Chapter 1 Chemistry: The Study of Change

No college chemistry textbook perfectly matches the AP Chemistry curriculum. My summaries will guide you through the topics in the text that you need to know for AP Chemistry. By using the summary, you will know which parts of the textbook you can skip. In this chapter you can skip three of the nine sections. Also do not do the problems in the textbook because they are not like AP Chemistry questions. Instead, the WebAssign problems will prepare you for the AP exam.

Keep each summary in your AP Chemistry binder.

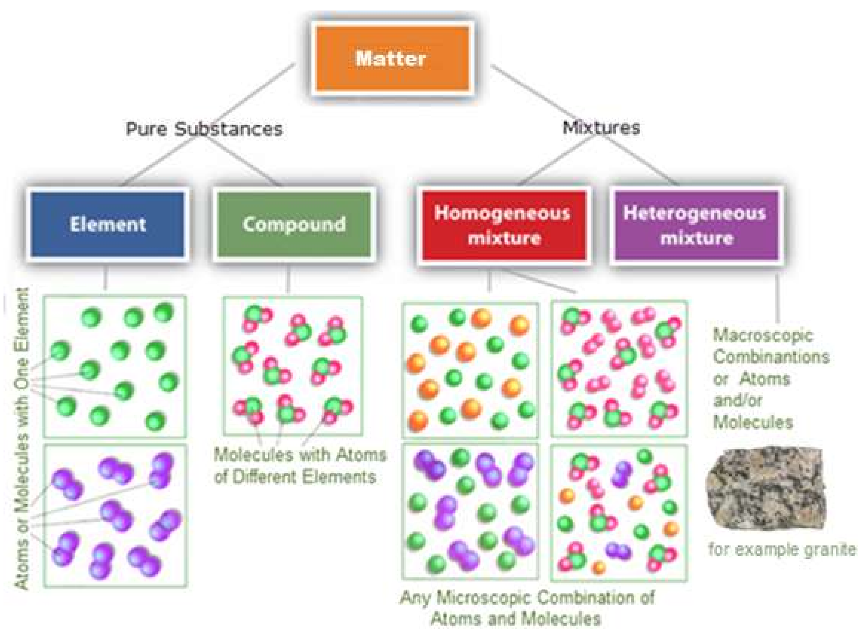
Chemistry by Chang and Goldsby

Chapter 1 – Chemistry: The Study of Change
~~1.01 Chemistry: A Science for the Twenty-First Century~~
~~1.02 The Study of Chemistry~~
~~1.03 The Scientific Method~~
1.04 Classifications of Matter
1.05 The Three States of Matter
1.06 Physical and Chemical Properties of Matter
1.07 Measurement
1.08 Handling Numbers
1.09 Dimensional Analysis in Solving Problems

1.1-1.3 Not in 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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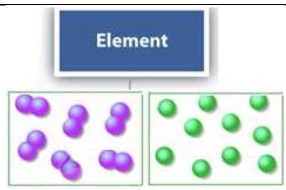
<td>
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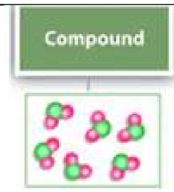
<td>
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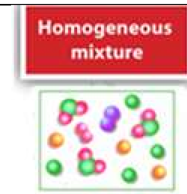
<td>
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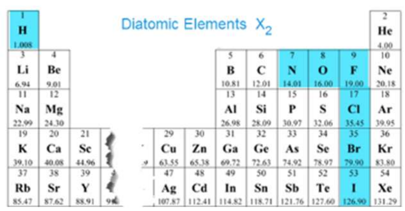
<td>There are only about 100 elements.

Atoms of an element may bond with one another to form molecules.
The most important elemental molecules are the diatomic elements.


