Dear AP Chemistry Student and Parent,
Welcome to AP Chemistry!! I am so excited about beginning this journey with you! For those of you that do not know me, my name is Chris Smith, and I will be your instructor for AP Chemistry for the 2024-25 school year. I am a graduate of the University of Georgia and have been teaching for 19 years, Chemistry for 15 years. The 2024-2025 school year will be my 9th year teaching AP Chemistry.

Overview of Course Expectations:

AP Chemistry is a college level course. **Students may earn up to 8 hours college credit for successful completion of this course and a good score on the AP Exam.** It is a time-consuming and challenging, yet extremely rewarding, course. This course moves at a very fast pace and classroom attendance is a MUST.

I will do my very best to provide a college level course/experience which not only prepares you for the AP exam, but provides a solid foundation in chemistry. I also intend for it to be fun!!

The AP Chemistry Exam is generally on the first day of the first week of AP testing in May, but that is subject to change. To be successful on the AP exam, students will need to spend on average five to ten additional hours per week outside of class working on AP Chemistry. This statement is not meant to discourage, but to point out and state the truth to avoid any misconceptions about the high expectations for this course.

Summer Assignment: Attached to this letter, you will find a packet that is divided into three sections. Section A consists of basic information that students need to memorize, and Section B contains practice of skills from Honors Chemistry. Section C is an appendix of useful information and resources. The memorization and practice needs to be COMPLETED prior to school starting. The majority of the material required in this assignment is review material that students should have learned in their first year chemistry class. Because this is a challenging problem-solving course, and for some of you, a year may have passed since you have had a chemistry course, it is imperative that you come to class the first day with Sections A and B completed. ***This summer assignment is not optional,*** and completing the assignment in a thorough and focused manner will contribute to a student's success in this course and on the AP Chemistry exam. I recommend that you work on this assignment periodically. **Do not wait until the last week of summer to begin. There will be a test covering the concepts included in this summer assignment during the first two weeks of school.**

I look forward to getting to know each of you! We will have fun and we will work hard. Students will receive a detailed course syllabus when school resumes in August. Please feel free to send me an e-mail over the summer if you have any questions or comments. I will check e-mail weekly during the summer months. I will likely send out Remind notifications as well, so make sure you signed up please.

My e-mail address is: **christopher.smith@cobbk12.org**

Have a great summer! I look forward to working with you in the upcoming year.
Sincerely,
Chris Smith

Section A: Information to be Memorized by the 1st Day of School

1. Know metric and SI base units and prefixes. Be able to convert between units within the metric system.
   - Metric units are: meter, liter, gram, second, mole (SI unit for length = meter, for mass = kg, for volume = m^3)
   - Prefixes to change the magnitude of metric units are: Giga-, Mega-, kilo-, hecto-, deka-, deci-, centi-, milli-, micro-, nano-

2. Rules for significant figures within measurements
   - All non-zeros digits are significant
   - Zeros before the first nonzero are NEVER significant
   - Zeros that are captive between two non-zeros are ALWAYS significant
   - Zeros after the last non-zero are significant ONLY IF there is a decimal point

3. Polyatomic Ions (most you learned in Honors Chemistry, but the ones you didn't are bolded)

-1	-2	-3
$C_2H_3O_2^-$ acetate	SO_4^{2-} sulfate	PO_4^{3-} phosphate
NO_3^- nitrate	SO_3^{2-} sulfite	**AsO_4^{3-} arsenate**
NO_2^- nitrite	CO_3^{2-} carbonate	**AsO_3^{3-} arsenite**
CN^- cyanide	**$C_2O_4^{2-}$ oxalate**	
OCN^- cyanate	**CrO_4^{2-} chromate**	
SCN^- thiocyanate	**$Cr_2O_7^{2-}$ dichromate**	
MnO_4^- permanganate	**$S_2O_3^{2-}$ thiosulfate**	
OH^- hydroxide	O_2^{2-} peroxide	
HSO_4^- bisulfate	**SeO_4^{2-} selenate**	
HCO_3^- bicarbonate	**HPO_4^{2-} hydrogen phosphate**	+1
O_2^- superoxide		NH_4^+ ammonium
$H_2PO_4^-$ dihydrogen phosphate		
ClO_4^- perchlorate *		
ClO_3^- chlorate *		
ClO_2^- chlorite *		
ClO^- hypochlorite *		

