2019-2020 AP Chemistry

Congratulations on making the decision to take AP Chemistry! This course will move at a fast pace and cover a substantial amount of material, starting with the first day of school. The primary goal of this course is to earn college credit by earning a grade of C or higher for dual enrollment credit or by passing the AP Chemistry exam with a score of 3 or higher in May 2020 (most colleges will not give credit for a score of 1 or 2).

So that we can spend more time on topics new to you in AP Chemistry, you are expected to be familiar answering questions and solving problems using the content covered in your first year chemistry course. The attached **review assignment** covers first-year chemistry topics that will not be taught in AP chemistry. Videos for each of these topics that I have created will be available through YouTube on my channel COACHSCOTTSCHEMISTRY. I will put the answers to these assignments on my webpage. You will have an opportunity to ask questions on this assignment during the first three class periods. The assignment will be collected prior to your in-class test on these topics during the fourth class period for a grade. *It is up to you whether or not you start work on this assignment before the school year*; if it has been a year since you took your first chemistry course you are strongly encouraged to begin work on this assignment at least the week before school starts.

Copies of the periodic table and the metric prefixes you will be using in AP Chemistry are included in this assignment. Please note that this periodic table does not include element names. Charges of monatomic ions and key polyatomic ions that need to be memorized by the first test are also included. You are encouraged to make flashcards or use the Quizlet ions card deck to begin learning these ions.

If you have any questions during the summer, you are welcome to contact me via email at scottal1@scsk12.org. I wish each of you a restful and enjoyable summer, and I look forward to seeing you next school year!

Coach Scott

TOPIC 0A: Chemistry, Scientific Method and Chemical & Physical Change What is chemistry?

Chemistry can be described as the science that deals with matter, and the changes that matter undergoes. It is sometimes called the ***central science*** because so many naturally occurring phenomena involve chemistry and chemical change.

Scientific problem solving

Scientific (logical) problem solving involves three steps;

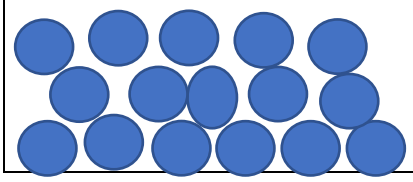
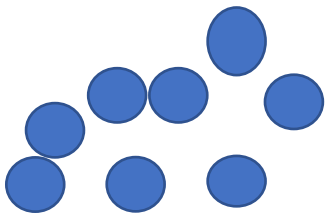
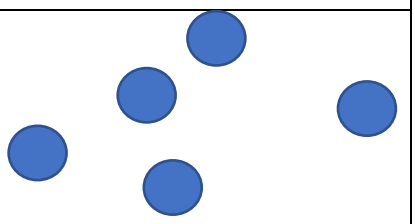
1. State the problem and make observations. Observations can be *quantitative* (those involving numbers or measurement) or *qualitative* (those not involving numbers).
2. Formulate a possible explanation (this is known as a *hypothesis*).
3. Perform experiments to test the hypothesis. The results and observations from these experiments lead to the modification of the hypothesis and therefore further experiments.

Eventually, after several experiments, the hypothesis may graduate to become a *theory*. A theory gives a universally accepted explanation of the problem. Of course, theories should be constantly challenged and may be refined as and when new data and new scientific evidence comes to light.

Theories are different to *laws*. Laws state what general behavior is observed to occur naturally. For example, the *law of conservation of mass* exists since it has been consistently observed that during all chemical changes mass remains unchanged (i.e., it is neither created nor destroyed).

States of matter and particle representations

All matter has two distinct characteristics. It has mass and it occupies space. Properties associated with the three states of matter, and the behaviors of the particles that make up each, are summarized below.

SOLIDS	GASES	LIQUIDS
Have a definite shape and definite volume. The particles in a solid are packed tightly together and only vibrate relatively gently around fixed positions.	Have no shape of their own but take the shape of their container. A liquid has a definite volume. The particles in a liquid are free to move around one another.	Have neither a definite shape nor a definite volume. The particles in a gas spread apart filling all the space of the container available to them and interactions between the particles are considered to be negligible
The circles in the diagrams below represent the relative positions and movements of the particles in the three states of matter. Expect to see many such *particulate representations* during the AP course		
		

Physical and chemical changes and properties

All matter exhibits physical and chemical properties by which it can be classified. Examples of *physical properties* are color, odor, density, hardness, solubility, melting point, and boiling point.

Chemical properties are those exhibited when a substance reacts with other substances. Examples of chemical properties are reactions with acids and bases, oxidation and reduction (REDOX) and a huge number of other chemical reactions. Changes in which the physical or chemical properties of a substance are altered are considered physical or chemical changes, respectively.

Physical change

If some aspect of the physical state of matter is altered, but the chemical composition remains the same, then the change is considered to be a physical change. The most common physical changes are changes of state. These are summarized below

Solid	→	Liquid	Melting
Liquid	→	Gas	Boiling
Gas	→	Liquid	Condensing
Solid	→	Gas	Sublimation
Gas	→	Solid	Reverse sublimation or deposition
Liquid	→	Solid	Freezing

In solids, the particles have relatively little energy and vibrate around fixed positions. If a solid is heated, the particles gain energy, move around move, and eventually gain enough energy to break away from their fixed positions and form a liquid. Continued heating leads to the liquid particles gaining sufficient energy to break away from one another and form a gas. In a gas the particles move freely and with relatively large amounts of energy

Chemical change

In a chemical change, which is often called a chemical reaction, the atoms of a substance are rearranged to form new substances. A chemical change requires that the new substance or substances formed have a different chemical composition to the original substance or substances. Chemical changes are often accompanied by observable changes such as color changes and energy changes.

There is a very important distinction to be made between these two types of change that you will encounter in UNIT 2. More on this later, but for now, note the following.

