Chapter 1 Chemistry: The Study of Change

Use the outlines to guide your way through the text. Don't read a chapter without my outline in front of you. No chemistry textbook perfectly matches the AP Chemistry curriculum. I chose Chang's textbook because I believe it has some of the clearest explanations of the important topics in chemistry. However, as with all chemistry textbooks, it has far more chemistry than any single chemistry course could possibly cover. My outlines will guide you through the topics you need to know for AP Chemistry. By using the outline, you will know which parts of the textbook you can skip. Keep each outline in your AP Chemistry binder for your final review.

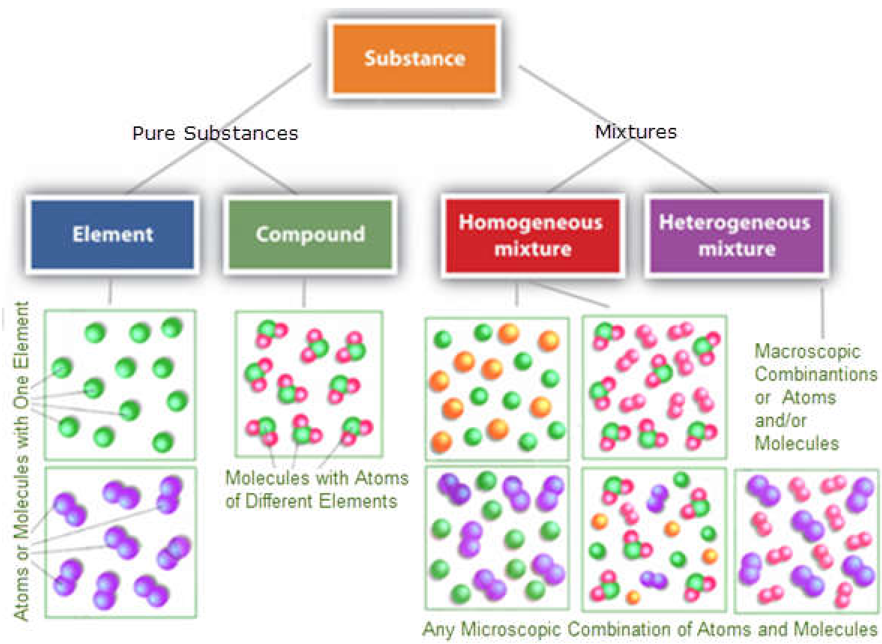
Chapter 1 – Chemistry: The Study of Change
~~1.1 Chemistry: A Science for the Twenty-First Century~~
~~1.2 The Study of Chemistry~~
~~1.3 The Scientific Method~~
1.4 Classifications of Matter
1.5 The Three States of Matter
1.6 Physical and Chemical Properties of Matter
1.7 Measurement
1.8 Handling Numbers
1.9 Dimensional Analysis in Solving Problems

1.1-1.3

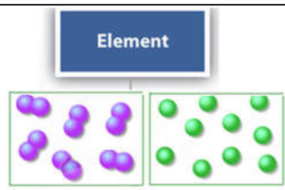
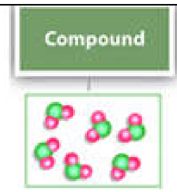
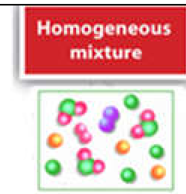
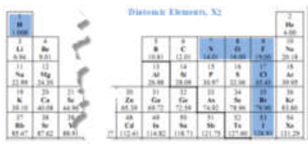
The material in sections 1.1 to 1.3 is not part of the AP Chemistry curriculum.

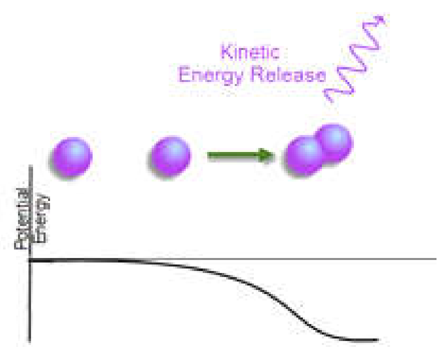
1.4. Classifications of Matter

Understand the differences between mixtures and (pure) substances.
After reading this section, look carefully at this drawing with atomic level drawings of the classifications.