* Br, I, and F may be substituted

4. Strong Acids (for all practical purposes, all others are weak acids): HCl, HBr, HI, H_2SO_4, HNO_3, $HClO_3$, $HClO_4$

5. Strong Bases (for all practical purposes all others are weak):
   Group I hydroxides and Group II hydroxides (except $Be(OH)_2$ and $Mg(OH)_2$)

6. Solubility Rules – you really only need to memorize the first two rows of the first table, but be able to use the table for all other solubilities

Soluble Ionic Compounds	Exceptions
Group IA and ammonium (NH_4^+) salts	none
nitrates (NO_3^-) and acetates ($C_2H_3O_2^-$)	none
Chlorides (Cl^-), bromides (Br^-) and iodides (I^-)	Compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}
Sulfates (SO_4^{2-})	Compounds of Sr^{2+}, Ba^{2+}, Ca^{2+}, and Pb^{2+}

Insoluble Ionic Compounds	Exceptions
Sulfides (S^{2-})	Compounds of NH_4^+, Group IA ions, or Ca^{2+}, Sr^{2+}, and Ba^{2+}
Carbonates (CO_3^{2-})	Compounds of NH_4^+ and Group IA ions
Phosphates (PO_4^{3-})	Compounds of NH_4^+ and Group IA ions
Hydroxides (OH^-)	Compounds of NH_4^+, Group IA ions, or Ca^{2+}, Sr^{2+}, and Ba^{2+}

7. Rules for assigning (determining) oxidation numbers:

- **Rule 1:** Atoms in a pure element have an oxidation number of zero.
- **Rule 2:** The more electronegative element in a binary compound is assigned the number equal to the negative charge it would have as an anion. The less-electronegative atom is assigned the number equal to the positive charge it would have as a cation.
- **Rule 3:** Fluorine has an oxidation number of -1 in all of its compounds because it is the most electronegative element.
- **Rule 4:** Oxygen has an oxidation number of -2 in almost all compounds.
  Exceptions:
  Peroxides, such as H_2O_2, in which its oxidation # is -1
  When oxygen is in compounds with halogens, such as OF_2, its oxidation # is +2.
- **Rule 5:** Hydrogen has an oxidation # of +1 in all compounds that are more electronegative than it; it has an oxidation # of -1 in compounds with reactive metals (hydrides).
- **Rule 6:** The algebraic sum of the oxidation numbers of all atoms in a neutral compound is zero.
- **Rule 7:** The algebraic sum of the oxidation numbers of all atoms in a polyatomic ion is equal to the charge of the ion.

- **Rule 8:** Rules 1-7 apply to covalently bonded atoms; however, oxidation numbers can also be assigned to atoms in ionic compounds.

Section B: Calculations and Short Answer

I. Dimensional Analysis and Significant Figures Review

1. How many significant figures (digits) are in the following measurements?
   a. 1234 g
   b. 0.023 L
   c. 890 m
   d. 91010 kg
   e. 9010.0 m/s
   f. 3.40×10^4 mol
   g. 7.030×10^{-3} cm

2. Use **factor labeling** method (dimensional analysis) to convert the following. Show all work!
   a.
   515 m = ___ miles.
   b. 200 in = ____ meters
   c. 325 days = _____ seconds.
   d. 20 gallons = _____ ml
   e. 3 meters into centimeters
   f. 10 kilometers into meters
   g. 15,050 milligrams into grams
   h. 3,264 milliliters into liters
   i . 9,674,444 grams into kilograms

3. Classify each of the following as units of mass, volume, length, density, energy, or pressure.
   a. mg
   b. mL
   c. cm^3
   d. mm
   e. kg/m^3
   f. kJ
   g. atm
   h. cal

4. Most laboratory experiments are performed at room temperature at 25˚C. Express this temperature in:
   a. ˚F
   b. K .