- During physical changes, the intermolecular forces (the forces *between* particles) are disrupted, e.g., boiling water separates one water molecule (H_2O) from another water molecule but does not break any, individual water molecule apart.

- During chemical changes the intra forces (the forces *within* substances) are disrupted, e.g., during the electrolysis of water, one water molecule (H_2O) splits up to form O and H atoms. Individual water molecules do break apart.

TOPIC 0B: Measurement

Measurements

Measurements, and subsequently calculations applied to those measurements, allow the determination of some of the quantitative properties of a substance; for example, mass and density.

Scientific notation

Measurements and calculations in chemistry often require the use of very large or very small numbers. In order to make handling them easier, such numbers can be expressed using *scientific notation*. All numbers expressed in this manner are represented by a number between 1 and 10 which is then multiplied by 10, raised to a particular power.

The number of places the decimal point has moved determines the power of 10. If the decimal point has moved to the left then the power is positive, if it has moved to the right then it is negative.

For example, the number 42000.0 is converted to scientific notation by using the number 4.2. In the process the decimal point has moved four places to the *left*, so the power of 10 used is +4.

$$42000.0 = 4.2 \times 10^{4}$$

The number 0.00012 is converted to scientific notation by using the number 1.2. In the process he decimal point has moved four places to the *right*, so the power of 10 used is -4.

$$0.00012 = 1.2 \times 10^{-4}$$

Task 0Bi

1 Convert the following numbers to scientific notation.

(a) 24500

(b) 356

(c) 0.000985

(d) 0.222

(e) 12200

2. Convert the following scientific notation numbers to non-scientific notation numbers.

(a) 4.2×10^{3}

(b) 2.15×10^{-4}

(c) 3.14×10^{-6}

(d) 9.22×10^{5}

(e) 9.57×10^{2}

SI units

Units tell us the scale that is being used for measurement. Prefixes are used to make writing very large or small numbers easier. Common SI (*System International*) units and prefixes are given below.

Base Quantity	**Name of Unit**	**Symbol**
Mass	Kilogram	kg
Length	Meter	m
Time	Second	s
Amount of substance	Mole	Mol
Temperature	Kelvin	K

Prefix	Symbol	Meaning
Giga	G	10^{9}
Mega	M	10^{6}
Kilo	k	10^{3}
Deci	d	10^{-1}
Centi	c	10^{-2}
Milli	m	10^{-3}
Micro	μ	10^{-6}
Nano	n	10^{-9}
Pico	p	10^{-12}

Converting units and dimensional analysis (the factor label method)

One unit can be converted to another unit by using a conversion factor. Application of the simple formula below will allow the conversion of one unit to another. This method of converting between units is called *dimensional analysis* or the *factor-label method*.

(unit a) (conversion factor) = unit b

The conversion factor is derived from the equivalence statement of the two units. For example, in the equivalence of 1.00 inch = 2.54 cm, the conversion factor will either be,

$$\boxed{\frac{2.54\ cm}{1.00\ inch}\ or\ \frac{1.00\ inch}{2.54\ cm}}$$

The correct choice is the one that allows the cancellation of the unwanted units. For example, to convert 9.00 inches to cm, perform the following calculation

$$\boxed{\frac{9.00\ \cancel{inch}}{}\ x\ \frac{1.00\ inch}{2.54\ \cancel{cm}} = 1.97\ inches}$$

Task 0Bii

1. Convert the following quantities from one unit to another, using the following equivalence statements; 1.000 m = 1.094 yd, 1.000 mile = 1760 yd, 1.000 kg = 2.205 lbs

(a) 30 m to miles

(b) 1500 yd to miles

(c) 206 miles to m

(d) 34kg to lbs

(e) 34lb to kg

2. In each case below, which is the larger quantity?

   (a) A distance of 3.00 miles or 3000. m.

   (b) A mass of 10.0 kg or 25 lbs.

Temperature

There are three scales of temperature that you may come across in your study of chemistry. They are Celsius (^{o}C), Fahrenheit (^{o}F) and Kelvin (K). The following conversion factors will be useful.

Temperature Conversion factors	
Celsius to Kelvin	T in K =T in ^{o}C + 273
Kelvin to Celsius	T in ^{o}C =T in K - 273
Celsius to Fahrenheit	T in ^{o}F = (1.8 (T in ^{o}C)) + 32
Fahrenheit to Celsius	T in ^{o}C = $\frac{(T\ in\ \deg F - 32)}{1.8}$

Task 0Biii

1. Convert the following temperatures from one unit to the other.

   (a) 263 K to ^{o}F

   (b) 38 K to ^{o}F

   (c) 13° F to ^{o}C

   (d) 1390°C to K

   (e) 3000°C to ^{o}F

2. When discussing a change in temperature, why will it not matter if the change is recorded in Celsius or Kelvin?

Derived units

All other units can be derived from base quantities. One such unit that is very important in chemistry is volume. Volume has the unit, length$^{3.}$ Common units for volume are liters (L) or milliliters (mL).

$$1.000 \text{ mL} = 1.000 \text{ cm}^3$$

And

$$1.000 \text{ L} = 1000. \text{ mL} = 1000. \text{ cm}^3 = 1.000 \text{ dm}^3$$

Density is the ratio of the mass to volume.

$$\boxed{density = \frac{mass}{volume}}$$

This relationship is particularly useful when dealing with liquids in chemistry. Liquids are most conveniently measured by pouring them into, say, a graduated cylinder. The graduated cylinder records a volume, not a mass. In order to calculate the mass of a known volume of a liquid (assuming the density is known) the relationship below can be applied.

$$\text{Mass} = (\text{density})(\text{volume})$$

Assuming that density has the units of g/L, volume has units of L, and by using dimensional analysis, it can be seen that the resultant unit for mass in this case is g.