Elements, compounds, and homogeneous mixtures are the classifications that are will be used most often.

		
There are only about 100 elements. Element atoms may bond with one another to form molecules of the element. The most important elemental molecules are the diatomic elements.  Some element atoms form more complex molecules.	Atoms of different elements bond to form compounds. There are many millions of known compounds. The properties of the elements are changed on the formation of the compound. Ratios of the atoms of any one compound are fixed and determined by the bonding abilities of the elements in the compound.	Atoms and molecules in a mixture are not bonded in fixed ratios as in compounds. Many properties of the components of mixtures are not changed. **Solutions and alloys are examples of mixtures.** The attractions between the components in mixtures are not as strong as bonds between atoms in molecules. If the mixture is uniform, it is homogenous.
The atoms of molecules can only be separated by breaking the strong bonds between atoms in the molecules. A chemical reaction or high energies are required to separate atoms of molecules.		Components of mixtures can be separated more easily than the atoms within molecules. Physical processes such as evaporation, condensation, differential solubility, fractional crystallization, and chromatography use the physical properties of each component to separate components of a mixture.

Memorize the seven diatomic elements. Not all elements exist as atoms. Seven elements spontaneously form pairs of atoms[1]. Under normal conditions, a single monatomic atom will bond with another monatomic atom to make a diatomic molecule. All bond formation releases kinetic energy. This release of energy lowers the potential energy of the substances. Chemists prefer to plot the formation of bonds as a drop in potential energy. A container of a monatomic element will have its atoms instantly chemically react to form diatomic molecules emitting light and heat in the process and producing diatomic atoms with less potential energy. Rather than describing diatomic elements as X, these elements are written as X_2 when writing their formulas. Since diatomic molecules consist of two atoms of the same element, they are not classified as compounds.	

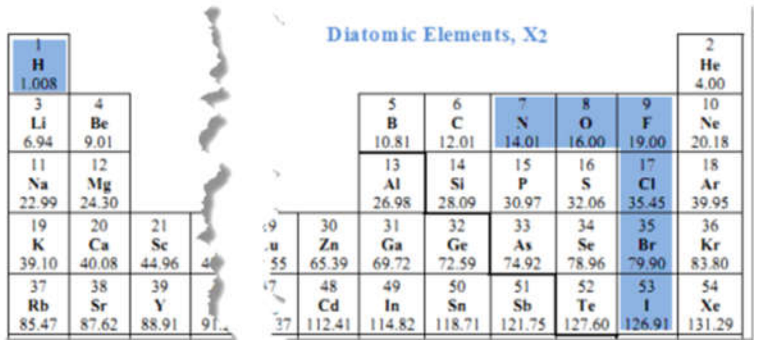
[1] Other elements bond to form elemental molecules also. P_4 and S_8 are two examples. However, they are not usually written as molecules in chemical equations and you are not expected to memorize their formulas.

Diatomic Substances

Hydrogen	H_2
Nitrogen	N_2
Oxygen	O_2
Fluorine	F_2
Chlorine	Cl_2
Bromine	Br_2
Iodine	I_2

The diatomic substances are written as diatomic when they appear in their normal, elemental form. Not writing these elements as diatomic in equations is almost always incorrect. In the few cases where the monatomic version is required, the element will be identified as atomic, e.g. atomic hydrogen would be H.

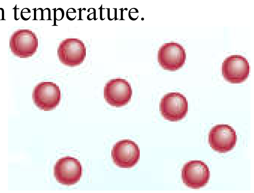
The diatomic substances can be easily remembered by their position on the periodic table.



