AP Chemistry
Summer Assignment 2013-2014

Welcome to AP Chemistry. Our mission this year is to understand Chemistry molecularly and to focus on understanding why Chemistry is the way it is rather than on memorizing different types of problems. You will learn to think analytically about the chemistry and be able to tackle any problem. That is the rigor of AP Chemistry.

As AP Chemistry students, this course is designed to be a rigorous college-level course that provides in-depth instruction in chemistry and in laboratory experiences. This course is demanding and will require your full commitment. It is expected that students spend at least 4-5 hours per week completing homework assignments and studying. Nothing can be forgotten in AP Chemistry, so a regular study session to refresh skills is very important. I would suggest forming a neighborhood study group to begin preparing. We will work diligently to prepare for the AP exam. It is my goal and expectation that each of you will receive a 3 or higher on the exam.

To prepare for the year ahead, this summer you will begin thoroughly reviewing your chemistry knowledge and reviewing the basics of how to speak "Chemistry". This assignment will cover all of the background material required. Please pay close attention to the assignment details to ensure all work is completed and turned in on-time. I have also attached a guide to writing in Chemistry. This should provide a clear expectation of the level of writing that is required for the course.

I'm very excited about working with each and everyone one of you to find great success in AP Chemistry. This course WILL be challenging, yet by the time you are ready for the exam, you will understand chemistry. If at any time you have a question regarding the assignments or the content itself, please feel free to email me at Justin.lambert@cmcss.net.

Mr. Justin W. Lambert
AP Chemistry Teacher

Summer Assignment:

1. Email me by June 1st. In this email please include your email address, extra-curricular and work commitments, your greatest fears in AP, and what your strength in Chemistry is.
2. We will be using a Moodle site to complete HW quizzes, share answers and provide important information. As soon as the website becomes functional, I will email each of the web address for the site. Please be watching your email for this.
3. Attached is a writing guide for AP. Please keep these guidelines in mind, as you write your lab reports next year.
4. From the textbook, read, take notes, and complete these homework problems:
    a. I will be collecting the HW when you return in August.
    b. Starting in August, I will have all the HW answers completed in a binder in my room for you to compare.
    c. This is designed to prepare you for the HW quizzes on Moodle.
    d. There will be a TEST over this material during the 1st week of school.

Chapter 1 - All	5, 7, 11, 14, 17, 18, 20, 22, 26, 31, 32, 33, 36, 38, 46, 50, 62, 64, 76, 100
Chapter 2 - All	1, 5, 12, 15, 16, 18, 24, 30, 36, 38, 44, 46, 47, 50, 58, 60, 62, 69, 73, 86, 88, 90, 94, 96, 100
Chapter 3 - All	4, 5, 6, 12, 14, 18, 20, 24,

	26, 28, 29, 30, 32, 34, 37, 40, 41, 42, 43, 44, 45, 46, 50, 52, 54, 58, 59, 60, 66, 68, 70, 71, 72, 76, 78, 84, 86, 89, 90, 94, 98, 100, 102, 103, 106, 117, 126, 132, 136, 140, 144, 146

5. It is also vital that you are able to SPEAK the language of Chemistry. There will be a test over this material in the first week of school. Please be sure to answer every question in these notes with work. This summer, you must memorize and know this information:
    a. Chemical Foundations Notes
    b. Atoms, Molecules, Ions Notes
    c. Stoichiometry Notes
6. By August 31st, on Moodle, you need to take the AP Chemistry Diagnostic Exam.

APPENDIX 2:

A BRIEF GUIDE TO WRITING IN CHEMISTRY

This document is a guide to assist students in chemistry courses with writing and formatting laboratory reports and research reports. An important goal of the Department is for our students to organize and communicate research results effectively and to write with acceptable scientific style. We hope that by providing many of the common stylistic, grammatical and organizational points in this single document, student can use this advice throughout their study of chemistry.

A. Formatting a Report

Layout. Use 12 point Times New Roman font and double spacing to allow space for comments and corrections. Number all pages, including those in appendices.

Organization. A standard lab report or research paper should be formatted with sections.

1. TITLE. List the title of the experiment or meaningful name for your research report. This is followed by your name and the date submitted. If you worked with partners, list their names next to yours, but put an asterisk after your name* to indicate that you wrote the report.

2. ABSTRACT. The abstract should be able to "stand alone." This means that someone should be able to read *only* your abstract and understand the basic nature of your report. For this reason, a good abstract clearly identifies the purpose of the experiment and the important results. Repeat: *a good abstract contains a summary of your results.*

Avoid pedagogical comments such as "this experiment helped us learn about the nature of chemical reactions" or "the goal of this experiment was to learn about dyes." Although those ARE important aspects and goals of the lab experience, the lab/research report should focus only on the data and results. Avoid starting your abstract with "The purpose of this experiment was..."