$$\boxed{\frac{g}{\cancel{L}} x \cancel{L} = g}$$

Uncertainty, significant figures and rounding

When reading the scale on a piece of laboratory equipment such as a graduated cylinder or a buret, there is always a degree of uncertainty in the recorded measurement. The reading will often fall between two divisions on the scale and an estimate must be made in order to record the final digit. This estimated final digit is said to be *uncertain* and is reflected in the recording of the numbers by using +/-. All of the digits that can be recorded with certainty are said to be *certain*. The certain and the uncertain numbers taken together are called *significant figures*

Determining the number of significant figures present in a number

1. Any non-zero integers are always counted as significant figures.
2. Leading zeros are those that precede all of the non-zero digits and are never counted as significant figures.
3. Captive zeros are those that fall between non-zero digits and are always counted as significant figures.
4. Trailing zeros are those at the end of a number and are only significant if the number is written with a decimal point.
5. Exact numbers have an unlimited number of significant figures. (Exact numbers are those which are as a result of counting e.g., 3 apples or by definition e.g., 1.000 kg = 2.205 lb).
6. In scientific notation the 10^x part of the number is never counted as significant.

Determining the correct number of significant figures to be shown as the result of a calculation

1. When multiplying or dividing. Limit the answer to the same number of *significant figures* that appear *in the original data with the fewest number of significant figures*.

2. When adding or subtracting. Limit the answer to the same number of *decimal places* that appear *in the original data with the fewest number of decimal places*.

i.e., don't record a greater degree of significant figures or decimal places in the calculated answer than the weakest data will allow.

Rounding

Calculators will often present answers to calculations with many more figures than the significant ones. As a result many of the figures shown are meaningless, and the answer, before it is presented, needs to be rounded. In a multi-step calculation it is possible to leave the rounding until the end i.e., leave all numbers on the calculator in the intermediate steps, or round to the correct number of figures in each step, or round to an extra figure in each intermediate step and then round to the correct number of significant figures at the end of the calculation. In most cases in the AP chemistry course you will leave numbers on the calculator and round at the end.

Whichever method is being employed, use the simple rule that if the digit directly to the right of the final significant figure is less that 5 then the preceding digit stays the same, if it is equal to or greater than 5 then the preceding digit should be increased by one.

Task 0Biv

1. Determine the number of significant figures in the following numbers.

   (a) 250.7

   (b) 0.00077

   (c) 1024

   (d) 4.7×10^{-5}

   (e) 34000000

   (f) 1003

2. Use a calculator to carry out the following calculations and record the answer to the correct number of significant figures.

   (a) (34.5) (23.46)

   (b) 123/3

   (c) (2.61×10^{-1}) (356)

   (d) 21.78 + 45.86

   (e) 23.888897 - 11.2

   (f) 6-3.0

Accuracy and precision

Accuracy relates to how close the measured value is to the actual value of the quantity. *Precision* refers to how close two or more measurements of the same quantity are to one another.

Task 0Bv

1. Consider three sets of data that have been recorded after measuring a piece of wood that is exactly 6.000 m long.

	SET X	SET Y	SET Z
	5.864 m	6.002 m	5.872 m
	5.878 m	6.004 m	5.868 m
Average Length	5.871 m	6.003 m	5.870 m

   (a) Which set of data is the most accurate?

   (b) Which set of data is the most precise?

Percentage error

The data that are derived in experiments will often differ from the accepted, published, actual value. When this occurs, a common way of expressing accuracy is *percentage error*.

$$Percentage\ Error = \left[\frac{(Actual\ Value - Calculated\ Value)}{Actual\ Value}\right] x\ 100$$

TOPIC 0C: Atomic Theory

Brief history of atomic theory

Circa. 400-5 BC. Greek philosopher Democritus proposes the idea of matter being made up of small, indivisible particles (*atomos*).

Late 18th Century. Lavoisier proposes the Law of conservation of mass and Proust proposes the Law of constant composition.

Early 19th Century. Using the previously unconnected ideas above, John Dalton formulates his Atomic Theory.

Dalton's atomic theory

1. Elements are made from tiny particles called atoms.
2. All atoms of a given element are identical (N.B., see isotopes).
3. The atoms of a given element are different to those of any other element.
4. Atoms of different elements combine to form compounds. A given compound always has the same relative numbers and types of atoms. (Law of constant composition).
5. Atoms cannot be created or destroyed in a chemical reaction they are simply rearranged to form new compounds. (Law of conservation of mass).

Structure of the atom and the periodic table

Several experiments were being carried out in the 19th and 20th centuries that began to identify the sub-atomic particles that make up the atom. A summary of those experiments is given below

Scientist	Experiment	Knowledge gained	Relating to
Crookes	Cathode Ray Tube	Negative particles of some kind exist	Electron
J.J. Thomson	Cathode Ray Deflection	Mass/charge ratio of the electron determined	Electron
Millikan	Oil Drop Experiment	Charge on the electron	Electron
Rutherford, Marsden, and Geiger	Gold Foil Experiment	Nucleus present in atom	The nucleus of an atom and the proton

In the first part of the 20th Century, Bohr built upon Rutherford's idea by introducing quantum theory to the *Solar System Model* and proposed the idea that the atom was made up of a nucleus containing protons, that was being orbited by electrons, *but only in specific, allowed orbits*. Schrödinger subsequently expanded upon Bohr's model, in order to incorporate the wave nature of the electrons. Once Chadwick's discovered the neutron in 1932, the modern picture of the atom *in its simplest form* was complete.

