AP Chemistry – Summer 2011

Welcome to AP Chemistry. Over the summer, you will be expected to read the first three chapters of *Chemistry: The Central Science (10th Edition)* and complete the assigned problems at the end of each chapter. **These problems will be collected on orientation day and count as a test grade worth 125 points.** You may type or handwrite your responses as long as they are legible and all your work is shown. The following document has been created to help refresh your memory about some topics, most of which you have already encountered in your prior chemistry course. It does not include every topic; rather, it is intended to supplement the material presented in the textbook. If you have any questions as you complete this assignment, please feel free to email me at bcrossen@bishopfeehan.com. I will check my email daily throughout the summer and will do my best to respond as quickly as possible. A word of warning: I have assigned a large number of problems to ensure that you are prepared for the AP exam, so do not wait until the last minute to complete them. Please let me know if you encounter any difficulties as you go along. I look forward to meeting all of you in August!

CHAPTER 1: 26 (1.5 pt), 30 (1.5 pt), 32 (2 pt), 34 a,b (1.5 pt), 50 (4 pt), 52 (1.5 pt), 74 (2 pt), 80 (1 pt)

CHAPTER 2: 16 a,b (2 pt), 18 (2 pt), 20 (1 pt), 24 (4 pt), 28 (1.5 pt), 30 (2 pt), 42 (3.5 pt), 48 (4.75 pt), 52 (1.5 pt), 54 (2.5 pt), 56 (2 pt), 60 a, c-g, i, j (8 pt), 62 (3.5 pt), 64 (3 pt), 66 (5 pt), 68 (4 pt), 70 a (.25 pt), 84 (4 pt), 90 (1.5 pt), 96 (3 pt)

CHAPTER 3: 14 (5 pt), 16 (2 pt), 18 (4 pt), 20 (5 pt), 28 (1 pt), 36 (7 pt), 42 (2 pt), 50 (6 pt), 58 (5 pt), 64 (6 pt), 74 (5 pt), 80 (4 pt)

CHAPTER 1: Introduction: Matter and Measurement

Units of Measurement

The **SI system** of units, which is the modern version of the metric system, is the primary system of units that we will use in this class.

Units in the SI system include **base units**, which are defined by standards, and **derived units**, which are defined in terms of the base units. The base units that we will use are:

Physical Quantity	Name of Unit	Abbreviation
Mass	Kilogram	kg
Length	Meter	m
Time	Second	s (or sec)
Temperature	Kelvin	K
Amount of substance	mole	mol

You should also be familiar with the SI prefixes:

Prefix	Abbreviation	Meaning	Example
Giga	G	10^9	1 Gm = $1 \bullet 10^9$ m
Mega	M	10^6	1 Mm = $1 \bullet 10^6$ m
Kilo	k	10^3	1 km = $1 \bullet 10^3$ m
Deci	d	10^{-1}	1 dm = .1 m
Centi	c	10^{-2}	1 cm = .01 m
Milli	m	10^{-3}	1 mm = .001 m
Micro	µ	10^{-6}	1 µm = $1 \bullet 10^{-6}$ m
Nano	n	10^{-9}	1 nm = $1 \bullet 10^{-9}$ m
Pico	p	10^{-12}	1 pm = $1 \bullet 10^{-12}$ m
femto	f	10^{-15}	1 fm = $1 \bullet 10^{-15}$ m

The prefixes in bold are the most commonly used ones in this class.

Temperature

$K = {}^{o}C + 273.15$

${}^{o}C = \frac{5}{9}({}^{o}F - 32)$ Hint: For a “mental math” approximation, subtract 32 from the Fahrenheit temperature and divide by 2.

Volume

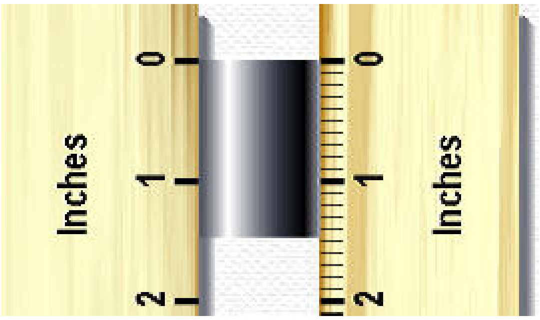
1 mL = 1 cm^3 = .001 L

Density

density = $\frac{mass}{volume}$ = g/cm^3 or g/mL

Significant Figures

Significant figures are one method that chemists can use to convey the uncertainty or error in a measurement. The accuracy of a measurement depends on the measurement scale of the instrument being used (5.1 m is a less accurate measurement than 5.12 m). The significant figures in any measurement are the digits that are known with certainty, plus one digit that is uncertain. When taking a measurement, write down the digits that can be read directly from the instrument, and obtain another digit by estimating the fraction of the smallest division of the instrument's scale.