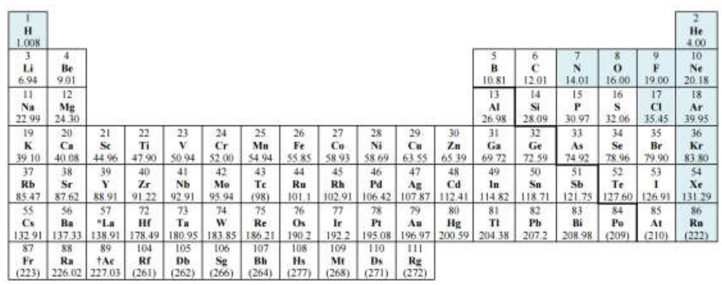
1.5 The Three States of Matter

Know the four abbreviations: (*s*) pure solid, (*l*) pure liquid, (*g*) pure gas, and (*aq*) aqueous mixture

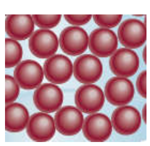
Memorize the elements that are gases at room temperature.



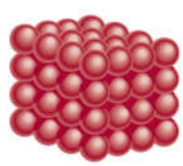
H_2,, N_2, O_2, F_2, Cl_2
He, Ne, Ar, Kr, Xe, Rn

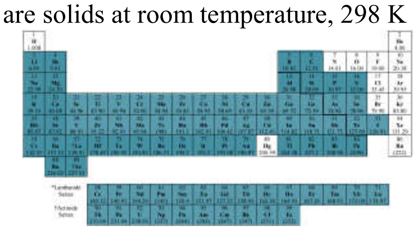


Memorize the elements that are liquids at room temperature, Br_2 and Hg.[2]



A majority of the elements are solids at room temperature, 298 K





[2] If the definition of room temperature 25°C were just a few degrees warmer, there would be two more elemental liquids, Cs and Ga. Several other elements would probably be liquids at room temperature, except macroscopic amounts of these substances have never been observed.

1.6 Physical and Chemical Properties of Matter

Physical properties can be directly measured and easily observed.

Mixtures can be separated using the physical properties of its components.

Examples of physical methods of separation are:

Distillation where the difference in boiling temperatures of the components allows for separation.
Freezing or cooling a solution so components selectively precipitate out.
Filtering a heterogeneous solution to separate small particles from a solution.
Adding a solvent to a mixture that selectively dissolves a component of a mixture.
Chromatography where the different attractions of each component to a separating media are used.

Chemical properties involve the reactions of a substance to change the substance's composition.
An example of a chemical property is sulfuric acid reacting with active metals to produce hydrogen gas and the sulfate ions.

Extensive properties depend on the amount of material. Mass is an extensive property and depends on the amount of material present.

Intensive properties do not change with the amount of material present. Density is an example of an intensive property.

1.7. Measurement

The terms *macroscopic* and *microscopic* are worth remembering. *Macroscopic* properties can be measured directly such as density. *Microscopic* relates to measurements that can only be made indirectly such as the electronic structure of atoms. Think of microscopic as atomic/molecular in scale.

SI Units
Units will help you understand chemistry and solve problems. Even if you do not understand a problem by using units and basic algebra, you can often manage your way through a chemistry problem using units.

If an official SI unit is named after a person, the first letter of its symbol is capitalized[3], but when writing the full name of the unit, the first letter is not capitalized. For example, meter, m, and kelvin, K.

[3] The rule is not followed exactly. As in this case of hertz, the derived unit for frequency named after Heinrich Hertz, Hz, and the first two letters are used. Historical roots and common usage will sometimes compromise the rule.

The most common prefixes used in chemistry should be memorized. The prefixes to know are:

k,	kilo,	10^{3}	
d	deci	10^{-1}	
c	centi	10^{-2}	
m	milli	10^{-3}	
µ	micro	10^{-6}	(mc is often used in place of the Greek letter mu)
n	nano	10^{-9}	
p	pico	10^{-12}	

Important tip to help with converting from one prefix to another:

The prefix letter can always be changed into its numeric value, e.g. 34 nm = 34×10^{-9} m. Likewise converting to a unit can be done by dividing by the numeric value of the prefix, e.g.

$$\frac{3.4 \times 10^{-9}\text{m}}{10^{-12}} = 3{,}400 \text{ pm}$$

Table 1.2 shows the **seven SI base units**. All other SI units of measurement can be derived from these base units. For example, the unit for energy, joule, J, is a derived unit. It is based on the kilogram, meter, and second. If you want to investigate this topic further, go to the US agency[4] on units:

National Institute for Science and Technology http://physics.nist.gov/cuu/Units/index.html

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The base unit most associated with chemistry is the unit for "Amount of substance" -- the mole

The amount of substance in chemistry is based on the number of particles of the substance, not its mass or volume.

