Chapter 1 Chemistry: The Study of Change

Use the outlines to guide your way through the text. No chemistry textbook perfectly matches the AP Chemistry curriculum. All college level chemistry textbooks have 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. Do not do the problems in the textbooks because they are not like AP Chemistry questions. The WebAssign problems will prepare you for my tests and for the AP exam.

Keep each outline in your AP Chemistry binder for your final review.

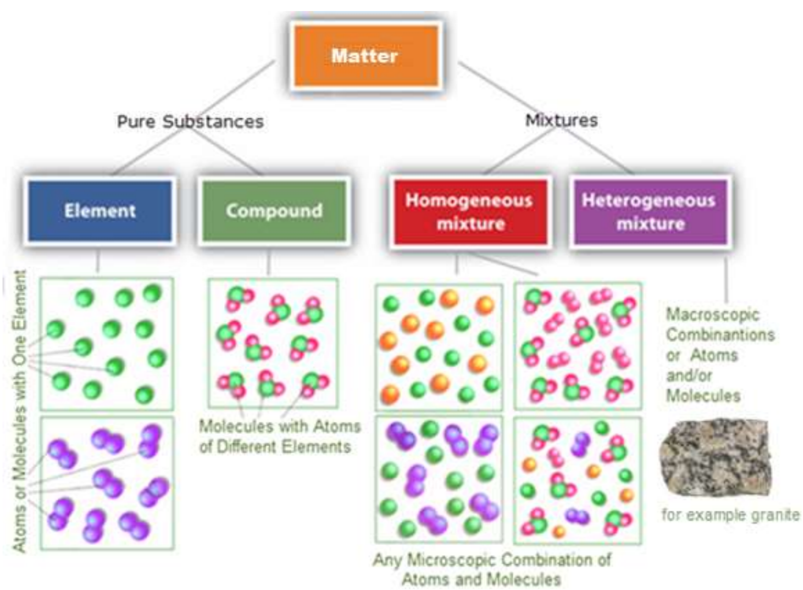
Chemistry by Chang and Goldsby	Chemistry 2e
Chapter 1 – Chemistry: The Study of Change	Chapter 1: Essential Ideas
~~1.1 Chemistry: A Science for the Twenty-First Century~~	~~1.1 Chemistry in context~~
~~1.2 The Study of Chemistry~~	1.2 Phases and Classification of Matter
~~1.3 The Scientific Method~~	1.3 Physical and Chemical Properties
1.4 Classifications of Matter	1.4 Measurements
1.5 The Three States of Matter	1.5 Measurement
1.6 Physical and Chemical Properties of Matter	1.6 Mathematical Treatment of Measurements
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.



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

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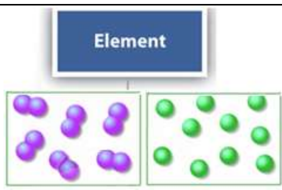
<td>
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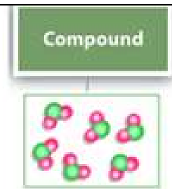
<td>
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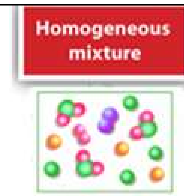
<td>
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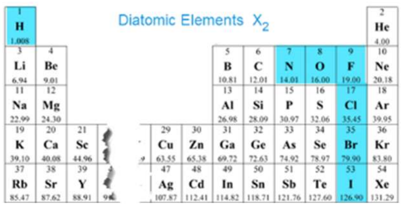
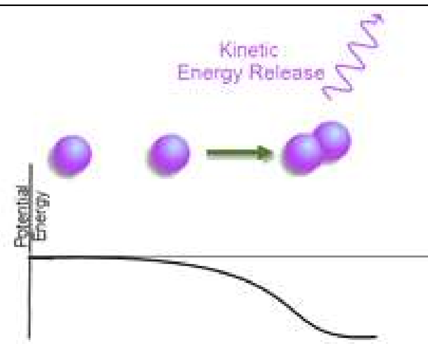
<td>There are only about 100 elements. Atoms of an element may bond with one another to form molecules. The most important elemental molecules are the diatomic elements.


Some element atoms (e.g. C, Si, P) form more complex molecules.</td>
<td>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.</td>
<td>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.
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.
Solutions and alloys are homogenous mixtures.</td>
</tr>
<tr>
<td colspan="2">The atoms of molecules can only be separated by breaking the strong bonds between atoms in the molecules. A chemical reaction, an electric current, or high energies are required to separate the atoms making up the molecules.</td>
<td>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 the components of a mixture.</td>
</tr>
</table>

Not all elements exist as atoms. Seven elements spontaneously form pairs of atoms[1]. Under normal conditions, a single atom of one of these elements will “bond” with another atom of the element to make a diatomic molecule. A container of single atoms of these elements will instantly chemically bond to form diatomic molecules releasing kinetic energy (emitting light and heat) with less potential energy than the monatomic atoms.	