Particle	Charge	Mass in atomic mass units (amu)	Position in atom
PROTON	+1	1	Nucleus
NEUTRON	0	1	Nucleus
ELECTRON	-1	$\frac{1}{1836}$	Outside of the nucleus

The atomic numbers (in the periodic table shown above the element symbol and sometimes referred to as Z) and mass numbers (in the periodic table below shown below the symbol and sometimes referred to as A) have specific meanings.

Atomic number =the number of protons in the nucleus of one atom in the element.

Since all atoms are neutral it also tells us the number of electrons surrounding the nucleus.

N.B., when atoms lose or gain electrons the proton and electron numbers become unbalanced and the atoms become charged particles, i.e., they are no longer neutral. These charged particles are called *ions*. A negative ion is formed when an atom gains electrons to possess a greater number of electrons than protons, and is called an *anion*. A positive ion is formed when an atom loses electrons to possess a fewer number of electrons than protons and is called a c*ation*.

Mass Number = the number of protons + the number of neutrons

Mnemonic **APE MAN**

Atomic Number = **P**rotons = **E**lectrons

Mass number = **P**rotons + **N**eutrons

Task 0Ci

1. Determine the number of protons, electrons and neutrons in,

   (a) $^{210}Pb_{82}$

   (b) $^{34}S_{16}$

2. Using your periodic table, determine how many elements within the first 20, have atoms with;

   (a) The same numbers of protons and electrons

   (b) The same numbers of protons and neutrons

TOPIC 0D: Nomenclature

Nomenclature

Nomenclature is the language of chemistry, and a grasp of it is essential to studying the subject.

Symbols

Each element has a symbol displayed on the periodic table. Some elements have a symbol that is a single letter while others have a symbol made up of two letters. It is important when writing the two letter symbols to ensure that you use a lower case letter for the second letter. This may sound trivial but is very important, for example, Co (cobalt), a metal element, is not the same as CO (carbon monoxide), a gaseous compound made from carbon (C) and oxygen (O).

Binary compounds of metals and non-metals (ionic compounds)

Binary compounds are those formed between only two elements. In compounds where one is a metal and one a non-metal an *ionic* compound is formed. An ion is a charged particle and ionic formulae and names can be determined by considering the charge on the ions. To find the formula of an ionic compound the positive and negative charges must be balanced, i.e., there must be no net charge.

To name a binary compound of a metal and a non-metal, the unmodified name of the positive ion is written first followed by the root of the negative ion with the ending modified to -ide. For example, NaCl is sodium chloride.

A few common ions, their charges and formulae are listed below.

Negative ions (ANIONS)			**Positive Ions (CATIONS)**		
Name	**Charge**	**Symbol**	**Name**	**Charge**	**Symbol**
Bromide	-1	Br^{-}	Aluminum	+3	Al^{+3}
Chloride	-1	Cl^{-}	Barium	+2	Ba^{+2}
Fluoride	-1	F^{-}	Calcium	+2	Ca^{+2}
Hydride	-1	H^{-}	Copper (I)	+1	Cu^{+}
Iodide	-1	I^{-}	Copper (II)	+2	Cu^{+2}
Nitride	-3	N^{-3}	Hydrogen	+1	H^{+}
Oxide	-2	O^{-2}	Iron (II)	+2	Fe^{+2}
Phosphide	-3	P^{-3}	Iron (III)	+3	Fe^{+3}
Sulfide	-2	S^{-2}	Lead (II)	+2	Pb^{+2}
			Lead (IV)	+4	Pb^{+4}
			Lithium	+1	Li^{+}
			Magnesium	+2	Mg^{+2}
			Manganese (II)	+2	Mn^{+2}
			Nickel (II)	+2	Ni^{+2}
			Potassium	+1	K^{+}
			Silver	+1	Ag^{+}
			Sodium	+1	Na^{+}

			Strontium	+2	Sr^{+2}
			Tin (II)	+2	Sn^{+2}
			Tin (IV)	+4	Sn^{+4}
			Zinc	+2	Zn^{+2}

Most transition metal ions (and a few other metal ions) include a Roman numeral after the name, for example, copper (II). These metals form ions with varying charges, and the Roman numeral identifies the charge in each case. Elements that commonly form an ion with only a single charge for example, sodium, do not have Roman numerals associated with them.

Task 0Di

1. Name these binary compounds.

   (a) NaCl

   (b) SrO

   (c) AlN

   (d) $BaCl_2$

   (e) K_2O (

   f) CuO

   (g) Cu_2O

2. Convert these names to formulae

   (a) Magnesium nitride

   (b) Barium bromide

   (c) Aluminum phosphide

   (d) Potassium iodide

   (e) Lithium chloride

   (f) Sodium fluoride

   (g) Tin (IV) bromide

Binary acids

Acids will be discussed at great length later in the course, but for the purposes of nomenclature, an acid can be defined as a compound that produces hydrogen ions (H^+) when it is dissolved in water, and the formulae of acids start with 'H'. *Binary acids* are formed when hydrogen ions combine with monatomic anions.

To name a binary acid use the prefix 'hydro' followed by the other non-metal name modified to an –ic ending. Then add the word 'acid'. For example, HCl is hydrochloric acid.

Polyatomic ions

Polyatomic ions are those where more than one element are combined together to create a species with a charge. Some of these ions can be named systematically, others names must be learned. Some common polyatomic ions, their charges and formulae are listed below.