5. A cylinder rod formed from silicon is 16.8 cm long and has a mass of 2.17 kg. The density of silicon is 2.33 g/cm^3. What is the diameter of the cylinder? (the volume of cylinder is given by $\pi r^2 h$, where r is the radius and h is the length)

6. How many significant figures are in each of the following?
   a. 1.92 mm
   b. 0.030100 kJ
   c. 6.022 x10$_{23}$ atoms
   d. 460.00 L
   e. 0.00036 cm_3
   f. 100
   g. 1001
   h. 0.001
   i. 0.010

7. Record the following in correct scientific notation:
   a. 350,000,000 cal
   b. 0.0000721 mol
   c. 0.0000000809 Å
   d. 765,400,000,000 atoms

8. Calculate the following to the correct number of significant figures.
   a. 1.27 g / 5.296 cm_3
   b. 12.235 g / 1.01 L
   c. 12.2 g + 0.38 g
   d. 17.3 g + 2.785 g
   e. 2.1 x 3.21
   f. 200.1 x 120
   g. 17.6 + 2.838 + 2.3 + 110.77

9. If you drive 154 miles in 3.0 hours, what is your average speed in meters per minute?

10. Calculate the mass of a sample of copper that occupies $4.2x10^3 cm^3$ if the density of copper is $8.94 g/cm^3$.

II. Atomic Structure Review

1. Fill in the following table:

Element or ion	Complete Symbol	# protons	# neutrons	# electrons
Fe-55				
K^+				
		27		25
O^{2-}				
		11	12	11
Pb-208				

2. Find the average atomic mass of an element if, out of 100 atoms, 5 have a mass of 176amu, 19 have a mass of 177amu, 27 have a mass of 178amu, 14 have a mass of 179amu and 35 have a mass of 180amu.

3. Strontium consists of four isotopes with masses and percent abundances as follows: 83.9134amu (0.5%), 85.9094amu (9.9%), 86.9089amu (7.0%), and 87.9056amu (82.6%). Calculate the atomic mass of strontium.

4. Write the complete and abbreviated ground state electron configurations for:
   a. Strontium
   b. Iron
   c. Sulfur
   d. Neodymium

III. Nomenclature Review

Forming binary ionic compounds

A. In a binary ionic compound the total positive charges must equal the total negative charges. The best way to write correct formula units for ionic compounds is to use the "least common multiple" method.

B. Sample problem: What ionic compound would form when calcium ions combine with bromide ions?

Steps to writing ionic formulas:

1. Write the ions with their charges, cations are always first. Ex: Ca^{2+} Br^{1-}
2. Determine the least common multiple of the charges. This is the total positive and total negative value that would result in a neutral compound. Ex: LCM = 2
3. Use subscripts after each element symbol to indicate the number of that ion needed to reach the least common multiple of charge. Ex: $CaBr_2$

Naming binary ionic compounds

A. Combine the names of the cation and the anion.
B. Example; $BaBr_2$ is named barium bromide.

Naming binary ionic compounds that contain polyatomic ions
A. The polyatomic ions on your common ions list should be memorized.
B. The most common oxyanions – polyatomic anions that contain oxygen, end in *–ate*. Oxyanions with one less oxygen end in *–ite*. For example:

NO_3^{-1} is nitrate SO_4^{2-} is sulfate
NO_2^{-1} is nitrite SO_3^{2-} is sulfite

C. Anions with one less oxygen than the *–ite* ion are given the prefix *hypo-*.
D. Anions with one more oxygen than the *–ate* ion are given the prefix *per-*.

ClO^{-1} is hypochlorite ClO_3^{-1} is chlorate
ClO_2^{-1} is chlorite ClO_4^{-1} is perchlorate

E. Naming compounds with polyatomic ions is the same as naming other compounds, just name the cation and then the anion. If there is a transition metal involved, be sure to check the charges to identify which ion (+1, +2, +3, +4….) it may be so that you can put the correct Roman numeral in the name.

Naming binary molecular compounds
With molecules, the prefix system is used.

Number Prefix	Number Prefix
1 mono-	7 hepta-
2 di-	8 octa-
3 tri-	9 nona-
4 tetra-	10 deca-
5 penta-	11 undeca-
6 hexa-	12 dodeca-

A. The less-electronegative element is always written first. It only gets a prefix if it has more than one atom in the molecule.
B. The second element gets the prefix and the ending *–ide*.
C. The *o or a* at the end of the prefix is dropped when the word following the prefix begins with another vowel, for example monoxide or pentoxide.