One mole is the amount of substance that is based on the **number of constituent particles**. Usually these constituent particles are atoms, molecules, ions, electrons, or photons.

6.02×10^{23} particles of a substance is 1.00 mole

Even though one mole is a very large number of particles, a mole of atoms or molecules is a manageable size. One mole of water which is 6.02×10^{23} molecules of water only weighs 18 grams.
The mole is rarely used with anything significantly larger than molecules (e.g. 1 mole of snowflakes would be about 20 trillion tons of snow).

When the mole is used, the elementary entities should be specified. For example, 0.32 mol of H_2O indicates that the particles that are being counted are H_2O molecules.

Memorize Avogadro's number to 3 significant digits: 6.02×10^{23} particles.

[4] All systems of weights and measures, metric and non-metric, are linked through a network of international agreements. The International System is called the SI. The key SI agreement is the Treaty of the Meter signed in Paris in 1875. All the major industrialized countries have signed the treaty. The United States is a charter member having signed the original document in 1875. The SI is maintained in Paris, by the Bureau International des Poids et Mesures.

Mass and Weight[5]

The mass of a substance is determined by the number of particles and the mass of the individual particles in the sample. Since temperature does not change the number of particles, the mass of any substance does not change with temperature. Also the mass of a substance does not change with the phase of a substance.
The mass of a substance is constant. Mass is conserved in all chemical and physical reactions.

The official base of mass is kg. Mass is the only base unit that relies on a prefix. Any SI derived units that incorporate mass into the derived unit (such as joules) use kg rather than g.

Volume

The volume of a substance is determined by the spacing between the particles in the sample. The spacing of the particles changes with temperature and phase.
The volume of a substance can change with temperature and phase. Accurate measures of volume will also include the temperature the volume is measured at.
The official SI unit for volume is the m^3, but chemists use the liter, L, (dm^3) for liquid volume measurements. While not an official unit, the liter is accepted for use by the SI.[6] There have been numerous attempts to get chemists to abandon the liter in favor of dm^3, and the cm^3 for mL but they have not been successful.

The mL and cm^3 are now defined as identical units. The proper SI unit for volume, cubic centimeter written as cm^3 is often abbreviated as cc. AP Chemistry and most chemists use the liter and mL for volume.

$$1 \text{ cm}^3 = 1 \text{ cc} = 1 \text{ mL}$$

Density

Density differs from mass and volume in that it is an intensive property. $\text{Density} = \dfrac{\text{mass}}{\text{volume}}$

Since density measurements include volume, the density of a substance varies with temperature and phase. Know how use the density equation. Its use has been required on almost every AP Chemistry exam.

Remember the density of water at room temperature is 1 g/mL.

Solid and liquid densities are measured in g/mL and range from 0.2 g/mL to 20 g/mL.
Solid and liquid densities vary slightly with temperature as long as there is no phase change.

Gas densities are measured in g/L because the densities of gases are 1/1000 that of solids and liquids. Densities of gases have larger variations than solids and gases with changes in pressure and temperature. The density of air at room temp and pressure is about 1 g/L.

[5] There is a fundamental difference between the scientific meaning of weight and mass that is explained in the text. In chemistry however, weight and mass are used interchangeably. All of our mass/weight measurements will be done at normal gravity conditions, so we will not need to differentiate between the two terms as we would in a physics course.

[6] A humorous ruse was used to give the liter, L, an air of legitimacy. Claude Émile Jean-Baptiste Litre was a fictional character credited as the father of the Liter (and the burette).
http://en.wikipedia.org/wiki/Claude_%C3%89mile_Jean-Baptiste_Litre

Before dealing with temperature scales you need some critical background material that is not in the text.