Common Polyatomic Ions		
Name	**Charge**	**Formula**
Ammonium	+1	NH_4^+
Carbonate	-2	CO_3^{2-}
Chromate (VI)	-2	CrO_4^{2-}
Dichromate (VI)	-2	$Cr_2O_7^{2-}$
Ethanedioate	-2	$C_2O_4^{2-}$
Hydrogen carbonate (bicarbonate)	-1	HCO_3^-
Hydrogen Sulfate	-1	HSO_4^-
Hydroxide	-1	OH^-
Manganate (VII)(permanganate)	-1	MnO_4^-
Nitrate	-1	NO_3^-
Nitrite	-1	NO_2^-
Phosphate	-3	PO_4^{3-}
Sulfate	-2	SO_4^{2-}
Sulfite	-2	SO_3^{2-}

Polyatomic anions where oxygen is combined with another non-metal are called oxoanions and can be named systematically. In these oxoanions certain non-metals (Cl, N, P and S) form a series of oxoanions containing different numbers of oxygen atoms. Their names are related to the number of oxygen atoms present and are based upon the system below.

Name	Number of oxygen atoms
Hypo (element) ite	1
(element) ite	2
(element) ate	3
Per (element) ate	4

Where there are only two members in such a series the endings are –ite and –ate. For example, sulfite (SO_3^{2-}) and sulfate (SO_4^{2-}). When there are four members in the series the hypo- and per- prefixes are used additionally.

Some oxoanions contain hydrogen and are named accordingly, for example, HPO_4^{2-}, hydrogen phosphate. The prefix thio- means that a sulfur atom has replaced an atom of oxygen in an anion.

To name an ionic compound that contains a polyatomic ion, the unmodified name of the positive ion is written first followed by unmodified name of the negative ion. For example, K_2CO_3 is potassium carbonate.

Oxoacids

Oxoacids are formed when hydrogen ions combine with polyatomic oxoanions. This gives a combination of hydrogen, oxygen and another non-metal.

To name an oxoacid use the name of the oxoanion and replace the -ite ending with –ous or the - ate ending with -ic. Then add the word 'acid'. For example, H_2SO_4 is sulfuric acid.

To illustrate the names of these oxoanions and oxoacids consider the following example using chlorine as the non-metal.

Formula and name of oxoacid		Formula and name of corresponding oxoanion	
$HClO$	Hypochlorous acid	ClO^-	Hypochlorite
$HClO_2$	Chlorous acid	ClO_2^-	Chlorite
$HClO_3$	Chloric acid	ClO_3^-	Chlorate
$HClO_4$	Perchloric acid	ClO_4^-	Perchlorate

Task 0Dii

1. What are the formulae for the following ionic compounds?

   (a) Ammonium nitrate

   (b) Copper (II) bromide

   (c) Copper (I) bromide

   (d) Zinc hydrogen sulfate

   (e) Aluminum sulfate

   (f) Sodium perchlorate

   (g) Copper (II) iodite

2. Convert the following formulae to names.

(a) $NaNO_3$

(b) $KMnO_4$

(c) CaC_2O_4

(d) $CuSO_4$

(e) Cu_2SO_4

(f) KNO_2

(g) $LiClO_4$

Binary compounds of two non-metals (molecular compounds)
If the two elements in a binary compound are non-metals, then the compound is *molecular*.

To name a molecular compound of two non-metals, the unmodified name of the first element is followed by the root of the second element with ending modified to -ide. In order to distinguish between several different compounds with the same elements present use the prefixes mono, di, tri, tetra, penta and hexa to represent one, two, three, four, five and six atoms of the element respectively. For example, SO_2 is sulfur dioxide.

Some other examples are given below.

Formula	Name
BCl_3	Boron trichloride
CCl_4	Carbon tetrachloride
CO	Carbon monoxide
CO_2	Carbon dioxide
NO	Nitrogen monoxide
NO_2	Nitrogen deoxide

Note that the prefix mono is only applied to the second element present in such compounds, if the prefix ends with ‘a’ or ‘o’, and the element name begins with ‘a’ or ‘o’, then the final vowel of the prefix is often omitted.

Some compounds have trivial names that have come to supersede their systematic names, for example, H_2O is usually ‘water’, not dihydrogen monoxide!

Task 0Diii

1. Write formula or names for the following molecular compounds.

    (a) Dinitrogen tetroxide

    (b) Phosphorous pentachloride

    (c) Iodine trifluoride

    (d) Nitrogen dioxide

    (e) Dihydrogen monoxide

2. Convert the following formulae to names.

    (a) N_2O_5

    (b) PCl_3

    (c) SF_6 (

    d) H_2O

    (e) Cl_2O

Hydrates

Hydrates are ionic formula units with water molecules associated with them. The water molecules are incorporated into the solid structure of the ions. Strong heating can generally drive off the water in these salts. Once the water has been removed the salts are said to be anhydrous (without water).

To name a hydrate use the normal name of the ionic compound followed by the term 'hydrate' with an appropriate prefix to show the number of water molecules per ionic formula unit. For example, $CuSO_4.5H_2O$ is copper (II) sulfate pentahydrate.

AP Chemistry Ions

Monatomic Cations	Monatomic Anions	Polyatomic Cations	Polyatomic Anions
Group 1 (including H) H^{+1}, hydrogen Li^{+1}, lithium Na^{+1}, sodium K^{+1}, potassium Cs^{+1}, cesium Group 2 Be^{+2}, beryllium Mg^{+2}, magnesium Ca^{+2}, calcium Sr^{+2}, strontium Ba^{+2}, barium Group 13 Al^{+3}, aluminum Transition and Heavier Metals Cr^{+2}, chromium (II) Cr^{+3}, chromium (III) Mn^{+2}, manganese (II) Mn^{+4}, manganese (IV) Mn^{+7}, manganese (VII) Cu^{+1}, copper (I) Cu^{+2}, copper (II) Fe^{+2}, iron (II) Fe^{+3}, iron (III) Pb^{+2}, lead (II) Pb^{+4}, lead (IV) Hg^{+2}, mercury (II) Ni^{+2}, nickel (II) Ni^{+3}, nickel (III) Sn^{+2}, tin (II) Sn^{+4}, tin (IV) Ag^{+1}, silver Zn^{+2}, zinc	Group 17 and H H^{-1}, hydride F^{-1}, fluoride Cl^{-1}, chloride Br^{-1}, bromide I^{-1}, iodide Group 16 O^{-2}, oxide S^{-2}, sulfide Group 15 N^{-3}, nitride P^{-3}, phosphide	Ammonium, NH_4^{+1} Mercury (I), Hg_2^{+2}	Acetate, $C_2H_3O_2^{-1}$ Bicarbonate (hydrogen carbonate), HCO_3^{-1} Carbonate, CO_3^{-2} Perchlorate, ClO_4^{-1} Chlorate, ClO_3^{-1} Chlorite, ClO_2^{-1} Hypochlorite, ClO^{-1} Permanganate, MnO_4^{-1} Cyanide, CN^{-1} Hydroxide, OH^{-1} Peroxide, O_2^{-2} Nitrate, NO_3^{-1} Nitrite, NO_2^{-1} Chromate, CrO_4^{-2} Dichromate, $Cr_2O_7^{-2}$ Sulfate, SO_4^{-2} Sulfite, SO_3^{-2} Phosphate, PO_4^{-3} Phosphite, PO_3^{-3}