[1] Other elements bond to form elemental molecules. P_4 and S_8 are two examples. These molecular elements are not usually written as molecules in chemical equations. While you may encounter P_4 and S_8 on the AP Chemistry test, their formulas will be given so there is no need to memorize them.

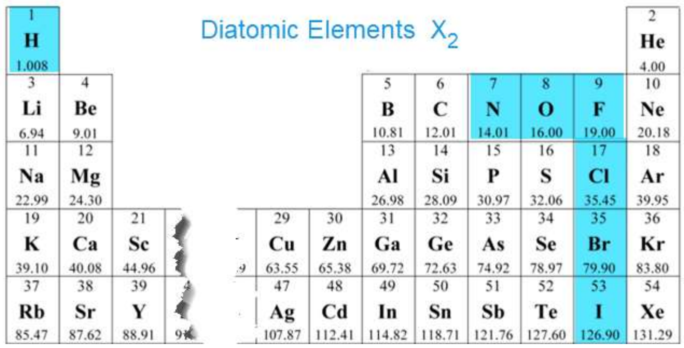
Memorize the Diatomic Elements.

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, while hydrogen would be H_2.

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

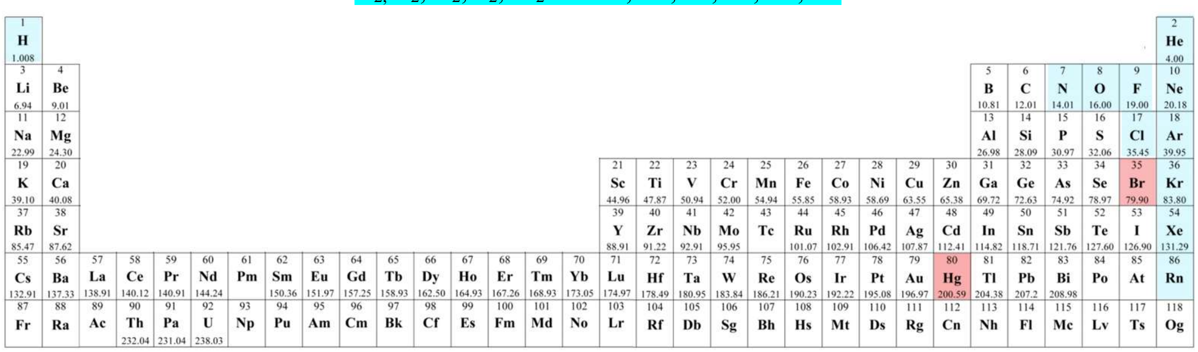


1.5 The Three States of Matter

Know the four abbreviations: (*s*) **pure** solid, (*l*) **pure** liquid, (*g*) **pure** gas, and (*aq*) aqueous **mixture.**
Formulas with (*s*), (*l*), or (*g*) are pure substances with fixed physical and chemical properties.
Formulas with (*aq*) are mixed with water whose properties vary with the concentration of the solutions.

Using a periodic table, you should be able to recognize the elements that are gases at room temperature.

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



Except for Br_2 and Hg, all of the rest of the elements are are solids at room temperature[2], 298 K.

[2] If the definition of room temperature, 25°C, was just a few degrees warmer, there would be two more elemental liquids, Cs and Ga. Several other elements could be liquids at room temperature, but macroscopic amounts of these substances have never been produced to observe their actual phase.

1.6 Physical and Chemical Properties of Matter

Physical properties can be directly measured and easily observed. Examples of physical properties are boiling temperature, density, and color.

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

Examples of physical methods of separation are:

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

Except for the last process, all of these will be covered in AP Chemistry.

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 and energy are extensive properties and depend on the amount of material present.

Intensive properties do not change with the amount of material present. Density and temperature are examples of intensive properties.

1.7 Measurement

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

SI Units

On any AP Chemistry test there will problems where not paying attention to units will result in an incorrect answer. Units will help you understand chemistry and solve problems. Even if you do not understand a chemistry problem it may be possible for you to manage your way to a correct solution by using units and basic algebra. Units are an excellent double check as to the validity of your method of solution.

Use the same rules for units as you use with variables in algebra	
$2x \times 3x = 6x^2$	$2\text{ m} \times 3\text{ m} = 6\text{ m}^2$
$2x + 3x = 5x$	$2\text{ m} + 3\text{ m} = 5\text{ m}$
$2x + 3y = ?$ cannot be added without additional information If $y = 0.01x$, then $3y \times \frac{0.01x}{y} = 0.03x$ $2x + 0.03x = 2.03x$	$2.00\text{ m} + 3\text{ cm} = ?$ cannot be added without additional information Since $c = 0.01$, then $3\text{ cm} \times \frac{0.01\text{ m}}{\text{cm}} = 0.03\text{ m}$ $2.00\text{ m} + 0.03\text{ m} = 2.03\text{ m}$

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. As in any language, historical roots and common usage will sometimes compromise the grammar rules.