III. Nomenclature Review (continued)

1. Write formulas for the following substances:

a. Barium sulfate _____
b. Ammonium chloride _____
c. Chlorine monoxide _____
d. Silicone tetrachloride _____
e. Magnesium fluoride _____
f. Sodium oxide _____
g. Sodium peroxide_____
h. Copper (I) iodide _____
i. Zinc sulfide _____
j. Potassium carbonate _____
k. Hydrobromic acid _____
l. Perchloric acid _____
m. Lead (II) acetate _____
n. Sodium permanganate _____
o. Lithium oxalate _____
p. Potassium cyanide _____
q. Iron (III) hydroxide _____
r. Silicone dioxide _____
s. Nitrogen trifluoride _____
t. Chromium (III) oxide _____
u. Calcium chlorate _____
v. Sodium thiocyanate _____
w. Cobalt (III) nitrate _____
x. Nitrous acid _____
y. Ammonium phosphate _____
z. Potassium chromate _____

2. Name each of the following compounds (Give acid names where appropriate)

a. $CuSO_4$ _____
b. PCl_3 _____
c. Li_3N _____
d. $BaSO_3$ _____

e. N_2F_4 _____
f. $KClO_4$ _____
g. NaH _____
h. $(NH_4)_2Cr_2O_7$ _____
i. HNO_2 _____
j. Sr_3P_2 _____
k. $Mg(OH)_2$ _____
l. Al_2S_3 _____
m. AgBr _____
n. P_4O_{10} _____
o. $HC_2H_3O_2$ _____
p. CaI_2 _____
q. MnO_2 _____
r. Li_2O _____
s. FeI_3 _____
t. Cu_3PO_4 _____
u. PCl_3 _____
v. NaCN _____
w. Cs_3N _____
x. $Zn(NO_3)_2$ _____
y. N_2O _____
z. HF _____

Nomenclature Review (continued)
Practice with acids!

Remember:

-IC from *–ATE; -OUS* from *–ITE; HYDRO-, -IC* from *-IDE*

Complete the Following Table:

Name of Acid	Formula of Acid	Name of Anion
*hydro*chlor*ic*	*HCl*	chlor*ide*
sulfuric acid	H_2SO_4	sulfate
	HI	
		sulf*ite*
chlor*ous* acid		
		nitr*ate*
	$HC_2H_3O_2$ *or* CH_3COOH	
*hydro*brom*ic* acid		
		sulf*ide*
	HNO_2	
chrom*ic* acid		
		phosphate

IV. Balancing Equations Review

Balance the following equations by adding coefficients as needed. Some equations may already be balanced.

1) __ C_6H_6 + __ O_2 → __ H_2O + __ CO_2

2) __ NaI + __ $Pb(SO_4)_2$ → __ PbI_4 + __ Na_2SO_4

3) __ NH_3 + __ O_2 →__ NO + __ H_2O

4) __ $Fe(OH)_3$ → __ Fe_2O_3 + __ H_2O

5) __ HNO_3 + __ $Mg(OH)_2$ → __H_2O + __ $Mg(NO_3)_2$

6) __ H_3PO_4 + __ NaBr → __ HBr + __ Na_3PO_4

7) __ C + __ H_2 → __ C_3H_8

8) __ CaO + __ MnI_4 → __ MnO_2 + __ CaI_2

9) __ Fe_2O_3 + __ H_2O → __ $Fe(OH)_3$

10) __ C_2H_2 + __ H_2 → __ C_2H_6

11) __ VF_5 + __ HI → __ V_2I_{10} + __ HF

12) __ OsO_4 + __ $PtCl_4$ → __ PtO_2 + __ $OsCl_8$

13) __ CF_4 + __ Br_2 → __ CBr_4 + __ F_2

14) __ Hg_2I_2 + __ O_2 → __ Hg_2O + __ I_2

15) __ $Y(NO_3)_2$ + __ $GaPO_4$ → __ YPO_4 + __ $Ga(NO_3)_2$

V. Chemical Reactions

In AP Chemistry, most of the reaction we write are called "net ionic." But before we can do that, you need to review and memorize some basic reaction types. For some basic review, go to the following website:
http://misterguch.brinkster.net/6typesofchemicalrxn.html
Now try these sample problems from the website.
Give the type for each of the following reactions:

1) $NaOH + KNO_3 \rightarrow NaNO_3 + KOH$ _____

2) $CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$ _____

3) $2\ Fe + 6\ NaBr \rightarrow 2\ FeBr_3 + 6\ Na$ _____

4) $CaSO_4 + Mg(OH)_2 \rightarrow Ca(OH)_2 + MgSO_4$ _____

5) $NH_4OH + HBr \rightarrow H_2O + NH_4Br$ _____

6) $Pb + O_2 \rightarrow PbO_2$ _____

7) $Na_2CO_3 \rightarrow Na_2O + CO_2$ _____

Memorize these types of decomposition reactions:
1. Metallic carbonates, when heated, form metallic oxides and $CO_{2(g)}$.
EX. $CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$