***Note: Transition metals are named with Roman numerals to indicate their oxidation state (charge) <u>if</u> they have multiple oxidation states. Silver and zinc are the only transition metals on this list that have a single oxidation state and therefore are not named with roman numerals. As long as you know which transition metals need Roman numerals, individual charges of these metals do not need to be memorized.

INFORMATION IN THE TABLE BELOW AND IN THE TABLES ON PAGES 3-5 MAY BE USEFUL IN ANSWERING THE QUESTIONS IN THIS SECTION OF THE EXAMINATION.

DO NOT DETACH FROM BOOK.

PERIODIC TABLE OF THE ELEMENTS

1 **H** 1.0079																	2 **He** 4.0026
3 **Li** 6.941	4 **Be** 9.012											5 **B** 10.811	6 **C** 12.011	7 **N** 14.007	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.179
11 **Na** 22.99	12 **Mg** 24.30											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.974	16 **S** 32.06	17 **Cl** 35.453	18 **Ar** 39.948
19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.90	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.938	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.39	31 **Ga** 69.72	32 **Ge** 72.59	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90	36 **Kr** 83.80
37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.94	43 **Tc** (98)	44 **Ru** 101.1	45 **Rh** 102.91	46 **Pd** 106.42	47 **Ag** 107.87	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.71	51 **Sb** 121.75	52 **Te** 127.60	53 **I** 126.91	54 **Xe** 131.29
55 **Cs** 132.91	56 **Ba** 137.33	57 ***La** 138.91	72 **Hf** 178.49	73 **Ta** 180.95	74 **W** 183.85	75 **Re** 186.21	76 **Os** 190.2	77 **Ir** 192.2	78 **Pt** 195.08	79 **Au** 196.97	80 **Hg** 200.59	81 **Tl** 204.38	82 **Pb** 207.2	83 **Bi** 208.98	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
87 **Fr** (223)	88 **Ra** 226.02	89 **†Ac** 227.03	104 **Rf** (261)	105 **Db** (262)	106 **Sg** (263)	107 **Bh** (262)	108 **Hs** (265)	109 **Mt** (266)	110 **§** (269)	111 **§** (272)	112 **§** (277)						

§Not yet named

*Lanthanide Series	58 **Ce** 140.12	59 **Pr** 140.91	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.4	63 **Eu** 151.97	64 **Gd** 157.25	65 **Tb** 158.93	66 **Dy** 162.50	67 **Ho** 164.93	68 **Er** 167.26	69 **Tm** 168.93	70 **Yb** 173.04	71 **Lu** 174.97
†Actinide Series	90 **Th** 232.04	91 **Pa** 231.04	92 **U** 238.03	93 **Np** 237.05	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (260)

GO ON TO THE NEXT PAGE.

Metric Conversions

Unit	Symbol	*Equivalent Expressions*	
mega	M	**1 Mg** = 1,000,000 g = 10^6 g	**1 Mg** = 1,000,000 g = 10^6 g
kilo	k	**1 kg** = 1,000 g = 10^3 g	**1 kg** = 1,000 g = 10^3 g
hecta	h	**1 hg** = 100 g = 10^2 g	**1 hg** = 100 g = 10^2 g
deca	da	**1 dag** = 10 g = 10^1 g	**1 dag** = 10 g = 10^1 g
o		1g = 10^0g	1g = 10^0g
deci	d	**1 g** = 10 dg = 10^1 dg	**1 dg** = 0.1 g = 10^{-1} g
centi	c	**1 g** = 100 cg = 10^2 cg	**1 cg** = 0.01 g = 10^{-2} g
milli	m	**1 g** = 1,000 mg = 10^3 mg	**1 mg** = 0.001 g = 10^{-3} g
micro	µ	**1 g** = 1,000,000 µg = 10^6 µg	**1 µg** = 0.000001 g = 10^{-6} g
nano	n	**1 g** = 1,000,000,000 ng = 10^9 ng	**1 ng** = 0.000000001 g = 10^{-9} g
pico	p	**1 g** = 1,000,000,000,000 pg = 10^{12} pg	**1 pg** = 0.000000000001 g = 10^{-12} g

* Any quantity can be substituted for g; ie. 1 L = 1000 mL just as 1 g = 1000 mg

A helpful pnemonic for memorizing prefixes (you need to know these):

Many kids have dropped over dead converting metric measurements in problems.