For example, if you were measuring the lead pellet above using the left ruler (divided into 1 inch increments):

- You would start by writing 1 inch, because you are certain it is 1 inch, and then estimate another digit based on the fraction. So 1.4 inches would be a valid measurement. It has two significant figures, one that is known with certainty and one that is estimated.

If you were measuring the lead pellet above using the right ruler (divided into .1 inch increments):

- You would start by writing 1.4 inches, because you are certain it is 1.4 inches, and then estimate another digit based on the fraction. So 1.45 inches would be a valid measurement. It has three significant figures, two that are known with certainty and one that is estimated.

The rules for determining how many significant figures a written number has are:

1. All non-zero numbers are always significant.

2. Zeros at the beginning of a number are not significant. They merely locate the decimal point. For example, 0.0254 m has three significant figures (2, 5, 4)

3. Zeros between other sig figs are significant ("sig fig sandwich" rule). For example, 104.6 m has four significant figures (1, 0, 4, 6)

4. Zeros at the end of a number and to the right of the decimal point are significant. For example, 2705.0 m has five significant figures (2, 7, 0, 5, 0)

5. In whole numbers without a decimal point that end in one or more zeros (trailing zeros) – for example, 500 kg – the zeros may or may not be significant. In such cases, it is not clear which zeros serve only to locate the decimal point and which are actually part of the measurement. To avoid this, use scientific notation.

500 kg	1 sig fig
$5.0 \bullet 10^1$ kg	2 sig figs
$5.00 \bullet 10^2$ kg	3 sig figs

6. Counted or exactly defined quantities (conversion factors) have unlimited sig figs.

16 puppies	unlimited sig figs
2.54 cm = 1 inch	unlimited sig figs
60 minutes = 1 hour	unlimited sig figs

When performing calculations, use the following rules:

1. When multiplying or dividing quantities, leave as many **significant figures** (rounded) in the answer as there are in the quantity with the **least** number of significant figures.

2. When adding or subtracting quantities, leave the same number of **decimal places** (rounded) in the answer as there are in the quantity with the **least** number of decimal places.

3. For calculations involving multiple steps, use the "order of operations" rules to perform one calculation at a time.

Use these rules for rounding:

1. If the first digit to be dropped is less than 5, leave the preceding digit as is.

2. If the first digit to be dropped is 5 or greater, increase the preceding digit by one.

For example,

$$\underset{\text{(3 sig figs)}}{(6.12\text{ m})}\underset{\text{(2 sig figs)}}{(2.9\text{ m})} = 17.748\text{ m}^2 = \underset{\text{(round to 2 sig figs)}}{18\text{ m}^2}$$

$$\frac{(1216.550g - 162.55g)}{.0910mL} = \frac{1054.00g}{.0910mL} = 11582g/mL = 1.16 \bullet 10^4\ g/mL$$

(round numerator to 2 decimal places) (round to 3 sig figs)

Dimensional Analysis

In order to convert a measurement from one unit to another, multiply by the correct **conversion factor** – a statement of equivalence in the form of a ratio. Many conversion factors are listed on the back inside cover of your textbook.