</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 compounds are fixed and are represented by a formula indicating the ratios of atoms of each element.</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, and chromatography use the physical properties of each component to separate the components of a mixture.</td>
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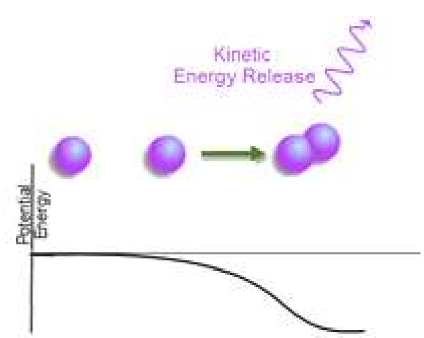
</table>

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. Single atoms of these elements will instantly chemically bond to form diatomic molecules releasing kinetic energy (light and heat). The resulting molecules have less potential energy and are more stable as molecules than as monatomic atoms. Bonded atoms have a lower potential energy than their original form..	

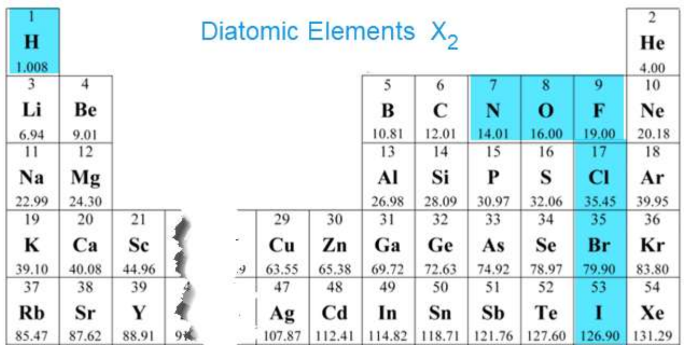
[1] Other elements bond to form elemental molecules. P_4 and S_8 are two examples. These molecular elements are not usually written as molecules in chemical equations. While you may encounter P_4 and S_8 on the AP Chemistry test, their formulas will be given so there is no need to memorize them.

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 molecules in equations is almost always incorrect.

In the few cases where the monatomic version is required, the element will be identified as atomic, for example atomic hydrogen would be H, while hydrogen would be H_2.

The diatomic substances can be easily found using their position on the periodic table.

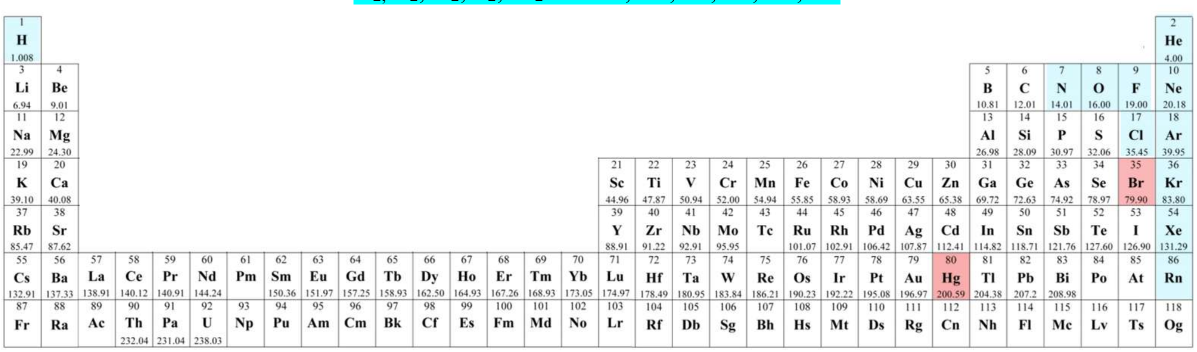


1.5 The Three States of Matter

Know the four abbreviations: (*s*) **pure** solid, (*l*) **pure** liquid, (*g*) **pure** gas, and in addition (*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, which are liquids, all of the rest of the elements are are solids at room temperature[2].

[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 of mixtures 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 magnesium to produce hydrogen gas and magnesium sulfate.

While you do not need to memorize the terms *extensive* and *intensive*, you should know the ideas behind the terms.

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

Macroscopic properties can be measured directly using simple lab equipment such as density which is measured from mass and volume of visible samples.