Energy is not temperature

Energy, an extensive property (like mass).
Energy is measured using the SI derived unit, the joule, J. Most chemistry problems use kJ.
Just as grams can be used as a measure of the mass of different substances, joules can be used to measure different forms of energy (kinetic, potential, electrical, solar, nuclear, and chemical).
Annoyingly non-SI units are often used for energy such as the electron volt ($1 \text{ eV} = 1.6 \times 10^{-19}$ joules), the kilowatt hour ($1 \text{ kwh} = 3.6\times 10^{6}$ joules), and the Calorie (1 Cal = 4,160 joules). The AP Chemistry exam will only use joules for its unit of energy.

The joule unit and kinetic energy.

The kinetic energy of moving objects is determined by the mass of the object and its speed. The kinetic energy equation is on the AP Chemistry equations sheet:

$$KE = \frac{1}{2}mv^2$$

m is mass and *v* is velocity. The mass unit for this equation is kilograms, not grams! The velocity unit is meters per second (m/s).

1 J ≈ kinetic energy released when a small apple falls 1 meter in earth's gravity
1 kJ ≈ kinetic energy released in an Olympic hammer throw

In chemistry when talking about heat, you are talking about a quantity of thermal energy.

Now to the unit measure that is often confused as being a measurement of energy:

Temperature is not a measure of total energy.
Temperature is a measure of the average kinetic energy of atoms and molecules.
Temperature is an intensive property (like density).

The SI unit of temperature is the kelvin, K.

The Kelvin scale is set up so that 0 K is the coldest possible temperature.

While commonly used, degrees Celsius is not a true measurement of average kinetic energy. 0°C is 273°C above the lowest temperature, 0 K.
Calculations using the true magnitude of average kinetic energy require kelvins.

The negative values used with degrees Celsius do not really indicate a negative temperature. There is no true negative temperature just as there is no negative density.

Memorize how to convert between °C and K: 273 + °C = K

When you just want to compare **changes** in temperature, ΔT, degrees Celsius can be used interchangeably with kelvins since the temperature span of a °C is the same as the span of a K.

ΔT, as in all Δ calculations, is the second measurement minus the first measurement, $T_2 - T_1$.

Don't worry about °F. There will never be an AP question using degrees Fahrenheit.

Remember for purposes of chemistry **temperature is not heat!**

1.8 Handling Numbers
Review how to use scientific notation without using your calculator. Half the AP Chemistry exam does not allow the use of a calculator and you will very likely need to estimate answers involving calculations without a calculator using estimation and basic math skills. Scientific notation will allow you to make difficult math calculations without a calculator as long as you can add and know your simple multiplication tables.

602×0.000786
$\approx 6\times10^{2} \times 8\times10^{-4}$
$\approx 6\times8 \times10^{(2-4)}$
$\approx 48 \times10^{-2}$
$\approx 0.48\ldots$ (answer using a calculator is 0.475)

Here is an example of a typical problem you would be expected to solve without a calculator.

$$1.2\times10^{-8} = \frac{x^2}{(0.20-x)}$$ where x is less than 1×10^{-3}

$$1.2\times10^{-8} = \frac{x^2}{0.20}$$

Since 1×10^{-3} is much less than 0.20 then $(0.20 - x) \approx 0.20$. Which makes solving the problem much easier.

$$1.2\times10^{-8} \times 0.20 = x^2$$
$$1.2\times10^{-8} \times 2.0\times10^{-1} = x^2$$
$$2.4\times10^{-9} = x^2$$
$$24\times10^{-10} = x^2$$
$$5\times10^{-5} \approx x$$

The answer done without rounding is 4.9×10^{-5}. Being able to approximate answers is one of the skills that separates 5's from 4's on the AP Chem exam.

Know how to enter and read scientific notation on your scientific calculator.

All scientific calculators have a special way of entering scientific notation.
Never use the multiplication key when typing in scientific notation.
Scientific and graphing calculators have a special key for scientific notation entry.