To convert 14.6 km into feet:

$$14.6km \bullet \frac{1000m}{1km} \bullet \frac{100cm}{1m} \bullet \frac{1inch}{2.54cm} \bullet \frac{1foot}{12inches} = 47900 \text{ ft} = 4.79 \bullet 10^4 \text{ ft} \text{ (round to 3 sig figs)}$$

To convert 92.400 in^3 into km^3:

$$92.400in^3 \bullet \left(\frac{2.54cm}{1inch}\right)^3 \left(\frac{1m}{100cm}\right)^3 \left(\frac{1km}{1000m}\right)^3 = 1.5142 \bullet 10^{-12} \text{ km}^3 \text{ (round to 5 sig figs)}$$

CHAPTER 2: Atoms, Molecules, and Ions

Atomic Structure

Atomic Number

- The number of protons in the nucleus of an atom
- Unique to every atom of an element
- Written as a subscript to the left of the element's symbol

Mass Number

- The number of protons and neutrons in the nucleus of an atom
- Not unique to every atom of an element
- Written as a superscript to the left of the element's symbol

An atom of carbon with an atomic number of 6 and a mass number of 14 can be written as ${}^{14}_{6}C$ or carbon-14. Assuming neutrality, this atom would contain 6 protons, 6 electrons, and 8 (14 - 6) neutrons. Carbon-12 and carbon-14 are known as **isotopes** since they contain the same number of protons (6) but different numbers of neutrons (6 and 8, respectively).

Atomic Mass (AM)

- An average mass of the atoms of an element based on their natural abundance in nature

To calculate an element's atomic mass, multiply the percent of each isotope by its mass (in amu) and add:

$$\text{AM} = \underbrace{\left(\frac{abundance}{100}\right)(\text{mass in amu})}_{\text{ISOPTOPE 1}} + \underbrace{\left(\frac{abundance}{100}\right)(\text{mass in amu})}_{\text{ISOTOPE 2}} + \underbrace{\left(\frac{abundance}{100}\right)(\text{mass in amu})}_{\text{ISOTOPE 3}}$$

Round to the least number of decimal places in masses (adding masses).

$$\text{carbon} = \left(\frac{98.89\%}{100}\right)(12.000amu) + \left(\frac{1.11\%}{100}\right)(13.003amu) = 12.011 \text{ amu}$$

Notice that the atomic mass of carbon is much closer to the mass of carbon-12 since it is the most abundant isotope.

Ions and Ionic Compounds

You will be expected to memorize the names and formulas of the following polyatomic ions:

Ammonium	NH_4^+	Acetate	$C_2H_3O_2^-$
Chlorate	ClO_3^-	Chlorite	ClO_2^-
Cyanide	CN^-	Bicarbonate	HCO_3^-
Hydroxide	OH^-	Nitrate	NO_3^-
Nitrite	NO_2^-	Perchlorate	ClO_4^-
Permanganate	MnO_4^-	Carbonate	CO_3^{2-}
Sulfate	SO_4^{2-}	Sulfite	SO_3^{2-}
Phosphate	PO_4^{3-}		

Naming Acids

Use the following rules and tips to name acids (molecules that form H^+ ions when dissolved in water):

1. For acids that contain an anion that ends in "-ide," add "hydro" to the anion name and change "-ide" to "-ic." For example, you know that HCl is **hydro**chlor**ic** acid (from the chloride anion).

2. For acids that contain an anion that ends in "-ate," change the "-ate" to "-ic." For example, you know that H_2SO_4 is sulfur**ic** acid (from the sulfate anion).

3. For acids that contain an anion that ends in "-ite," change the "-ite" to "-ous." For example, $HClO_2$ is chlor**ous** acid (from the chlorite anion). I use the term "sp**ite us**" to remember this rule.

Naming Molecules

1. Write the name of the element furthest to the left on the periodic table first.

2. If both elements are from the same group, write the name of the one with the higher atomic number first.

3. Use the following prefixes to indicate the number of atoms of each element

Prefix	Mono-	Di-	Tri-	Tetra-	Penta-	Hexa-	Hepta-	Octa-	Nona-	Deca-
Meaning	1	2	3	4	5	6	7	8	9	10

4. Do not use the prefix "mono-" for the first element.

5. Drop the final “a” or “o” from the prefix if the element’s name begins with a vowel.

Naming Organic Compounds

Alkanes

- Organic compounds made only of carbon and hydrogen atoms joined by single bonds

When naming such organic compounds, the prefixes “meth-” (1), “eth-” (2), “prop-“ (3), and “but-“ (4) are used to indicate the number of carbon atoms. The prefixes used for molecules (see table on page 6) are used for compounds with 5 or more carbon atoms.