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

Units make an abstract number real and will help you understand chemistry and solve problems. Even if you do not understand a chemistry problem it may be possible to find 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	
$2x + 3x = 5x$	$2\text{ m} + 3\text{ m} = 5\text{ m}$
$2x \times 3x = 6x^2$	$2\text{ m} \times 3\text{ m} = 6\text{ m}^2$
$2x + 3y = ?$ 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 $3\text{ cm} \times \frac{0.01\text{ m}}{\text{cm}} = 0.03\text{ m}$ $2.00\text{ m} + 0.03\text{ m} = 2.03\text{ m}$

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

M	**mega**	10^6	
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 $\times 10^{-9}$ m.

The seven SI Base units[3] are the basis of measuring all phenomena in the universe.

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 an official SI unit is named after a person, the first letter of its symbol is capitalized, but when writing the full name of the unit, the first letter is not capitalized. For example, the unit for temperature, kelvin, has the symbol K, because the unit is named after William Thomson, who became Baron Kelvin of Largs, and was first British scientist to be elevated to the House of Lords.

SI Base Units

Base Quantity	Name of Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Electrical current	ampere	A
Temperature	kelvin	K
Amount of substance	mole	mol
Luminous intensity	candela	cd

The AP Chemistry Equation sheet has units and prefixes on its first page.

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

The mole is the amount of substance 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.

6.022×10^{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 6.022×10^{23} molecules of water, only weighs 18 grams.
One mole of uranium atoms, the heaviest primordial element, weighs about a half pound, 238 g.
One mole of a solid or liquid will be usually less than a small handful.
One mole of a gas at room temperature and pressure would be the amount gas in found in a small beachball.
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 constituent entities should be specified. For example, 0.500 mol of H_2O indicates that the 3.01×10^{23} particles that are being counted are H_2O molecules.

[3] The only base unit not used in AP Chemistry is the candela.

Mass and Weight[4]

The mass of a substance is constant.

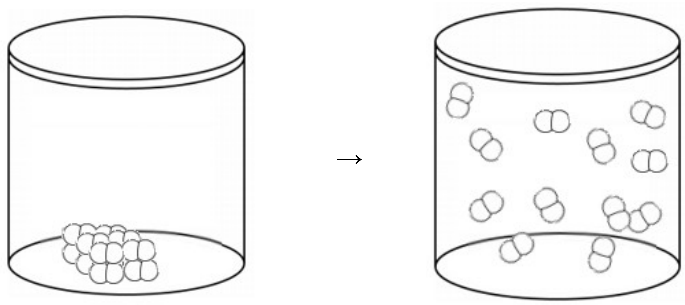
In chemistry we will measure the mass of a substance by its gravitational attraction to the earth using balances or scales.
The mass of any substance does not change with temperature.
Also, the mass of a substance does not change with the phase of a substance.

Mass is conserved in all physical and chemical reactions.

Physical Reaction

$I_2(s)$ → $I_2(g)$

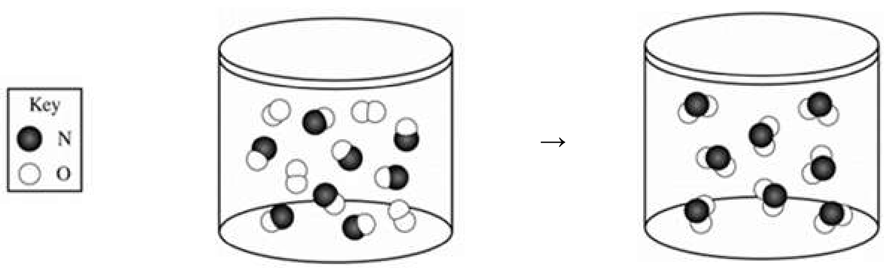
284 g   284 g



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 changes with temperature and phase.

The volume of a substance is determined by the spacing between the particles in the sample.

[4] 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. Since 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. Most scales measure mass by measuring the weighed objects attraction to earth. Gravitation attraction varies slightly depending on location. So extremely accurate scales are calibrated for their location.