Often this key involves using a second function as with Texas Instrument graphing calculators.
To enter the number of particles in one mole, 6.02×10^{23}, on a TI calculator[7], you would enter

6.02 [2ND][EE]23

Try the next problem on your scientific calculator to see if you can use scientific notation on your calculator. The multiplication signs that are part of the scientific notation in this problem will require the use of the scientific notation key on your calculator. If you are typing in ×10 using the multiplication key and the number ten, you are not using your scientific calculator correctly! Only use the multiplication key once in this problem.

$(6.02 \times10^{23}) \times (6.63 \times10^{-34}) = 3.99 \times10^{-10}$ (rounded to three significant figures)

If you are not certain how to enter scientific notation on a TI calculator, watch the video on the TI-84 and scientific notation at the ChemAdvantage website.

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[7] Don't be misled by the [10^x] function on the TI-83 and TI-84 calculators. That key is the antilog base 10 key.

Significant Figures (Summary of Chang's excellent explanation)

All non-zero digits are significant 1456 (4 significant figures)	Trailing zeros are only significant with a decimal point 100.0 g (4 significant figures) 100 g (1significant figure)
Zeros bracketed by nonzero digits are significant 1001cm (4 significant figures)	Leading zeros are never significant. 0.001 kg has only one significant figure

Definitions and exact counts have a non-limiting number of significant figures.

Sometimes it is difficult to round and answer to a specific number of significant figures without using scientific notation.

For example, rounding 297 to two significant figures would require the use of scientific notation since rounding to 300 would show 1 sig fig and 300. would show 3 sig figs. When in a situation like this, use scientific notation, 3.0×10^2.

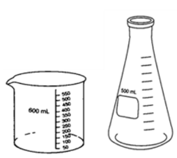
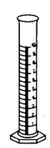
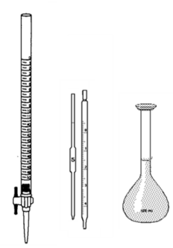
The Argyle Sweater by Scott Hillburn

http://www.theargylesweater.com/

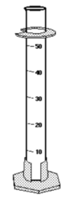
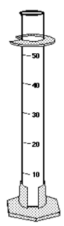
Significant Figures and Lab Measurements

The accuracy of a measurement is reflected in its significant figures.

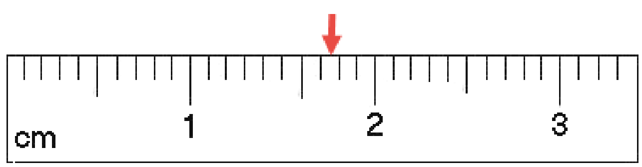
Volume measurements in a lab often limit the accuracy of a measurement. In terms of accuracy and precision.

Least accurate	Intermediate accuracy	Highest accuracy
beakers and flasks 	graduated cylinders 	burettes, pipets, volumetric flasks 

Typically, lab measurements can be considered as being ± 1 of the last (rightmost) significant figure. For example, these measurements of a 50-mL volume were made using three types of volumetric glassware

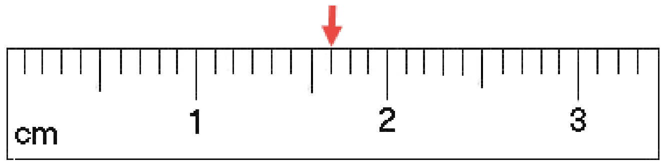
150 mL beaker (± 10-mL)  50 mL 1 sig fig The measured volume could be between 60 mL and 40 mL. However, 120 mL of liquid would be measured to 2 significant figures. Greater volumes can lead to increased accuracy.	50-mL Student Grade graduate cylinder (± 1-mL)  50. mL 2 sig figs The measured volume could be between 51 mL and 49 mL.
50-mL Grade A graduate cylinder (± 0.1-mL)  50.0 mL 3 sig figs The measured volume could be between 50.1 mL and 49.9 mL.	Most Accurate volumetric flask (± 0.01-mL)  50.00 mL 4 sig figs The measured volume could be between 50.01 mL and 49.99 mL. This level of accuracy would require that the volume be measured close to the temperature the volumetric flask was designed for.