CH_4	methane
C_2H_6	ethane
C_3H_8	propane
C_4H_{10}	butane
C_5H_{12}	pentane
C_6H_{14}	hexane

Alcohols

- An alkane with a hydroxyl group (-OH) bonded to one or more carbon atoms

When naming alcohols, add “-ol” to the alkane from which it is derived.

CH_3OH	methanol
C_2H_5OH	ethanol
C_3H_7OH	propanol
C_4H_9OH	butanol
$C_5H_{11}OH$	pentanol
$C_6H_{13}OH$	hexanol

CHAPTER 3: Stoichiometry

Types of Chemical Reactions

1. Combination
   - 2 or more elements combine to form one product

     Magnesium + oxygen → magnesium oxide
     $2\,Mg + O_2 \rightarrow 2\,MgO$

     Iron + sulfur → iron (III) sulfide
     $2\,Fe + 3\,S \rightarrow Fe_2S_3$

     Carbon + oxygen → carbon monoxide
     $2\,C + O_2 \rightarrow 2\,CO$

2. Decomposition
   - One compound breaks down into 2 or more products

   Hydrogen peroxide → water + oxygen
   $2\,H_2O_2 \rightarrow 2\,H_2O + O_2$

   Water → hydrogen + oxygen
   $2\,H_2O \rightarrow 2\,H_2 + O_2$

   Calcium bromide → calcium + bromine
   $CaBr_2 \rightarrow Ca + Br_2$

3. Single-Replacement
   - One element replaces another element in a compound (like love triangle)

A) metal A (Angelina Jolie) vs. metal B (Jennifer Aniston) for nonmetal C (Brad Pitt)

$$\underset{\substack{\downarrow\\ \text{metal}}}{A} \quad + \quad \underset{\substack{\downarrow\\ \text{metal}}}{B}C \rightarrow B \quad + \quad AC$$

$$\underset{A}{2\,K} + \quad \underset{B}{Cu}(\underset{C}{NO_3})_2 \rightarrow \underset{B}{Cu} \underset{+}{+} \underset{AC}{2\,KNO_3}$$

No reaction if metal A is lower than B on activity series.

$Na + H_2O \rightarrow$

$2\ Na + 2\ HOH \rightarrow H_2 + 2\ NaOH$
Hint: Think of H_2O as an ionic compound made of H^+ and OH^-.

B) halogen A vs. halogen C

$$\underset{\substack{\downarrow \\ \text{halogen}}}{A} \quad + \quad B\underset{\substack{\downarrow \\ \text{halogen}}}{C} \rightarrow C \quad + \quad BA$$

$Br_2 + 2\ NaI \rightarrow I_2 + 2\ NaBr$
A + BC C + BA

No reaction if halogen A is lower than C in group 7.

4. Double-Replacement
   - The cation and anion of one ionic compound switch partners with the cation and anion in another ionic compound

$AB + CD \rightarrow AD + CB$

$K_2O + BaCl_2 \rightarrow 2\ KCl + BaO$

$3\ Li_2CO_3 + Mg_3N_2 \rightarrow 2\ Li_3N + 3\ MgCO_3$

$Ca(OH)_2 + 2\ NH_4Cl \rightarrow CaCl_2 + 2\ NH_4OH$

$Fe_2S_3 + 3\ ZnCl_2 \rightarrow 2\ FeCl_3 + 3\ ZnS$

5. Combustion
   - A compound with hydrogen and carbon (and sometimes oxygen) reacts with oxygen to form carbon dioxide and water (exothermic reaction)

$2\ C_3H_8O + 9\ O_2 \rightarrow 6\ CO_2 + 8\ H_2O$
$2\ C_8H_{18} + 25\ O_2 \rightarrow 16\ CO_2 + 18\ H_2O$
$CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$

Molar Mass

When calculating a substance's molar mass, round to the same number of decimal places as the mass with the least number of decimal places (since you are adding these masses). Use the atomic masses listed on the front inside cover of your textbook:

CO_2	12.0107 g/mol + 2(15.9994 g/mol) = 44.0095 g/mol
H_2S	2(1.00794 g/mol) + 32.065 g/mol = 34.081 g/mol

Empirical Formula

- Shows the lowest whole-number ratio of atoms in a compound

A new compound contains 25.9 % nitrogen and 74.1 % oxygen by mass. What is the empirical formula of this compound?