Advanced Placement Chemistry Review Assignment

Topic 1: Significant Figures & Scientific Notation

1. Count the number of significant figures in the following measurements.
   a. 2.71 g _____ b. 0.00047 kg _____ c. 7.0×10^5 m _____ d. 1,030 L _____
   e. 150 pencils _____ f. 37500 µg _____ g. 0.1010 cm _____
2. Express each of the following in proper scientific notation (Pay attention to sig figs and units).
   a. 0.000125 m _____ b. 155.0 mL _____
   c. 123,030,000 ng _____ d. 481.9×10^{-9} cm _____
3. Calculate the correct answer with proper units and significant figures for each of the following:
   a. 12 g + 0.677 g + 86.33 g = _____
   b. (355.78 g) / (0.056 g) = _____
   c. 97.34 mL – 34.1 mL = _____
   d. 14.68 x 5 = _____
4. Perform the following calculations with scientific notation and report your answer with the correct number of significant figures.
   a. $0.14 \times (6.02 \times 10^{23})$ = _____
   b. $\frac{(9.875 \times 10^4) - (9.795 \times 10^4)}{9.875 \times 10^4} \times 100\ \%$ = _____ (assume 100 is exact)
   c. $\frac{(3.8 \times 10^{-12} \times 4.0 \times 10^{-13})}{(4 \times 10^{12} \times 6.3 \times 10^{13})}$ = _____

Topic 2: Dimensional Analysis

Show work using dimensional analysis. No work = no credit even if answer is correct. Follow significant figures and rounding rules unless the number of significant figures is specified. Include units where appropriate.

5. How many hours are in a week? Report your answer to three significant figures.

6. Find the number of centimeters in 1.00×10^2 yards. (1 yd = 3 ft, 1 ft = 12 in, 2.54 cm = 1 in)

7. If Jules Verne expressed the title of his famous book, Twenty Thousand Leagues under the Sea in basic SI units, what would the title be? Round your answer to three significant figures.
   (1 league = 3.45 mi, 1 mi = 1609 m)

8. How many μL are present in 250 mL of H_2O?

9. Wavelengths are often represented in nm. What is the diameter of a helium (He) atom in nm if it is equivalent to 1.0×10^{-13} km?

10. The area of a rectangular room has a length of 10.5 m and a width of 4.50 m. What is this area in m^2? In cm^2?

11. The acceleration of a sphere is determined to be 9.52 m/s^2. What is the acceleration in km/min^2?

<u>Topic 3: Density and Temperature</u>

Show all work. No work = no credit even if answer is correct. Follow significant figures and rounding rules. Include units where appropriate.

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

13. The density of pure silver is 10.5 g/mL at 20°C. If 5.25 grams of pure silver pellets are added to a graduated cylinder containing 11.2 mL of water, to what volume will the water in the cylinder rise?

14. You can figure out whether a substance floats or sinks if you know its density and the density of the liquid. In which of the liquids listed below will high-density polyethylene, HDPE, float? HDPE, a common plastic, has a density of 0.97 g/cm^3. It does not dissolve in any of the following liquids.

Substance	Density (g/cm^3)
ethylene glycol	1.1088
water	0.9997
ethanol	0.7893
methanol	0.7914
acetic acid	1.0492
glycerol	1.2613

15. Mercury is found as a liquid at room temperature. If it has a boiling point of 630. K, what is this boiling point in degrees Celsius?

Topic 4: Precision and Accuracy

16. The density of ethanol was determined experimentally at 25°C in a series of trials to be 0.608 g/mL, 0.705 g/mL, and 0.689 g/mL. The accepted density of ethanol is reported to be 0.789 g/mL.
    a. Are the experimental densities precise? Why/Why not?

    b. Calculate % error for this experiment. Use the average experimental density in your calculation and report your answer to 0.1%. Show your work.

    c. Are the experimental densities accurate? Why/Why not?

Topic 5: Properties and Changes

17. Categorize each of the following as an element, a compound, or a mixture:
    a. carbonated water _____
    b. tungsten _____
    c. aspirin (acetylsalicylic acid) _____
    d. air _____
    e. lye (sodium hydroxide) _____
    f. fluorine _____

18. Iron pyrite, also known as fool's gold, has a shiny golden metallic appearance. Crystals are often in the form of perfect cubes. A cube of iron pyrite measuring 0.40 cm on each side has a mass of 0.064 g.

   a. Which of these observations are qualitative and which are quantitative?

   b. Which of these observations are extensive (dependent on the amount of substance present) and which are intensive (independent of the amount of substance present)?

19. Identify the following as a physical property, physical change, chemical property, or chemical change:

   a. Ethanol has a density of 0.697 g/mL. _____

   b. The solution turns blue upon mixing water and food coloring. _____

   c. Wood burns in an oven. _____

   d. Methyl alcohol is highly flammable. _____

   e. Ice melts in a beaker. _____

   f. Methyl ethanoate smells like apples. _____

   g. Iron rusts on a car. _____

   h. Alkali metals react strongly in hydrochloric acid. _____

Topic 6: Atom Structure & History

20. How many protons and neutrons are contained in the nucleus of each of the following atoms? How many electrons are present in each of these neutral atoms?

   a. $^{13}_{6}C$ ____ protons ____ neutrons ____ electrons

   b. $^{208}_{82}Pb$ ____ protons ____ neutrons ____ electrons

21. Complete the following table:

Name	Mass #	Atomic #	# of Protons	# of Neutrons	# of Electrons	Symbol
Gallium-70					31	
						$^{31}_{15}P^{-3}$
Strontium-80					36	
						$^{55}_{25}Mn^{+2}$

22. The natural abundance for boron isotopes is 19.9% boron-10 (exact mass 10.013 amu) and 80.1% boron-11 (exact mass 11.009 amu). Calculate the average atomic mass of boron using the exact masses instead of mass numbers in your calculations. Show your work. Follow significant figures and rounding rules. Include appropriate units.