Background information on the theory or applications of your experiment belongs in the Introduction section. Avoid referencing any other sources or parts of the report, because the abstract should be able to "stand alone."

Be specific about what was done: name the reagents or types (not models) of instruments that were used, the products of a reaction, numerical values that were measured or calculated, etc. Avoid vague statements such as "a metal complex was prepared and the percent yield was calculated." A better abstract would read "hexaammine cobalt(III) chloride was prepared from cobalt (II) chloride, ammonia, ammonium chloride and hydrogen peroxide. The yield was 8.45 g (64 % based on cobalt)."

The best way to learn how to write a good abstract is to READ some published abstracts. These can be found in chemistry journals (for example, *The Journal of the American Chemical Society*) which are in the library.

TIP: When writing a full report, write the Abstract last.

3. INTRODUCTION. The introduction section explains to the reader what basic scientific question is being addressed. It includes general background material or a brief historical perspective on the topic being investigated. It presents brief summaries, with references, of previous work. An effective introduction funnels the reader from a larger area of research, through examples of progress in the field to a clear statement of the research problem or approach being addressed in the current report.

4. EXPERIMENTAL. This section includes a description of your experimental procedure, and names of instruments used. For lab courses, the procedure can simply reference the lab manual, listing any changes to the published procedure. DO NOT REWRITE THE LAB MANUAL. For advanced labs or independent research, the experimental section should provide all the necessary detail for someone to be able to reproduce your work. Often, an Experimental section is subdivided into **Materials** (sources and purity of reagents used), **Preparation of Compounds** (with procedure, and summary of characterization by NMR, IR, UV-Vis spectroscopy, melting point, chromatography, or elemental analysis) and **Instrumentation** (manufacturer, description of any adaptation or sample preparation) sections. Consult *JACS* to see examples of Experimental sections for various types of reports.

TIP: a good experimental section should allow another person, using what you have written and a lab manual, to completely reproduce what you did in the lab.

5. RESULTS & DISCUSSION. (may be single or separate sections) The Results should include a summary of your raw data (preferably in tabular form) and important observations. Do NOT include long tables of raw data; for those experiments simply present the results of your calculations. Calculations may be included in this section or in an Appendix, and a description of equations used in your calculations must be presented. Handwritten calculations are acceptable for lab reports.

A Discussion section should take the form of an analysis of your results. Comment on the purpose of the experiment. What do the results indicate? What are sources of error (experimental uncertainty/precision)? What additional experiments could help address any dangling ends? Do the results agree with what others have found? Do the results support a model or hypothesis? For some lab courses, you can use this section to answer any questions presented in the manual or in class. Although you should answer the questions in the lab manual, this section should have the style of flowing prose, not simply answers to numbered questions.

6. CONCLUSION. Summarize your results and discussion with a short conclusion that is more than simply a reiteration of your results. Phrase it in terms of the broader questions addressed in the Introduction.

7. REFERENCES. Citations of the literature used in the previous sections (see section F)

8. APPENDIX. Graphics may appear here, along with lengthy calculations or additional material not needed when reading through the report.

Graphics. Graphics include Tables, Figures, Schemes and chemical structures. Tables are columns of measured and/or calculated values or observations. All quantities should have units and be expressed using proper significant figures and scientific notation. Important experimental conditions should be listed as footnotes, especially when the table includes data obtained under different experimental conditions. Figures include: spectra, graphs, cartoons of experimental set-up or other drawings intended to show an *object.* Schemes include: reaction mechanisms, experimental flow charts or other drawings that are intended to show a *process*. All Tables, Figures and Schemes should be numbered sequentially and must be mentioned in the text. All graphics should be a full page in size and included at the end of the manuscript in the Appendix. Chemical structures can appear in the text and should be labeled with the same name, formula or compound number that appears in the text.

B. Sentence Structure and Writing Style

1. Beginning a sentence. Avoid beginning a sentence with a symbol, numeric value or equation.

incorrect: 315.6 mg of ammonium chloride was added to the solution, which was then heated to 50 °C.
correct: After the addition of 315.6 mg of ammonium chloride, the solution was heated to 50 °C.

incorrect: ν is both the vibrational frequency and the IR radiation frequency.
correct: The frequency ν refers to both the vibrational frequency and the frequency of IR radiation.