The spacing of particles changes with temperature and phase.
Accurate measures of volume will also include the temperature associated with the volume measurement.
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.[5]

The mL and cm^3 are now defined as identical units.[6] 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.

Unit for Solid and Liquid densities: **g/mL** (or 1 g/cm^3)
Range from 0.2 g/mL to 20 g/mL.
Remember the **density of water** at room temperature is **1.0 g/mL**.

Unit Gas densities: **g/L**
The density of air at room temp and pressure[7] is about 1 g/L.

Before dealing with temperature, you need some critical background relating to temperature and energy.

The SI unit of Energy is joules.

Energy is a conserved, extensive property (like mass) and is defined as the ability to do work (apply a force over distance) or to heat a substance.
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).
Non-SI units are often used for energy such as the electron volt (1 eV = 1.6×10^{-19} joules), the kilowatt hour (1 kwh = 3.6×10^{6} joules), and the Calorie (1 Cal = 4,160 joules).
The AP Chemistry exam will only use the joule for its unit of energy.

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[5]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

[6] Prior to 1964, 1 mL = 1.000028 cm^3

[7] Typically, room temp and pressure is 25° and 1.00 atm. Of course, there is variation in "rooms".
If you live in La Rinconada, Peru (highest city in the world) typical room pressure would be 0.55 atm and the density of air would be 0.5 g/L

Potential (stored) energy is an important part of chemistry. Stored chemical energy is called enthalpy, and its symbol is *H*, and its typical unit is kJ.

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

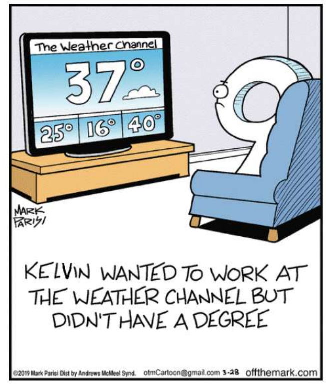
The SI unit of temperature is the kelvin, K.

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.

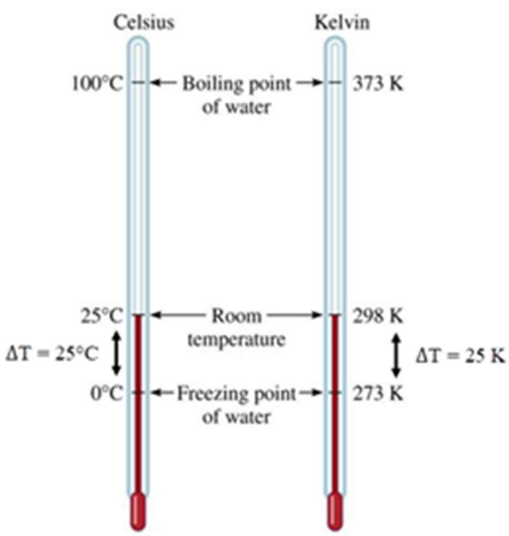
The kelvin scale is set up so that 0 K is the coldest possible temperature. There is no negative temperature just as there is no negative density. While commonly used, degrees Celsius is not a true measurement of average kinetic energy. Negative degrees Celsius do not really indicate a negative temperature.

Even though the equation for converting °C and K is on the AP Chemistry Equations Sheet, memorize how to convert[8] between °C and K: 273 + °C = K



Off the Mark by Mark Parisi [9]

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. $T_1 = 0.°C$ $T_2 = 25°C$ $\quad$ $T_1 = 273$ K $T_2 = 298$ K

$\Delta T = T_2 - T_1$ $\quad$ $\Delta T = T_2 - T_1$

$\Delta T = 25°C - 0.°C$ $\quad$ $\Delta T = 298\text{ K} - 273\text{ K}$

$\Delta T = +25°C$ $\quad$ $\Delta T = +25$ K

When dealing with ΔT, 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.

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

[9] *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

You will be allowed to use a calculator for both sections of the AP Chemistry exam, but you should be able to estimate answers without the use of a calculator.