2. Most metallic hydroxides, when heated, decompose into metallic oxides and water.
EX. $Ca(OH)_{2(s)} \rightarrow CaO_{(s)} + H_2O_{(g)}$

3. Metallic chlorates, when heated, decompose into metallic chlorides and oxygen.
EX. $2KClO_{3(s)} \rightarrow 2KCl_{(s)} + 3O_{2(g)}$

4. Some acids, when heated, decompose into nonmetallic oxides and water.
EX. $H_2SO_4 \rightarrow H_2O_{(l)} + SO_{3(g)}$

5. Some oxides, when heated, decompose.
EX. $2HgO_{(s)} \rightarrow 2Hg_{(l)} + O_{2(g)}$

6. Some decomposition reactions are produced by electricity.
EX. $2H_2O_{(l)} \rightarrow 2H_{2(g)} + O_{2(g)}$
EX. $2NaCl_{(l)} \rightarrow 2Na_{(s)} + Cl_{2(g)}$

Memorize these types of synthesis reactions:
1. Metal + oxygen → metal oxide
EX. $2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$

2. Nonmetal + oxygen → nonmetallic oxide
EX. $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$

3. Metal oxide + water → metallic hydroxide
EX. $MgO_{(s)} + H_2O_{(l)} \rightarrow Mg(OH)_{2(s)}$

4. Nonmetallic oxide + water → acid
EX. $CO_{2(g)} + H_2O_{(l)} \rightarrow$; $H_2CO_{3(aq)}$

5. Metal + nonmetal → salt
EX. $2\ Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$

6. A few nonmetals combine with each other.
EX. $2P_{(s)} + 3Cl_{2(g)} \rightarrow 2PCl_{3(g)}$

Now try these decomposition reactions: (Rewrite as a balanced equation with the products predicted):
1. barium hydroxide (heated)

_____
2. sodium carbonate (heated)

_____
3. lithium chlorate (heated)

_____
4. electrolysis of aluminum oxide

_____
5. sulfuric acid heated gently
_____

Now try these synthesis reactions: (Rewrite as a balanced equation with the products predicted):
1. magnesium burned in oxygen

_____
2. hydrogen gas + nitrogen gas

_____
3. sulfur burned (complete combustion)

_____
4. calcium oxide added to water

_____

Chemical Reactions (continued)

For each equation below, identify the type (synthesis, decomposition, single replacement, double replacement, or combustion), predict the products, and then write the balanced reaction. Remember to use the solubility rules for double replacement reactions and the activity series for single replacement reactions. Hint: when writing these reactions, ignore all of the information about heat, 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.
$$2AgNO_{3(aq)} + MgI_{2(aq)} \rightarrow 2AgI_{(s)} + Mg(NO_3)_{2(aq)}$$

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. Solutions of lead nitrate and calcium iodide are combined.

8. Sulfuric acid is combined with sodium hydroxide.

9. Isopropyl alcohol (C_3H_7OH) is burned in oxygen.

10. Iron metal shavings are added to hydrochloric acid.

11. Solid sodium carbonate is heated in a crucible.

12. Sodium metal is added to distilled water.

VI. Stoichiometry Review

Show all of your work for the following problems:

1. Find the mass percent of nitrogen in each of the following compounds:

a. NO

b. NO_2

c. N_2O_4

d. N_2O

2. Benzene contains only carbon and hydrogen and has a molar mass of 78.1 g/mol. Analysis shows the compound to be 7.74 % hydrogen by mass. Find the empirical and molecular formulas of benzene.

3. Calcium carbonate decomposes upon heating, producing calcium oxide and carbon dioxide.

a. Write a balanced chemical equation for this reaction.

b. How many grams of calcium oxide will be produced after 12.25 grams of calcium carbonate are completely decomposed?

c. What is the volume of carbon dioxide gas produced 12.25 grams of calcium carbonate at STP?

d. What is the volume of carbon dioxide in L if the pressure is 785mm Hg and the temperature is 30°C? (R =0.0821 L·atm/mol·K) (HINT: use the ideal gas law to relate mole amount to volume)

4. Hydrogen gas and bromine gas react to form hydrogen bromide gas.

a. Write a balanced equation for this reaction.

b. 3.2 grams of hydrogen react with 9.5 grams of bromine. Which is the limiting reagent?

c. How many grams of hydrogen bromide gas can be produced using the amounts in (b)?

d. How many grams of excess reactant are left unreacted?

e. What volume of HBr, measured at STP is produced in (b)?