This list has the most common prefixes used in chemistry and they should be memorized:

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, μ)
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.

Converting units can be done by using the ratio of the unit and its numeric value.
e.g. changing m into pm:

$$3.4 \times 10^{-9}\text{m} \times \frac{1\text{pm}}{10^{-12}\text{m}} = 3.4 \times 10^{-9}\cancel{\text{m}} \times \frac{1\text{pm}}{10^{-12}\cancel{\text{m}}} = 3.4 \times 10^{-9+12}\ \text{pm} = 3.4 \times 10^{3}\ \text{pm} = 3{,}400\ \text{pm}$$

Table 1.2 in the text 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 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.

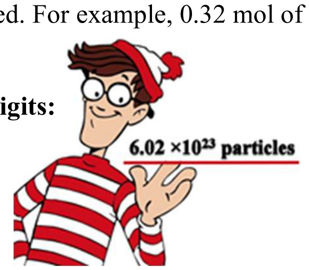
One mole is the amount of substance the **Avogadro number of constituent particles**. Usually these constituent particles are atoms, molecules, ions, electrons, or photons.

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.
One mole of uranium atoms, the heaviest primordial element, weighs about a half pound, 238 g.
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:



[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 constant.

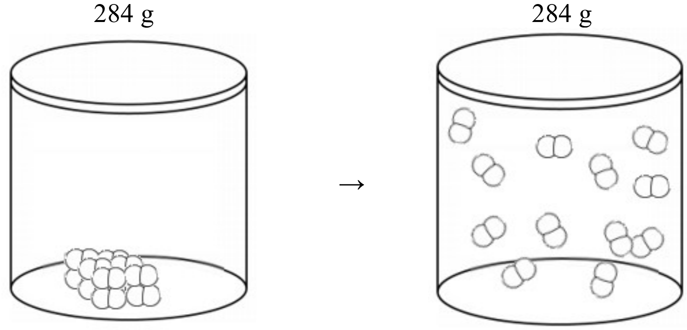
The mass of a substance is based on the number of particles **and** the mass of the individual particles. The mass of any substance does not change with temperature. Also, the mass of a substance does not change with the phase of a substance.

Mass is conserved in all chemical and physical reactions.

Physical Reaction

$I_2(s)$ → $I_2(g)$

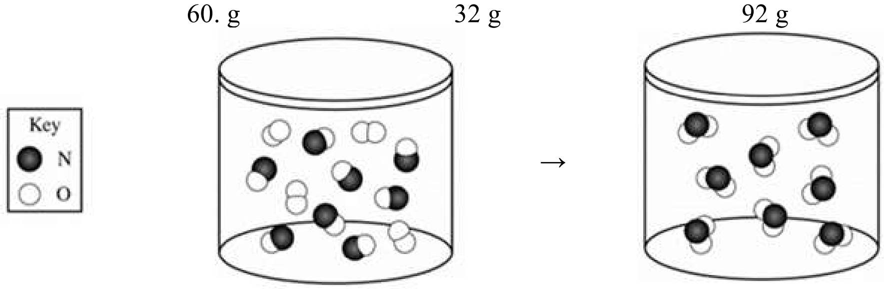
284 g   284 g



Chemical Reaction

$2\ NO$ + O_2 → $2\ NO_2$

60. g   32 g   92 g



While grams are the primary unit of mass used in AP Chemistry questions, kg is the official base unit. Any SI derived unit that incorporates mass as part of its derivation (such as joules) uses kg rather than g.

Volume

The volume of a substance can change with temperature and phase.

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.

Accurate measures of volume will also include the temperature associated with the volume measurement.

[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.

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 uses the liter, L, and milliliter, mL, for volume.

$$1\ cm^3 = 1\ cc = 1\ mL$$

Density

Know how to use the density equation. Its use has been required on almost every AP Chemistry exam.

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

Density differs from mass and volume in that it is an intensive property. Since density measurements include volume, the density of a substance varies with temperature and phase. While the SI derived unit for mass density is kg/m^3, AP Chem uses two other units of density.

AP Chem solid and liquid densities: g/mL (or 1 g/cm^3) and range from 0.2 g/mL to 20 g/mL.
Remember the **density of water** at room temperature is **1.0 g/mL**.

AP Chem Gas densities: g/L
The density of air at room temp and pressure is about 1 g/L.