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

e.g Find the product of 602×0.000786 without a calculator

simplifying numbers: $602 \approx 600$ $\quad 0.000786 \approx 0.0008$
scientific notation: $600 = 6\times10^{2}$ $\quad 0.0008 = 8\times10^{-4}$
$\approx 6\times10^{2} \times 8\times10^{-4}$
$\approx 6\times8 \times10^{(2\,+\,-4)}$
$\approx 48 \times10^{-2}$
$\approx$ 0.48… (answer using a calculator is 0.475)

Here is another example of simplification and scientific notation.

$$1.2\times10^{-8} = \frac{x^{2}}{(0.20-x)}$$ where x is less than 0.001

Since 0.001 is much less than 0.20 then $(0.20 - x) \approx 0.20$.
Eliminating the $-x$ from the equation makes solving the problem much easier.

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

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

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

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

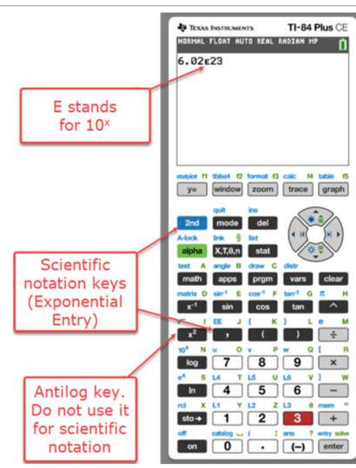
$$24\times10^{-10} = x^{2}$$

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

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

Most scientific calculators and WebAssign use E notation where the “E” is used as a substitute for 10^{x} in scientific notation.

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



Never use the multiplication key when typing in scientific notation on scientific or graphing calculators. [10]

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

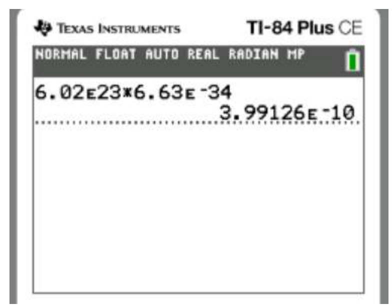
Try the next problem on your scientific calculator to see if you can use scientific notation on your calculator. The scientific notation in this problem will require the use of the scientific notation key on your calculator. If you type ×10 using the multiplication key and the number ten, you are not 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. Also you must use the negative value key (-) rather than the subtraction function key.

6.02 EE23 × 6.63 EE(-)34



The videos for this unit include a demonstration on how to use a Ti graphing calculator.

On entering scientific notation in WebAssign use E notation, where “E” stands for ×10.

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.

2 sig figs allow for ≈ 10% error
3 sig figs allow for ≈ 1% error
4 sig figs allow for ≈ 0.1% error

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.

[10] The Desmos online digital calculator is an exception to this rule.

Here are the rules to find the number of significant figures for a measurement in AP Chemistry.

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 (3 significant figures) 100 g (1 significant 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.

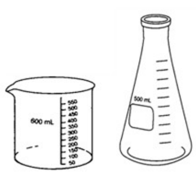
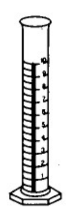
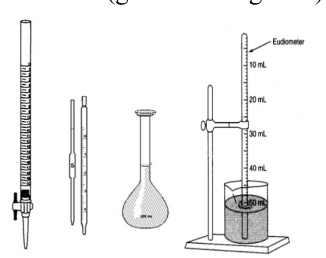
For example, rounding 297 to two significant figures would require the use of scientific notation since rounding to 300 would show 1 sig fig and 300. would show 3 sig figs. When in a situation like this, use scientific notation, 3.0×10^2.

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

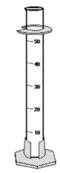
Significant Figures and Lab Measurements

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

Volume measurements in a lab often limit the accuracy of a measurement.

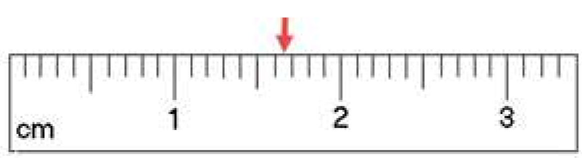
Least accurate	Intermediate accuracy	Highest accuracy
beakers and flasks 	graduated cylinders 	burettes, pipets, volumetric flasks, eudiometers (gas collecting tube) 

From least accurate to most accurate:

150 mL beaker (least accurate)  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 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 > 10 mL A grade A buret can dispense a 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.