5. When ammonia gas, oxygen gas and methane gas (CH_4) are combined, the products are hydrogen cyanide gas and water.

a. Write a balanced chemical equation for this reaction.

b. Calculate the mass of each product produced when 225 grams of oxygen gas is reacted with an excess of the other two reactants.

c. If the actual yield of the experiment in (b) is 105 grams of HCN, calculate the percent yield.

VII. Solutions Review

1. What mass of magnesium nitrate is needed to prepare 500.mL of 2.41M magnesium nitrate?

2. What volume of a 0.465M solution of calcium chloride contains 1.28g of solute?

3. Determine the Molarity of a solution containing 4.20g of ammonium acetate in 750.0mL of solution.

4. A solution contains 1.50moles of ethanol in 1200g of water.
   a. Calculate the percent by mass of ethanol (C_2H_6O) in the solution

   b. Calculate the molality of the solution.

5. Calculate the mass of Ag_2S produced when 125 mL of 0.200 M $AgNO_3$ is added to excess Na_2S solution. ("Excess" Na_2S means that you have more than you need - therefore, $AgNO_3$ is limiting.)

6. What mass of $Mg(OH)_2$ is produced when 100 mL of 0.42 M $Mg(NO_3)_2$ is added to excess NaOH solution?

7. What mass of $Fe(OH)_3$ is produced when 35 mL of 0.250 M $Fe(NO_3)_3$ solution is mixed with 55 mL of a 0.180 M KOH solution? (this is a **limiting reactant problem**).

8. When solutions of potassium iodide and lead (II) nitrate are combined, the products are potassium nitrate and lead (II) iodide.

   a. Write a balanced equation for this reaction, including (aq) and (s).

   b. Calculate the mass of precipitate produced when 50.0mL of 0.45M potassium iodide solution and 75mL of 0.55M lead (II) nitrate solution are mixed.

9. Calculate the volume of 0.50M potassium iodide required to react completely with 50.0mL of 0.50M lead (II) nitrate.

Section C: Appendix

How to Study Chemistry

You, like many students, may view chemistry and physics as difficult. You may seem overwhelmed by new terms, ideas, equations, and methods of problem-solving. You may find it hard to transpose and apply the information from your textbook and classes to actual problems. We hope that this document can help you. It will act as a guide for:

- reading science textbooks
- note taking
- test preparation
- problem solving techniques
- preparation for final exams

Overview

It's important to recognize that chemistry is a problem-solving class. Major themes and principles are stressed, and one major goal is that the student will be able to apply these principles to understand and solve problems. You should understand that in a science course, *a significant portion of your time will be spent solving problems*. THIS IS A MATHEMATICS BASED COURSE, and mathematics will be used to solve problems.

An overview of the topic can help you organize your thoughts and allow you to use your studying time to its maximum benefit. Your goal is not to remember formulas – but to understand the underlying principles. It is inevitable that you will forget a formula, and if you have an understanding of the underlying principle, you can generate the formula for yourself.

Take these steps to get an overview of the topics that will be covered. Use the overview to "get the big picture" and integrate new material into your overall view of the subject.

1. Examine the course syllabus carefully and read the description of the course that the teacher provides. Ask about the underlying themes that will be covered. (Your teacher will not choose topics at random, rather "there is a method to the madness.")

2. Preview the textbook:

   A. Read the introduction and table of contents.

   B. Read any notes to the student that are included and the preface.

   C. Check the course outline to see what chapters are assigned and which are omitted. If they are not assigned in the same order as in the table of contents, ask about your teacher's decision to alter the order of presentation. However, remember that the textbook is not the "be all and end all" of the topic.

   D. Look at the appendix of the book.
      1. Tables and location.
      2. Are answers to some or all of the problems there?

E. Glance at some of the problems in the textbook. How are they worked out? Is this the same approach that your teacher uses?

Reading Your Science Textbook

Most students make the mistake that reading a science textbook is like reading a mystery novel. Read it as fast as possible, cover to cover (or chapter by chapter) and then figure out who did it in the last chapter. This approach will not work in a science class. You must learn to be an *active reader*.