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

Energy is the property that can heat an object or perform work on the object. In AP Chemistry we will not deal with the work aspect of energy

The SI unit of Energy is joules.

Energy is an extensive property (like mass).
The SI derived unit for energy is the joule, J, but most chemistry problems use kJ.

1 J ≈ can power a smart phone for about 10 seconds
1 kJ ≈ can power a refrigerator for about 10 seconds

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 ($1\ eV = 1.6 \times 10^{-19}$ joules), the kilowatt hour ($1\ kwh = 3.6\times 10^{6}$ joules), and the Calorie (1 Cal = 4,160 joules).
The AP Chemistry exam will only use the joule for its unit of energy.

In chemistry we will dealing with potential energy, where chemicals have the ability to release heat, kinetic energy in a chemical reaction.

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

[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

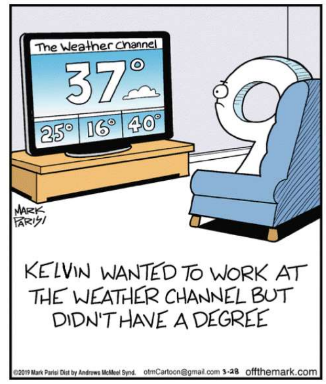
The SI unit of temperature is the kelvin, K.

Temperature is a measure of the **average kinetic energy** of atoms and molecules.

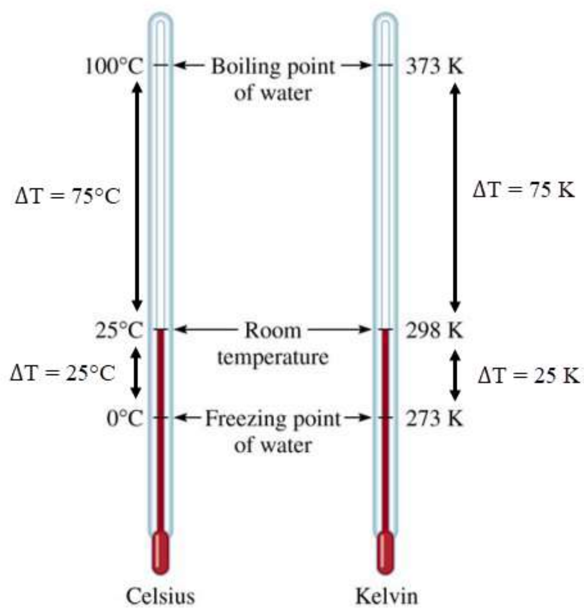
Temperature is as intensive property like density. Just as there is a significant difference between density and mass, there is a difference in temperature and energy.

The kelvin scale is set up so that 0 K is the coldest possible temperature. There is no true negative temperature just as there is no negative density. While commonly used, degrees Celsius is not a true measurement of average kinetic energy. Negative degrees Celsius do not really indicate a negative temperature.

Even though the equation for converting °C and K is on the AP Chemistry Equations Sheet, memorize how to convert[7] between °C and K: 273 + °C = K



Off the Mark by Mark Parisi [8]



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

When comparing changes in temperature, ΔT, degrees Celsius can be used interchangeably with kelvins. The reason for this is that the span of a °C is the same as the span of a K.

e.g. T_1 = 0.°C = 273 K T_2 = 25°C = 298 K

$\Delta T = T_2 - T_2$
ΔT = 25°C – 0.°C ΔT = +25°C
ΔT = 298 K – 273 K ΔT = +25 K

When dealing with Δ*T*, do not add 273 to convert °C to K.

A positive ΔT indicates warming, an increase in average kinetic energy.
A negative ΔT indicates cooling, a decrease in average kinetic energy.

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

[7] The actual equation is 273.15 + °C = K, but for all the AP questions 273 + °C = K will work.

[8] *Off the Mark* by Mark Parisi March 28, 2019
https://www.gocomics.com/offthemark/2019/03/28?ct=v&cti=1289522

1.8 Handling Numbers

The Multiple Choice (MC) section of the AP Chemistry exam does not allow the use of a calculator. Therefore, you will need to be able to do math operations without a calculator.

Scientific notation and simplifying numbers will allow you to make difficult math calculations without a calculator.

e.g 602×0.000786 without a calculator

simplifying numbers:	$602 \approx 600$	$0.000786 \approx 0.0008$
scientific notation:	$600 = 6\times10^{2}$	$0.0008 = 8\times10^{-4}$

$\approx 6\times10^{2} \times 8\times10^{-4}$

$\approx 6\times8 \times10^{(2+-4)}$

$\approx 48 \times10^{-2}$

$\approx 0.48\ldots$ (answer using a calculator is 0.475)

Being able to approximate answers without a calculator is a skill that required on the AP Chem exam.