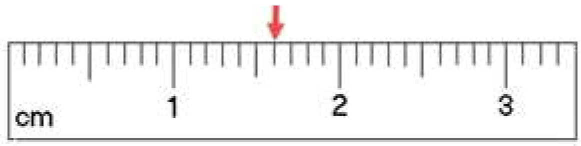
Many lab measurements are done using digital equipment, but some measurements involve interpreting scales. Measurements done with rulers and volumetric labware such as burettes, pipets, eudiometers, and graduated cylinders all require interpreting the accuracy of a scale.

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 is 0.1 cm. 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 considered as correct.

Here is another point on the ruler.

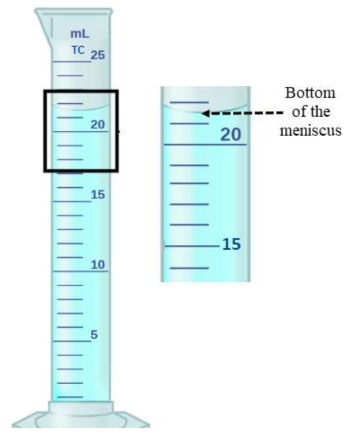


Reporting this measurement as 1.6 cm is incorrect. The measurement should be reported as 1.60 cm to indicate that the measurement's estimate includes the hundredth of a cm accuracy. When using a scale, you should estimate the to 1/10 of the value of the smallest scale division.

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.

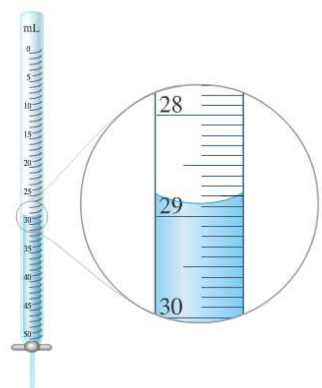


The graduations in this graduated cylinder are in 1 mL increments. So, you would be expected to estimate the volume to the next decimal place, 0.1 mL.

Read and report between the lines!

An acceptable reading would be 21.4 mL, 21.5 mL, or 21.6 mL.

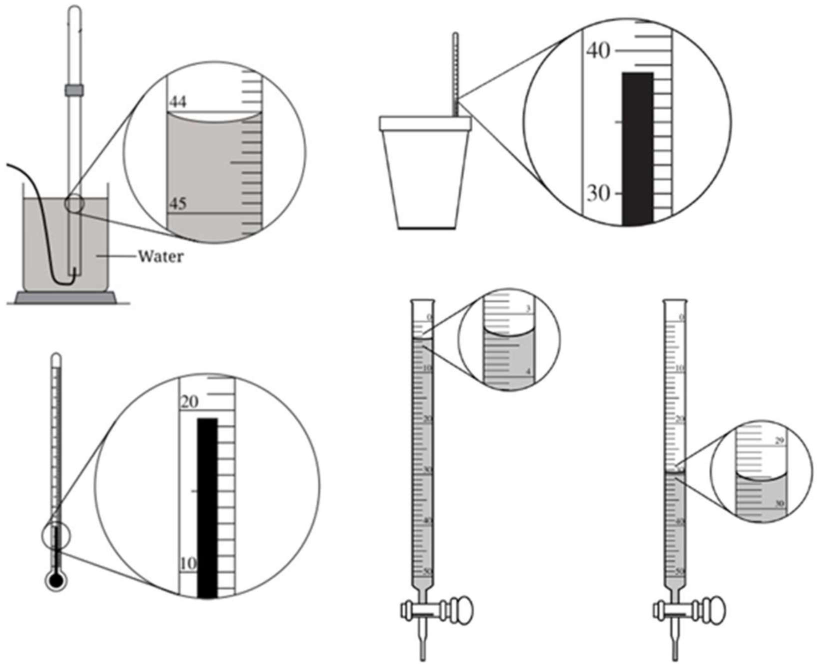
Burettes and pipettes are designed to pour out their contents and are designated as TD (to deliver). Burettes and pipettes are first filled and then the contents poured out. So that you can easily determine the amount of liquid dispensed, burette and pipette scale numbers increase from top to bottom.