An active reader is one that pauses to think about what was just read. They relate the material just read to previous material. They make sure that they understand the material and its application to the topic at hand. They ask themselves questions so that they are sure that they understand the material. Reading a science textbook requires a pen, paper, and a calculator. Work the practice problems and the examples. This is the only way that you can make sure that you understand the material.

Reading the text and solving homework problems is a cycle: Questions lead to answers that lead back to more questions. An entire chapter will often be devoted to the consequences of a single basic principle. You should look for these basic principles. Science is built around these basic laws. All of the problems that you will face in a science course can be analyzed by means of one or more of these laws.

Many times in science, a problem is first analyzed in great detail. Then the results of the problem are generalized into more abstract concepts. You should understand the generalizations made, and you should refer back to the specific case that was studied and make sure that you understand how the general theory applies to the specific problem.

Guidelines to follow for reading:

1. Use the topic or section headings as a guide. Use the preview that you did prior to the class. Understand why you are doing this chapter or topic. Start by reading the chapter summary first. This will give you an idea of what you are trying to read and what you are to get out of reading the chapter.

2. Make connections between your notes and your readings. Write down connections between the example problems done in class and the reading.

3. Write down ANY questions that you have in your notes. Leave room for answers and make sure that your teacher answers those questions.

4. Make a list of all terms in the chapter.

5. Read the homework problems first. If specific homework problems have not yet been assigned, select several and look these over. This gives you an idea of why you are studying this topic and what your teacher expects you to be able to do.

6. Read actively with questions in mind. A passive approach wastes your time. Read with a pencil and paper beside the book to jot down questions and notes. **Read to learn, not to cover material.** If it takes you three times to read through a section, so be it...just make sure you understand it. **DO THE PRACTICE PROBLEMS**...they are there for a reason.

7. Add notes from the textbook into the margins of your lecture notes.

8. *Read the margins of your textbook*. Many important connections are covered in the margins and sometimes mnemonic devices are discussed.

9. During your reading you will notice sections, equations, or ideas that apply directly to assigned problems. After you have read such a section, stop and try to apply it to the homework problem. Often textbooks break their homework problems into the sections of the chapter. Also, many textbooks "pair" their questions with two (or more) problems covering the same material. Look over the problem that wasn't assigned – it just may be on the test, and it is one way that you can check yourself to see if you truly understand the concept.

The interplay of reading and problem solving is part of the cycle:

question → answer → question → answer.

Active reading helps your understanding and is far more effective than reading alone.

Lectures and Notetaking

Just as you are to be an active reader, you must be an active listener. If you are, then your notes will be more complete and accurate.

1. Read ahead. Ask the teacher what will be covered next (or check the syllabus) and preview the sections to be covered. Make a list of the new terminology, units of measurement, and concepts that you will encounter. Look at the diagrams, figures, and tables - try to formulate what point each is trying to illustrate.

2. Re-read your notes from the previous class. They are the basis for your new material.

3. Listen very carefully at the beginning and end of the class. This is when many of the key concepts will be introduced or reiterated. Too often, students miss this information.

4. Use a list of abbreviations consistently throughout your notes. Make a table of them for future reference. Just remember that these are shortcuts and not to be used in formal writing.

Examples are:

A. → Leads to
B. ⇒ Implies
C. b/c Because
D. b/f Before
E. ∀ For all
F. ∃ There exists
G. ∋ Such that
H. ∴ Therefore

5. Use a comfortable note taking method for science classes. Use the margin for later comments, questions, and textbook references.

6. Take notes in outline form showing major topics, sub-topics and their relationship.

7. Copy not only what is on the board but also the important points that the teacher talks about. When copying diagrams, artwork is less important than completeness.

8. If you fall behind in your note taking, leave a space in your notes and go on. You can fill in your notes later with the help of a classmate or your textbook.

9. The only dumb question is the one that you don't ask. Don't be embarrassed to ask your teacher questions. If you don't understand, **ASK** – that is what your teacher gets paid for. Since your teacher is human, they can even make a mistake when writing something on the board.

10. That night, review and edit your notes. You don't need to rewrite them. At this time you may want to add an outline to your notes. Also, use the margin to make reference from your textbook to your notes.

11. As you review your notes, write out any questions that come to mind. Leave space for the answer and then ask your teacher the next day.

12. Remember that most of your notes will be examples of problems. *Copy the problem down completely, follow all steps and make reference as to why each step was done*.