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

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

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

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

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

$$1.2\times10^{-8} \times 2.0\times10^{-1} = x^2$$

$$2.4\times10^{-9} = x^2$$

$$24\times10^{-10} = x^2$$

$$5\times10^{-5} \approx x$$

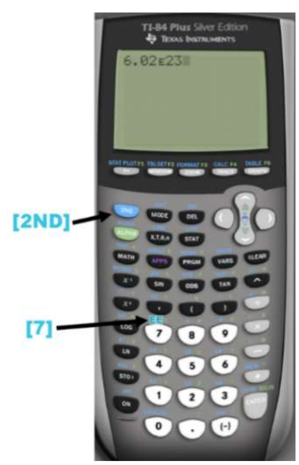
The answer without rounding is 4.9×10^{-5}.

You are allowed to use a calculator on the Free Response Question (FRQ) section of the AP Chemistry exam. In fact, you must sign a waiver if you do not have a calculator for this part of the exam.

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

Never use the multiplication key when typing in scientific notation. Scientific and graphing calculators have a special key for scientific notation entry. Often the scientific notation key involves using a second function as with TI graphing calculators. To enter the number of particles in one mole, 6.02×10^{23}, on a TI calculator , you would enter

6.02 [2ND][EE]23



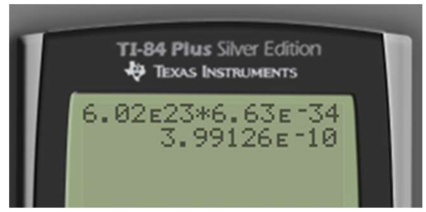
Try the next problem on your scientific calculator to see if you can use scientific notation on your calculator. The scientific notation in this problem will require the use of the scientific notation key on your calculator. If you are typing in $\times 10$ by using the multiplication key and the number ten, you are not using your scientific calculator correctly!

In performing the following calculation involving scientific notation

$(6.02 \times 10^{23}) \times (6.63 \times 10^{-34})$

you would only use the multiplication key once

6.02 EE23 × 6.63 EE(-)34



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.

On entering scientific notation in WebAssign, E notation, where "E" stands for $\times 10$, must be used.

3.99×10^{-10} would be entered in a WebAssign answer as 3.99E-10.

Significant Figures

The number of significant figures in a measurement indicates the precision and accuracy of a measurement. A measurement with one significant figure is crude approximation. The range of significant figures on the AP Chemistry exam is a minimum of one to a maximum of five significant figures. The most common number of significant figures used on AP Chemistry exams is 3.

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 show a specific number of significant figures without using scientific notation.

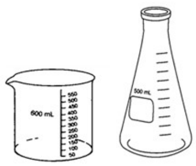
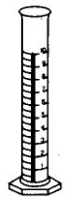
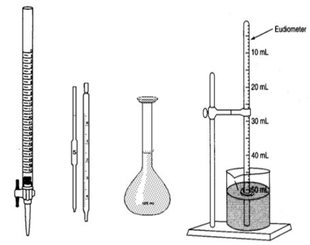
For example, rounding 297 to two significant figures would require the use of scientific notation since rounding to 300 would show 1 sig fig and 300. would show 3 sig figs. When in a situation like this, use scientific notation, 3.0×10^2.

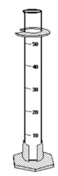
The Argyle Sweater by Scott Hillburn
http://www.theargylesweater.com/

Significant Figures and Lab Measurements

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

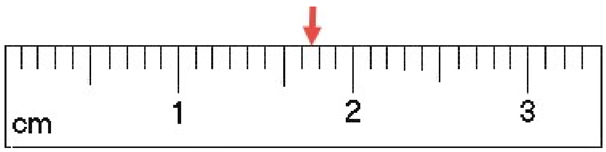
Volume measurements in a lab often limit the accuracy of a measurement. In terms of accuracy and precision.

Least accurate	Intermediate accuracy	Highest accuracy
beakers and flasks 	graduated cylinders 	burettes, pipets, volumetric flasks, eudiometers (gas collecting tube) 

150 mL beaker (± 10-mL)  1 sig fig Beakers and flasks volume markings are approximations and **cannot be used for accurate measurements**.	25-mL Grade A graduate cylinder (± 0.1-mL)  25.0 mL - 3 sig figs Graduate cylinders can be used to measure varying amounts of liquid but typically are only accurate to ± 0.1mL
50-mL Buret (± 0.01-mL)  0-50.00 mL – 4 sig figs for measurements greater than 10 mL. A grade A buret can dispense a measures volume with an accuracy of a couple of drops, 0.01 mL. A starting and final volume are required, and the accuracy is determined following the subtraction rules.	50-mL volumetric flask (± 0.01-mL)  50.00 mL - 4 sig figs A volumetric flask can only be used to measure the prescribed volume of the flask. The measured volume of a 50-mL volumetric flask 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.