1. Assume 100 g of the sample. Then, convert the grams of each element to moles

$$25.9\% = 25.9 \text{ g N} \bullet \frac{mol}{14.0067g} = 1.85 \text{ mole N} \qquad 74.1\% = 74.1 \text{ g O} \bullet \frac{mol}{15.9994g} = 4.63 \text{ mole O}$$

2. Find the ratio of elements by dividing by the lowest number of moles. If the ratios are not whole numbers, multiply each ratio by the smallest integer value that will make them whole numbers.

$$\frac{1.85}{1.85} = 1.00 \text{ N} \qquad 2(1.00) = 2.00 \text{ N} \qquad \frac{4.63}{1.85} = 2.50 \text{ O} \qquad 2(2.50) = 5.00 \text{ O}$$

3. Use the ratios calculated above as the subscripts in the formula.

   Empirical Formula = N_2O_5

If the molar mass of the compound is 216.018 g/mol, what is its **molecular formula**?

1. Find the molar mass of the empirical formula.

   MM of N_2O_5 = 108.0104 g/mol

2. Divide the molecular MM by the empirical MM.

   $216.018/108.0104 = 1.99997 \approx 2$

3. Multiply the subscripts in the empirical formula by the answer to step 2.

   $N_2O_5 \rightarrow N_4O_{10}$

Mole-Mole Conversions

$4\ Al + 3O_2 \rightarrow 2Al_2O_3$

How many moles of Al_2O_3 can be created when 14.8 moles of Al reacts with O_2?

$$14.8 \text{ mol Al} \bullet \frac{2molesAl_2O_3}{4molesAl} = 7.40 \text{ moles } Al_2O_3$$

Mass-Mass Conversions

$4\ Na + O_2 \rightarrow 2\ Na_2O$

How many grams of Na_2O can be created when 14.07 g of Na reacts with O_2?

$$14.07 \text{ g Na} \bullet \frac{1moleNa}{22.990g} \bullet \frac{2molesNa_2O}{4molesNa} \bullet \frac{61.979g}{1moleNa_2O} = 18.97 \text{ g } Na_2O$$

Limiting Reagents and Theoretical Yield

$2\ Cu + S \rightarrow Cu_2S$

What is the limiting reagent when 80.0 g of Cu reacts with 25.0 g of S?

1. Calculate the "moles present" for both reactants.

$$80.0 \text{ g Cu} \bullet \frac{1molCu}{63.546g} = 1.26 \text{ mol Cu present} \qquad 25.0 \text{ g S} \bullet \frac{1molS}{32.065g} = .780 \text{ mol S present}$$

2. Choose the "moles present" for one reactant and calculate the "moles needed" for the other reactant.

$$1.26 \text{ mol Cu } \frac{1molS}{2molCu} = .630 \text{ mol S needed} \quad \Big| \quad .780 \text{ mol S } \frac{2molCu}{1molS} = 1.56 \text{ mol Cu needed}$$

3. Compare the "moles present" (step 1) and "moles needed" (step 2) for the reactant you chose.
    - If "moles present" is greater than "moles needed," that reactant is the **excess reagent**.
    - If "moles present" is less than "moles needed," that reactant is the **limiting reagent**.

.780 mol S present > .630 mol S needed
S = EXCESS REAGENT
Cu = LIMITING REAGENT

1.26 mol Cu present < 1.56 mol Cu needed
Cu = LIMITING REAGENT
S = EXCESS REAGENT

How many grams of the excess reagent are **left over** after the reaction is complete?

.780 mol S present - .630 mol S needed = .150 mol S left over

$$.150 \text{ mol} \bullet \frac{32.065 g}{1 mol} = 4.81 \text{ g S left over}$$

Using the information above, what is the **theoretical yield** of Cu_2S?

$$1.26 \text{ mol Cu} \bullet \frac{1 mol Cu_2S}{2 mol Cu} \bullet \frac{159.157 g}{1 mol Cu_2S} = 1.00 \bullet 10^2 \text{ g } Cu_2S$$

If the actual yield of Cu_2S is 92 g, what is the **percent yield** of Cu_2S?

$$\text{PY} = \frac{AY}{TY} \bullet 100 = \frac{92 g}{1.00 \bullet 10^2 g} \bullet 100 = 92\%$$