2. Dangling Modifiers and Illogical Construction. Check that a modifier phrase or the pronoun "it" actually refers to the intended subject. (see also: subject-verb agreement.)

incorrect: Being coated with grease, I cleaned the flask before adding reagents
was I coated with grease or was the flask?
correct: Because the flask was coated with grease, it was cleaned before...

incorrect: After transferring to a larger flask, the solution was heated to a boil.
did the solution transfer itself?
correct: The solution was transferred to a larger flask and heated to a boil.

incorrect: A diagram of the influenza virus is now available. To obtain it, contact the instructor. *The instructor is making the influenza virus available?*
correct: A diagram of the influenza virus is now available from the instructor.

<u>incorrect</u>: To prevent decomposition, the reaction flask must be purged of air.
does the flask want to prevent decomposition?
<u>correct</u>: To prevent decomposition, purge all air from the reaction flask.

3. Equations. Equations typically appear as a separate line from the text and are numbered sequentially throughout the manuscript. Equations can then be referred to by number.

<u>example:</u>

"The quenching rate constant can be calculated using the Stern-Volmer equation:

$$\Phi_0/\Phi_q = 1 + k_q\tau_0[Q] \qquad (2)$$

4. Hyphens. Hyphenate compound adjectives.

5-mL aliquots were added but, aliquots of 5 mL were added
crystal deposited from the slowly-cooled solution.

5. Spaces. There should be a space between a quantity and its units and between a quantity or word and subsequent parenthetical phrase.

6.626 J s
25.15 K = 298.15 °C
45 mL
456 nm (34,000 M^{-1} cm^{-1})

6. Personal Pronouns. By tradition, scientists avoid using the personal pronouns "I" and "we" and "you" in most technical communications. The use of third person instead of first person is preferred when reporting results. (see also: active voice)

first person: I heated the solution at 100 °C for 1 h. and I noticed that it turned blue.
third person: When heated at 100 °C for 1 h., the solution turned blue.

7. Pedagogical comments. Avoid including pedagogical comments in a report or scientific communication. Phrases such as "this experiment helped us learn about the nature of chemical reactions" or "the goal of this experiment was to learn about dyes" are addressing the process of learning not the science of the experiment. Although those ARE important aspects and goals of the lab experience, the lab report should focus only on the data and results.

Also, try to avoid starting your abstract with "The purpose of this experiment was..."

8. Personification. Molecules and equipment are not people, so do not personify them in your writing.

<u>incorrect:</u> Sugar really wants to dissolve in water.

correct: Sugar is very soluble in water.

incorrect: Sodium wants to lose one electron to form Na$^+$.
correct: Oxidation of Na to Na$^+$ is thermodynamically favorable.

incorrect: The spectrum shows two bands of equal intensity
correct: Two bands of equal intensity appear in the spectrum.

9. Plural nouns. "Data" is plural for "datum," "spectra" is plural for "spectrum," "phenomena" is plural for "phenomenon," and "formulae" is plural of "formula." The amount of chemical reagent is singular, so use the correct verb tense.

incorrect: Data was acquired and a spectra is in the appendix.
correct: Data were acquired and a spectrum is in the appendix

incorrect: While the solution boiled, 5.0 g of KBr were added.
correct: While the solution boiled, 5.0 g of KBr was added.

10. Prepositions. Don't forget "of" between quantities and substance name.

incorrect: "... and 10 mL MeOH was added."
correct: "... and 10 mL of MeOH was added."

11. Redundant or unnecessary phrases (pleonasm).

incorrect: A photon of light having a wavelength of 530 nm...
if not "of light," what was the photon made of?
correct: Light having a wavelength of 530 nm...

incorrect: In this experiment, aspirin was prepared from oil of wintergreen.
If not this experiment, then in which experiment?
correct: Aspirin was prepared from oil of wintergreen.

12. Subject-verb agreement. Are you stating that an inanimate object is drawing a conclusion, or suggesting a strange cause and effect? (see also: dangling modifers)

incorrect: The IR spectrum implies that water is in the aspirin sample. *(spectra don't imply, people do*)
correct: The presence of water in the aspirin sample is inferred from the IR spectrum.

incorrect: Water was present in the aspirin product because of the peak at 3200 cm^{-1} in the IR spectrum. (*the peak in the spectrum didn't cause water to be present*)
correct: The peak at 3200 cm^{-1} in the IR spectrum indicates that water was present in the aspirin product. (*water caused the peak in the spectrum)*

C. Verbs

1. Active voice. By avoiding personal pronouns, scientists often depend excessively on the passive voice, which can weaken the writing style. *When possible,* replace passive voice with active voice.

passive voice: A vapor was observed when the solution was heated.
active voice: A vapor formed above the hot solution.

passive voice: There was some solid that did not dissolve.
active voice: Some solid did not dissolve.