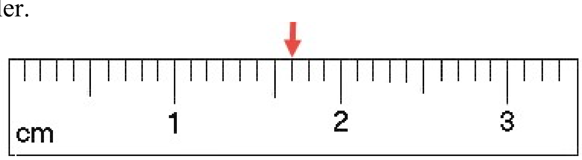
Many lab measurements are done using digital equipment, but some measurements involve interpreting scales. Measurements done with rulers and volumetric labware such as burettes, pipets, eudiometers, and graduated cylinders all require interpreting the accuracy of a scale.

When you report a measurement made using a scale with graduations (lines), you are expected to estimate the value between the graduations on the scale.



In reporting the position of the arrow on this ruler, you must understand that the spacing of the marks on the scale are 0.1 cm. Therefore, the arrow is between 1.6 cm and 1.7 cm. You would be expected to estimate the value between the lines in the 0.01 cm range. On an AP exam, reporting the length as 1.64 cm, 1.65 cm, or 1.66 cm would be a considered correct measurement.

Here is another point on the ruler.



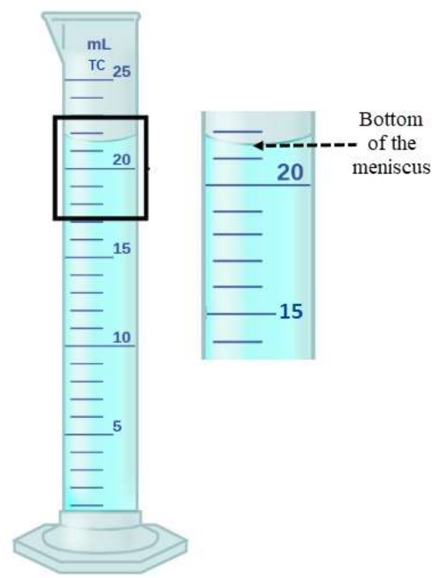
Reporting this measurement as 1.6 cm is incorrect. The measurement should be reported as 1.60 cm to indicate that the measurement's estimate includes the hundredth of a cm accuracy. When using a scale you should estimate the to 1/10 of the value of the smallest scale division.

Bottom line: Read and report between the lines!

Special rules used on the AP Exam involving graduated cylinders, burettes, and pipettes.

Graduate cylinders are designed to measure the volume of contents put in the container. Graduate cylinders are usually marked as TC (to contain). Water and many liquids will stick to the side of the glass and create a "lens" rather than a horizontal surface. The lens is called a meniscus and you are expected to measure from the lowest part of the meniscus.

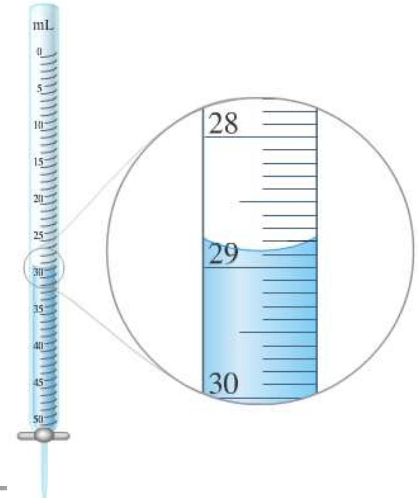
The graduations in this graduated cylinder are in 1 mL increments. So, you would be expected to estimate the volume to the next decimal place, 0.1 mL. An acceptable reading would be 21.4 mL, 21.5 mL, or 21.6 mL.



Burettes and pipettes are designed to pour out their contents and are designated as TD (to deliver). Burettes and pipettes are first filled and then the contents poured out. So that you can easily determine the amount of liquid dispensed, burette and pipette scale numbers increase from top to bottom.

In the case of this burette with graduations in 0.1 mL increments, you would be expected to estimate the 0.01 mL value (approximately 1 drop) which is typical for most burettes. This recorded measurement should be given as 28.89 mL.

While many labs are switching to digital pipettes and burettes, the AP Chemistry curriculum expects that students should be able to read burette scales. The use of burettes and pipettes will be revisited later in the course.



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.

WARNING! **DO NOT APPLY** THE MULTIPLICATION AND DIVISION SIGNIFICANT FIGURE RULE WHEN ADDING OR SUBTRACTING!

Addition and Subtraction

Significant figure counts are not used when adding and subtracting! Instead, the following four step procedure must be used to correctly display the accuracy of measurements.