Many lab measurements are done using digital equipment, but some measurements involve interpreting scales. Examples are measurements done with rulers and accurate volumetric lab ware such as burettes and graduated cylinders. When using graduations (lines), you are expected to estimate the value between the graduations on the scale.



The distances between the graduations on this ruler can easily be estimated. So giving the position of the arrow either as 1.6 cm or 1.7 cm would be considered a poor measurement. It would be expected that a student reporting this distance would estimate the arrow at 1.65 cm. The 0.05 is an estimate, but still it has some value. The ideal report of the point on the ruler would be 1.65 +/- 0.1 cm. However, reporting the latitude of the accuracy of measurements can be tedious and is not required in AP. On an AP exam reporting the length as 1.64 cm or 1.65 cm or 1.66 cm would all be considered correct answers.

Here is another point on the ruler.



Entering this position on the ruler as 1.6 cm would be considered incorrect because 1.6 cm would have only 2 sig figs and would not correctly represent the accuracy of the measurement. This point would be better reported as having three significant figures, 1.60 cm. Bottom line: When it is possible, read between the lines!

Significant Figures and Addition and Subtraction
Adding or subtracting with significant figures:

1. Set up the problem in column form with the decimal points aligned.	2. Perform the calculation ignoring the significant figures in the measured values.	3. Since the values of the hundredths and thousandths of 1.0 g are not certain, the addition or subtraction in those columns is not certain.	4. From right to left round your answer to the first complete column.
12.051 g - 2.0 g	12.051 g - 2.0 g 10.051 g	12.051 g - 2.0?? g 10.0~~51~~ g	12.051 g -2.0 g 10.1 g

While in almost all answers on the AP Chemistry exam using three significant figures will be satisfactory, in lab experiments problems, rules must be followed exactly. Especially for addition and subtractions. There will be one lab question in the FRQ section that will have lab data. Keep on your best behavior with significant figures as this is where traps are set to differentiate 5's from 4's. There is a more detailed explanation of this at the end of the outline.

Significant Figures and Multiplication and Division

When multiplying and/or dividing, the final answer is rounded to the lowest number of significant figures of the values used in the computation.

Precision and Accuracy
The dart board analogy is excellent, but you must correlate it to experimental data.

High precision within the limits of the experimental values indicates consistent measurements. However, the consistent answer may or not be close to the correct answer. **The most common reason for an inaccurate, but precise answer is a methodical error.** For example, even if you find the density of a liquid precisely, the liquid could be contaminated so the value for that precise answer would be inaccurate.

Accuracy compares a result to a correct answer. An accepted method of determining accuracy is % error.

$$\%\text{ Error} = \frac{\text{experimental results - correct value}}{\text{correct value}} \times 100$$

A positive % error indicates the error is higher than the correct value.
A negative % error indicates an answer below the correct value.
The absolute value of the % error indicates the accuracy.

A series of imprecise measurements can be averaged together to make a more accurate answer if the variations in data are random so that high and low values cancel.

The greater the number of significant figures in a measurement, the greater the precision and accuracy.

1.9 The Factor-Label Method of Solving Problems

Showing the method of solution on free response problems is as important as the answer.
The Factor-Label Method is the ideal way to solve most chemistry problems and also show your work at the same time.

Some students like to use a box method. Some like to use a fraction format.

e.g. The density of liquid nitrogen at its boiling point , 77 K, is 0.808g/cm^3. Convert the density units to kg/m^3

Box Method	Fraction method
$\begin{array}{\|c\|c\|c\|c\|c\|} \hline 0.808\ \cancel{g} & 1\ kg & (1\ \cancel{cm})^3 & = & 808\ kg \\ \hline 1\ \cancel{cm}^3 & 1000\ \cancel{g} & (1\times10^{-2}\ m)^3 & & m^3 \\ \hline \end{array}$	$\frac{0.808\ \cancel{g}}{1\ \cancel{cm}^3}\times\frac{1\ kg}{1000\ \cancel{g}}\times\frac{(1\ \cancel{cm})^3}{(1\times10^{-2}\ \cancel{m})^3} = 808\ kg/m^3$

5 steps to solving a numerical problem is worth your time.