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.

Here are examples of measurements that have been used on AP Exams.



Measurements: 44.10, 38.5, 19.5, 3.38, 29.57
(variations +/- 2 for last sig fig accepted)

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. While the 2.0 mL's tenth of an mL is known, the values of the hundredths and thousandths of 2.0 g are not certain, thus the addition or subtraction in those columns is not certain.	4. From right to left round your answer to the first complete column.
12.131 g - 2.0 g	12.131 g - 2.0 g 10.131 g	12.131 g - 2.0?? g 10.13~~1~~ g	12.131 g -2.0 g 10.1 g

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.

CAVEAT! When explaining the deviation from the correct answer on an experiment on any AP Chemistry test question **never say that the deviation was due human error!** That is an automatic loss of a point. Instead, you must work to find the experimental factor or procedure that produced that error.

The terms accuracy and precision are not part of AP Chem and will not be tested but the terms are often used in College Chemistry.

Accuracy of a measurement is an indication of how close the measurement is to the correct value.
Precision is used for a series of measurements and indicates variance within those measurements.

Summary:

The greater the number of significant figures measurements, the greater the precision and likely accuracy of measurements.

The % Error calculation determines the accuracy of a measurement.

When explaining error in an AP question NEVER USE HUMAN ERROR as a source of error in any FRQ.

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 Always found on the AP Chemistry Exam!

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 or 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) is 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 will undoubtedly result in a loss of one point on the AP Chemistry exam.

1.9 The Factor-Label Method of Solving Problems AKA Dimensional Analysis

The Factor-Label Method relies on equivalences which are 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.

$$0.876 \text{ g} = 1 \text{ mL can be written as } \frac{0.876 \text{ g}}{1 \text{ mL}} = 1 \text{ or } \frac{1 \text{ mL}}{0.876 \text{ g}} = 1$$

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?

$$56.0 \cancel{\text{ mL}} \times \frac{0.876 \text{ g}}{1 \cancel{\text{ mL}}} = 49.1 \text{ g}$$

(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.

Factor labels are the most effective way to convert prefixed units.

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

Box Method	Fraction method
$\begin{array}{\|c\|c\|c\|} \hline 0.808 \cancel{\text{g}} & 1 \text{ kg} & (1 \cancel{\text{cm}})^3 \\ \hline 1 \cancel{\text{cm}}^3 & 1000 \cancel{\text{g}} & (1 \times 10^{-2} \text{ m})^3 \\ \hline \end{array}$ = 808 kg/m^3	$\frac{0.808 \cancel{\text{ g}}}{1 \cancel{\text{ cm}}^3} \times \frac{1 \text{ kg}}{1000 \cancel{\text{ g}}} \times \frac{(1 \cancel{\text{ cm}})^3}{(1 \times 10^{-2} \text{m})^3} = 808 \text{ kg/m}^3$

You must show some indication of your work for a solution on AP Free Response Question answers involving calculations.

If you only show the answer for a problem requiring calculations, the AP Reader must assume you copied the answer from another student or had been given answers to the test. Even if the answer only requires a simple subtraction, the subtraction calculation must be shown for credit for the answer.

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

Your shown work doesn't have to be perfect but some indication of how the calculation is made is an absolute must.

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	1.03×10^{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: 1.616×10^{-38} m.

Planck time: 5.391×10^{-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*

This quote by William Thomson, Lord Kelvin is a good way to end this unit:

"When you can measure what you are speaking about, and express it in numbers, you know something about it, but when you cannot measure it, when you cannot express it in numbers, your knowledge is of a meagre and unsatisfactory kind."

[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