1. Set up the problem in column form with the decimal points aligned.	2. Perform the calculation.	3. Since the values of the hundredths and thousandths of 2.0 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

Significant Figures and logarithmic values[9]

There are numerous instances where logarithmic values are used in chemistry. The accuracy of a logarithmic value is determined solely by the numbers to the right of the decimal point. A logarithm value of 12.05 would only have 2 significant figures. When converting a decimal value to a logarithmic value, the logarithm should be reported to the same number of decimal places as the number of significant figures in the decimal measurement.

e.g. The logarithm of 0.012 $\log_{10}(0.012) = -1.92$

[9] Logarithmic significant figure counts are not "tested" on the AP Chem exam. But still, you should report answers following the logarithm sig fig rules to be able to give reasonable answers on the exam.

Precision and Accuracy

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

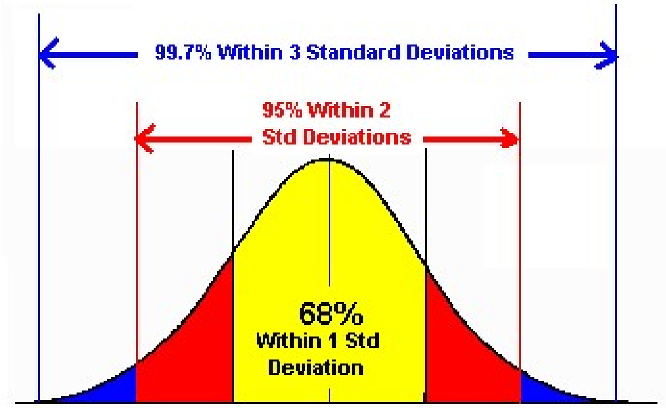
$$\% \text{ Error } = \frac{\text{experimental results - correct value}}{\text{correct value}} \times 100$$

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.

Standard deviation is not part of AP Chem, but it is useful in understanding experimental error and precision and accuracy.

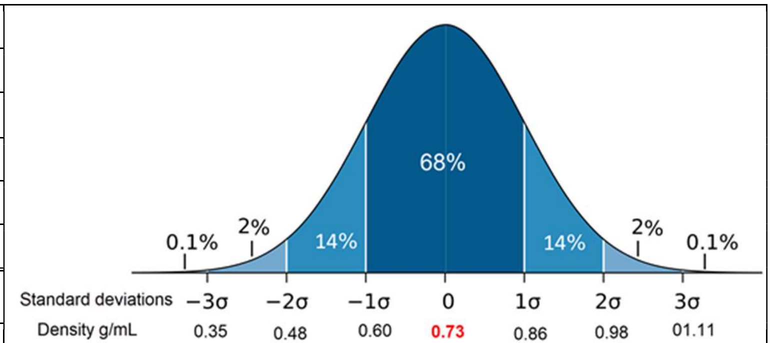
Standard deviation is a statistical value which shows the level of variance in a group of measurements and is an indicator of the level of precision. You are not expected to know how to calculate the standard deviation[10].

Typically, a set of measurements will show a range of values. The distribution of these measurements follows a pattern called a normal distribution curve. The center of the curve is the arithmetic mean (average) of the data. In a normal curve most of the data will be within 1 standard deviation of the average.



For example, the series of density measurements:

Trial 1	0.90	g/mL
Trial 2	0.79	g/mL
Trial 3	0.48	g/mL
Trial 4	0.73	g/mL
Trial 5	0.75	g/mL
Trial 6	0.74	g/mL
Average	0.73	g/mL
Standard Deviation	0.13	g/mL



[10]Standard Deviation calculator: https://www.mathportal.org/calculators/statistics-calculator/standard-deviation-calculator.php

The standard deviation of 0.13 g/mL indicates that 68% of the measurements fell within ±0.13 of the average. Of the six trials, Trial 3 looks dubious because it is more than 2 standard deviations from the average.

Caveat: While a small standard deviation indicates a high degree of precision in measurements, this does not guarantee accuracy. If there is a systemic error, the accuracy could be low despite the high precision. For example, in measuring density of a pure liquid, if the liquid had been contaminated, the density values would all be equally affected.

Summary:

> **The greater the number of significant figures in a measurement, the greater the precision and likely accuracy of that measurement.**
> **The % Error calculation determines the accuracy of a measurement.**
> **You should know how to do a % Error calculation.**

On WebAssign and on the AP Chemistry FRQ problems, it is generally expected that you will report your answers to 3 (or 4) sig figs. This will allow your answers to be within the accepted tolerances for grading for both WebAssign and the grading on the AP Chemistry test.

On WebAssign if you see the 4.0✓ icon, your answer will be expected to have the correct number of significant figures for the calculation and be within ± 1 on the last sig fig.

EXCEPTION TO THE 3-4 SIG FIG RULE!