1. Read the questions carefully. Understand what is given in the problem and what you are asked to solve. A sketch or outline of the problem may be helpful.
2. Find the appropriate equation(s). Dimensional analysis is often helpful in solving problems.
3. Check your answer for the correct sign, units, and significant figures.
4. Look at the answer to see if it is reasonable. As you go through the course you should get a sense as to what the magnitudes of many answers will be. Always check your answer to see if it falls within these magnitudes (e.g. densities are never greater than 25 g/mL)

When you do multiple step problems, do not round off your calculator's intermediate answers. Rounding through a series of calculations leads to rounding errors! This is especially important in WebAssign problems. If you round off your intermediate answers, you may exceed the numerical tolerance in WebAssign, and your answer will be marked as incorrect.

Whenever possible do not retype intermediate answers in your calculator. Either chain your calculations or use the [ANS] (last answer) key. This will prevent transcription and cumulative rounding errors.

e.g. For example this calculation for the number of molecules in 3.73 liters of gas at 1.07 atm of pressure and a temperature of 10. °C

Cancelling units in the calculation	$\frac{1.07\ atm \times 3.73\ L}{0.0821\frac{atm\ L}{mol\ K}\times 283\ K}\times\frac{6.02\times10^{23}\ molecules}{1\ mol}$ $\frac{1.07\ \cancel{atm} \times 3.73\ \cancel{L}}{0.0821\frac{\cancel{atm}\ \cancel{L}}{\cancel{mol}\ \cancel{K}}\times 283\ \cancel{K}}\times\frac{6.02\times10^{23}\ molecules}{1\ \cancel{mol}}$
As a single, longer entry, on your calculator:	1.07 × 3.73 ÷ .0821 ÷ 283 × 6.02E23 [ENTER] 1.034092785E23

The answer rounded to the correct number of significant figures is 1.03×10^{23} molecules. The huge value for the answer makes sense since you would expect a large number of molecules to be present in a macroscopic sample of gas.

Important! FRQ problem points may be deducted if your answer's significant figures are more than 1 off.

Even though you may have shown your work and had the correct calculations, you can lose the point for the calculation if you display your answer to too many or too few sig figs.

Normally you are allowed a latitude of +/- 1 on the sig fig count. Therefore, if you use 3 significant figures in your answers on the AP Exam you will almost always have the correct number of significant figures.

When in doubt use 3 Significant figures on AP Exam numerical answers

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An exception to the 3 sig fig guideline: Lab Based Problems.

On lab data problems you must have exact number of significant figures. There is no +/- 1 latitude for the number of sig figs.

If a problem has lab measurements **especially with addition or subtraction** such as changes in temperature or subtracting the weight of a container, use the rules of adding or subtracting significant figures carefully and abandon the 3 sig fig guideline.

e.g. A filter paper is weighed to be 2.54 g. The filter paper is used to separate a solid from a heterogeneous mixture. After filtering the solid, the filter paper dried until it reaches a constant mass and then reweighed. Its new mass of solid and paper was found to be 2.12 g. What is the mass of the solid that was filtered?

2.54	g	mass of filter paper and solid
-2.12	g	mass of filter paper
0.42	g	mass of solid

The answer is only reported to two significant figures since there is no data that gives you any information on the thousandths of a gram value. Answering 0.420 g following the 3 sig fig rule, would cause the loss of 1 point. Knowing this exception can be the difference between a 4 and 5.

Extra topic on measurement:

Current theories are unable to describe space and time below certain limits. The concepts of distance and time lose meaning below these levels.

Planck length: 1.616×10^{-38} meters.
Planck time: 5.391×10^{-44} seconds

Please keep your lab measurements above these values lest you destroy the space time continuum.[8]

[8] "Out of the Fabric," Tom Siegfried, Science News, April 23, 2011 page 29