13. These are your notes, but remember several points:
    A. Your classmates may need to copy them.
    B. You may need them in college as a reference.
    C. Make your notes so that a reasonable, intelligent person can follow them.

Taking the Test

1. If your test has multiple choice questions on it, check to see how the points are awarded. If all of the multiple choice questions are the same value, then don't spend a long time doing a hard problem when you could have done five easier questions in the same length of time. Save the hard questions until the end.

2. Estimate the answer to the question. Determine the **units** that the answer **MUST** have.

4. Show all of your work on the problems. Most teachers will give partial credit for problems that you have not finished or solved correctly. Make sure that your work is in the format that your teacher wants.

5. Check your answer with your estimate. Check to make sure that ALL of the units work out. Often problems can be done simply by unit analysis.

Problem Solving

Chemistry is a problem-solving class. Problem solving requires you to answer several questions:

- What am I asked to find?
- What information am I to use?
- Do I have all of the information to solve the problem?
- If not, are there other problems that I must do first to get me all of the information?
- What principles or laws apply?
- Have I seen other problems like this one?
- If so, what can I apply from them to solve this problem?
- How can I go about applying the information to solve the problem?
- Does my solution make sense?

You must be able to see beyond the surface of the problem. Look for the real meaning of the problem...find the "meat" of the problem. Many students just look at a problem and say, "I can't do this" but yet when they analyze the problem they can solve it.

The steps in solving a problem are:

1. Read the problem
   A. List out all information given – ***include the units***.
   B. List the unknown – what the problem wants you to find.

2. Draw a diagram of the problem

   1. Identify all compounds in the problem and write their formulas correctly.
   2. If it involves a chemical reaction, BALANCE THE REACTION first.
   3. Set up the problem the way that your teacher instructs you to. Each type of problem has its own method of diagramming the information.

3. Plan the problem.
   A. Write out the *basic equation(s)* that will be used to solve the problem.
   B. Solve the basic equation to find the *working equation(s)*, the equation that is solved for the unknown.
   C. Make sure that your formulas are appropriate, don't try to use the Ideal Gas Law on a solid.
   D. Think of the relationships involved in the problem. Are there ways to get missing information that you will need to solve the overall problem? Remember, if you are given distance and rate, you also have the time (after you solve for it).

4. Solve the problem. Put the numbers into the working equation, **WITH THE UNITS** and work out the problem. Use ***unit analysis*** to make sure that you arrived at an answer that is consistent with your work.

5. **SHOW ALL OF YOUR WORK**.

6. Ask yourself:
   A. Does this answer make sense?
   B. Are the units appropriate for the problem?

Final Exam Preparation

The first step in preparing for an exam is to organize your time. Write out a schedule listing which topic you are going to cover at which time. Remember, ***spend most of your time working problems***, most science tests are problem based and the more problems you do, the better prepared you will be. **You cannot cram for chemistry: it is a science in which one topic builds upon another**.

1. Organize your time and workspace
   A. Plan what you are going to be studying and when.
   B. Organize your study area to have all material needed and to minimize distractions.

2. Work with a study group to review notes, labs, and problems.
   A. Outline the material.
   B. Use concept maps for the material.
   C. Work problems together.
   D. Explain the meaning of new terminology to each other, or illustrate its use in a problem.

3. Quickly review the notes of the major topics the night before.

4. For your final exam:
   A. Start studying two weeks before the test date.
   B. Review your notes in order of presentation.
      1. Remember that science is cumulative and that topics covered the first day of class may be on the final.
      2. Rework example problems from the beginning of the year.
   C. Review your old homework and rework old problems.

Useful Websites:

There is a multitude of awesome chemistry resources available via the Internet. With the hundreds of tutorial websites out there, I feel confident that you will find adequate information on any topic within this packet. In addition, below are some general chemistry websites for you to check out this summer. You may want to bookmark these and use them as a reference throughout the year.

http://www.khanacademy.org/science/chemistry/#science/chemistry

http://www.brightstorm.com/science/chemistry/

http://www.chemmybear.com/stdycrds.html

http://www.chemteam.info/ChemTeamIndex.html

http://www.adriandingleschemistrypages.com/apquiz.html

http://www.webelements.com/

https://www.youtube.com/channel/UCj3EXpr5v35g3peVWnVLoew This one I like, Tyler DeWitt is great.

Recommended Study Books:

5 Steps to a 5, AP Chemistry

Cracking the AP Chemistry Exam, The Princeton Review

AP Chemistry, Barron's