> On the AP Chemistry exam **FRQ section**, **one question** will require the correct reporting of a value according to significant figure rules. However, that sig fig dependent question is not identified.

The significant figure graded question is almost always a simple calculation in a lab-based problem.

e.g.

Mass of thoroughly dried filter paper	12.46 g
Mass of filter paper + precipitate after drying	12.70 g

A precipitate (solid) produced on mixing two solutions. To collect the precipitate the resulting mixture is filtered, washed, and dried. The data from the experiment are shown in the table above. What is the mass of the precipitate?

12.70	g	mass of filter paper and solid
-12.46	g	mass of filter paper
0.24	g	mass of solid

This is a simple subtraction. There is no data that gives information on the thousandths of a gram value, thus the answer cannot be reported as 0.240 g. In this case following the 3-sig fig rule would cause the loss of 1 easy point.

Take special care with sig figs for easy numeric answers involving labs on the AP Chem exam.

Blindly using the three-sig fig generalization in lab calculations will definitely lead to a loss of a point on the AP Chemistry exam.

1.9 The Factor-Label Method of Solving Problems also called Dimensional Analysis

The Factor-Label Method relies on equivalences which are often shown as ratios.

e.g. The density of ethanol, a common alcohol, is 0.876 g/mL. This means that the mass of 1 mL of ethanol[11] is 0.876 g, or 0.876 g = 1 mL.

$$0.876\text{ g} = 1\text{ mL can be turned into } \frac{0.876\text{ g}}{1\text{ mL}} = 1 \text{ or } \frac{1\text{ mL}}{0.876\text{ g}} = 1$$

Thus, the density equivalence can be used to change a volume of ethanol into its corresponding mass, or a mass of ethanol into its volume.

(a) What is the mass of 56.0 mL of pure ethanol?

$$56.0\ \cancel{\text{mL}} \times \frac{0.876\text{ g}}{1\ \cancel{\text{mL}}} = 63.9\text{ g}$$

(b) What is the volume of 38.2 g of pure ethanol?

$$38.2\ \cancel{\text{g}} \times \frac{1\text{ mL}}{0.876\ \cancel{\text{g}}} = 43.6\text{ mL}$$

If you know how to use factor label equivalences, there is no need memorize the density equation.

I will point out the important equivalences for use in factor label conversions as we go through the course.

e.g.

1 kg = 1000 g

1 cm = 1 × 10^{-2} m.

Use the above equivalences to find the density of methanol, 0.808 g/cm^3, to kg/m^3.

When using factor labels, some students like to use a box method. Some like to use a fraction format.

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

You must show the method of solution on AP Free Response Question answers involving calculations.

If you only show a calculated answer, the AP Reader must assume you copied the answer from another student rather than calculating it yourself. **Shown calculations are proof of student work.**

When doing calculations, always include units with the numbers. Use units, all the time, in all the work.

The Factor-Label Method is the ideal way to solve most chemistry problems and show your work at the same time.

[11] The per denominator value, 1, in the ratio is a definition, and does not limit the significant figures of the ratio.

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When you do multiple step problems in WebAssign, do not round your calculator's intermediate answers.

Rounding done over a series of calculations leads to compounding rounding errors!
Rounding intermediate answers, may cause the final answer to 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. calculation for the number of molecules in 3.73 liters of gas at 1.07 atm of pressure and a temperature of 10.°C

$$\frac{1.07\ \text{atm} \times 3.73\ \text{L}}{0.0821\dfrac{\text{atm L}}{\text{mol K}} \times (10. + 273)\ \text{K}} \times \frac{6.02\times10^{23}\ \text{molecules}}{1\ \text{mol}}$$

$$\frac{1.07\ \cancel{\text{atm}} \times 3.73\ \cancel{\text{L}}}{0.0821\dfrac{\cancel{\text{atm}}\ \cancel{\text{L}}}{\cancel{\text{mol}}\ \cancel{\text{K}}} \times (10. + 273)\cancel{\text{K}}} \times \frac{6.02\times10^{23}\ \text{molecules}}{1\ \cancel{\text{mol}}}$$

As a single entry, on your calculator:	1.07 × 3.73 ÷ .0821 ÷ (10+273) × 6.02[EE]23 [ENTER] 1.034092785E23
Would be reported as	1.03 $\times10^{23}$ molecules

Extra non-AP topic on measurement and accuracy:

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 $\times10^{-38}$ m. Planck time: 5.391 $\times10^{-44}$ s

Please keep your lab measurements greater than these values lest you destroy the space time continuum.[12]

Salvador Dali's[13] *Melting Clock*

[12] “Out of the Fabric,” Tom Siegfried, Science News, April 23, 2011 page 29

[13] https://www.dalipaintings.com/images/paintings/the-melting-watch.jpg

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