23. Europium has two stable isotopes, ^{151}Eu and ^{153}Eu, with masses of 150.9197 u and 152.9212 u, respectively. Calculate the percent abundances of these isotopes of europium to 0.1%. Hint: The percent abundances of these two isotopes must add to 100%. Show your work. Follow significant figures and rounding rules. Include appropriate units.

24. Identify the scientist(s) noted for the following events in atomic history.
   a. identified the electron; noted for the plum pudding model _____
   b. noted for the first atomic theory of the atom; solid sphere model _____
   c. developed the planetary model; electrons in fixed orbits _____
   d. developed the quantum mechanical model; electrons are localized to orbitals _____
   e. identified the proton and the nucleus; nuclear model _____
   f. determined the charge of an electron _____
   g. described wave theory _____
   h. known for the uncertainty principle _____
   i. developed quantum numbers _____

25. Identify the model of the atom described in the following statements.
   a. currently accepted model _____
   b. model that first included a subatomic particle _____
   c. model developed using the gold foil experiment _____
   d. original model of the atom; atom was thought to be "indivisible" _____
   e. model that only showed the movement of hydrogen's electron accurately; involved "quantums" _____

Topic 7: Periodic Table Structure

Identify by name the group or section of the periodic table noted for the following features.

26. a. group containing the most reactive nonmetals; all are diatomics; form -1 ions _____
   b. group containing metals that only form +2 ions _____
   c. set of metals that often form colored ions in solution; the majority have multiple charges as ions _____
   d. group containing the most reactive metals; form +1 ions _____
   e. group containing least reactive elements on periodic table, typically inert _____

27. These elements start with the letter B: B, Ba, Bk, Bi, and Br. Identify which of these elements match the following descriptions. You may use elements once, more than once, or not at all.
   a. Which are metals? _____
   b. Which are liquids? _____
   c. Which are actinides? _____
   d. Which are main block elements? _____

Topic 8: Compound Nomenclature

28. Name or give the formula for the following compounds. All ions included in the summer letter are required to be memorized by name and by formula.

	Name	Formula
a.	lithium fluoride	_____
b.	_____	K_2O
c.	calcium phosphate	_____
d.	_____	$MnCl_2$
e.	silver sulfide	_____
f.	_____	Cu_2O
g.	aluminum sulfate	_____
h.	_____	$ZnCO_3$
i.	chromium (III) phosphide	_____
j.	_____	SO_3
k.	lead (IV) hydroxide	_____
l.	_____	N_2O_5
m.	ammonium sulfite	_____
n.	_____	$BaCr_2O_7$
o.	sodium peroxide	_____
p.	_____	NH_3 (use common names; see ppt/videos if necessary)
q.	nickel (II) hypochlorite	_____
r.	_____	$Fe(CN)_3$
s.	rubidium chromate	_____
t.	_____	$Mg_3(PO_4)_2$

Topic 9: Equations

29. Balance the following equations using the lowest whole-number coefficients.

a. $\_\_Fe + \_\_P_4 \rightarrow \_\_Fe_3P_2$

b. $\_\_Ca + \_\_H_2O \rightarrow \_\_Ca(OH)_2 + \_\_H_2$

c. $\_\_Ba(OH)_2 + \_\_H_3PO_4 \rightarrow \_\_Ba_3(PO_4)_2 + \_\_H_2O$

d. $\_\_(NH_4)_2CO_3 + \_\_Al(ClO_3)_3 \rightarrow \_\_Al_2(CO_3)_3 + \_\_NH_4ClO_3$

30. Write balanced chemical equations for the following word equations. Use the lowest possible whole-number coefficients to balance the equations.

a. Aqueous solutions of ammonium sulfate and barium nitrate form a precipitate of barium sulfate and aqueous ammonium nitrate.

b. Elemental magnesium and oxygen gas combine to form solid magnesium oxide.

c. Chlorine gas and aqueous potassium bromide react to form bromine liquid and aqueous potassium chloride.

d. Solid copper (II) carbonate decomposes to form crystals of copper (II) oxide and carbon dioxide gas.

e. Sulfuric acid is neutralized by lithium hydroxide to form water and aqueous lithium sulfate.

f. Liquid benzene, C_6H_6, undergoes combustion in oxygen gas, making carbon dioxide gas and steam.

<u>Topic 10: Mole Conversions & Stoichiometry</u>

Show your work. No work = no credit. Follow significant figures and rounding rules. Include appropriate units.

31. a. Calculate the number of moles in 500. atoms of iron (Fe).

b. What is the molar mass of lead (IV) carbonate, $Pb(CO_3)_2$?

c. How many formula units are present in 87.2 grams of lead (IV) carbonate?

d. What percentage of oxygen is found in lead (IV) carbonate? Round your answer to 0.1%.

32. The reusable booster rockets of the U.S. space shuttle employed a mixture of aluminum and ammonium perchlorate for fuel. A possible reaction for this is:

__Al(s) + __ NH_4ClO_4(s) → __ Al_2O_3(s) + __ $AlCl_3$(s) + _ NO(g) + __H_2O(g)

a. Balance the above reaction using the lowest possible whole-number coefficients.

b. If 4.00 g of aluminum reacted completely, how many grams of aluminum oxide would be made?

c. If 4.18 g of aluminum chloride was produced, how many moles of ammonium perchlorate would be consumed?

d. How many molecules of nitrogen monoxide would form if 6.3×10^{25} formula units of aluminum oxide were also produced?

33. The decomposition of ammonia is shown in the following equation: $2NH_3(g) \rightarrow N_2(g) + 3H_2(g)$.

a. 42.0 g of nitrogen has what volume in liters at STP?

b. 150 L of NH_3 undergoes decomposition to form how many liters of hydrogen gas at STP?

c. How many liters of ammonia were decomposed at STP if 3.0×10^{23} nitrogen molecules were made?