used to produce 15.0 g of iron?

106. Complete as is – don't forget to balance the equation first!

109. Complete a.

3.11 The Concept of Limiting Reactants

A. Take notes – Read this section. This is our starting point in fall!
- Define limiting reactant – leave a page or two for more notes on this
- Define theoretical yield, actual yield (a.k.a. experimental yield),percent yield
- Formula for calculating percent yield

B. Memorize/Flashcards - n/a

C. Practice Problems

13. A reaction with a theoretical yield of 9.23 g produced 7.89 g of product. What is the percent yield for this reaction?

14. 5.96 g of ammonia (NH_3) reacts completely according to the following reaction:

$$2\ NH_3(g) + CO_2(g) \longrightarrow CN_2OH_4(s) + H_2O\ (l)$$

After running the reaction in lab, 10.2 g of urea (CN_2OH_4) is formed. What is the percent yield for the reaction?