2. Subject-verb agreement. Based on whether the subject is singular or plural, use the correct verb tense. A quantity used is a singular subject, even when that quantity is in a plural form of units.

incorrect: 12 g **were** added
correct: 12 g **was** added

3. Verb tense. Past tense is used to describe a procedure that you followed in an experiment. Present tense is used to describe a scientific fact, such as the properties of a molecule.

examples: Hydrochloric acid was added to the flask slowly in order to prevent decomposition of the product. Hydrochloric acid is a caustic substance that must be used with caution.

4. "Verbing" a Noun. Don't turn nouns into verbs.

incorrect: ammonia complexes to cobalt ions
correct: ammonia forms complexes with cobalt ions.

incorrect: the mixture was centrifuged to separate the solid.
correct: The solid was separated from the mixture using a centrifuge.

incorrect: The solution was rotovapped to dryness
correct: The solvent was removed by rotary evaporation

D. Abbreviations, Formulae and Numerals

1. Standard Abbreviations. Use standard *JACS* abbreviations (note: not all journals use exactly the same abbreviations:

examples:
mL = milliliter; μg = microgram; nM = nanomolar
h = hour; min = minute; s = second
K = degrees Kelvin, °C = degrees Celsius

2. Chemical Formulae. Use subscripts, superscripts, parentheses, and symbols appropriately in chemical formulae.

examples: $Cr^{3+}(aq)$
$K_2[PtCl_4]$
$[Ru(bpy)_3{}^{2+}](PF_6)_2$

3. Compound Numbers. Compounds can be numbered if repeated long compounds names become cumbersome. The number should be defined (usually in bold or underlined) somewhere early in the manuscript, often when it is first presented. The numbers should appear in parenthesizes when used as adjectives, but not when used as nouns.

example:

"Investigations into the fluorescence of 8-hydroxyquinoline (**1**), 4-iodo-8-hydroxyquinoline (**2**) and 2-methyl-4-iodo-8-hydroxyquinoline (**3**) are described in this paper. Recrystallization of **1** and **2** afforded analytically pure samples, but vacuum sublimation of the methyl derivative (**3**) was necessary to remove fluorescent impurities."

4. Decimal Places. For values less than unity, use a leading zero. Avoid writing values having too many zero; use scientific notation.

examples:	"0.15 μL"	not ".15 μL"
	"2.3×10^{-5} M"	not "0.000024 M"

5. Defining Abbreviations. Abbreviations for chemical compounds, ligand, instruments or methods should be defined in the text before using throughout the manuscript.

examples:

"The complex cation $Ru(bpy)_3{}^{2+}$, where bpy = 2,2'-bipyridine, is luminescent . . ."
"Surfactants such as sodium dodecyl sulfate (SDS) lead to lower drag . . ."
"Peptide structures were minimized using the empirical force filed (EFF) method."
"The American Chemical Society (ACS) sponsors two annual national meetings."

6. Organic Abbreviations. Standard organic abbreviations can be used in text and formulae.

examples:

Me = methyl
Et = ethyl
iPr = *iso*-propyl
tBu = *tert*-butyl
Ch – cyclohexyl

7. Reagents and Solvents. Use chemical formulae for standard reagents and solvents, but not when the name is shorter or more precise examples:

NaOH (*aq*)	in place of "sodium hydroxide"
H_2SO_4 (*aq*)	in place of "sulfuric acid"
CH_2Cl_2	in place of "dichloromethane"
"caffeine"	in place of $C_8H_{10}N_4O_2$

E. Chemical Terms and Expressions

1. Chemical names. The names of chemicals are not capitalized, unless they are trade names such as "Tylenol" or "Viagra."

incorrect: The reaction of aqueous Cobalt(II) with Aspirin was investigated.
correct: The reaction of aqueous cobalt(II) with aspirin was investigated.

2. Create. Chemistry involves "synthesizing" new compounds, "preparing" solutions, "characterizing" products. Avoid using phrases such as "products were *created*." Too divine.

3. Measurements. Spectra are measured "with" or "using" a spectrometer, not "on" a spectrometer (ouch!)

4. Machines. Spectrometers (UV-Vis, IR, NMR, etc.) are "instruments," not "machines."

5. React. As an intransitive verb, "react" should not have an object and should not have a passive voice. Chemical reagents react with each other, they are not reacted.

incorrect: "Potassium hydroxide and hydrochloric acid were reacted to produce water and potassium chloride."
correct: "The reaction of potassium hydroxide and hydrochloric acid produced water and potassium chloride."

6. Tested. A hypothesis can be "tested" and a student can be "tested." For most laboratory work, the terms "measured," "investigated," "determined," "calculated" or "obtained" often work better.

incorrect: The absorbance of the solution was tested using the UV-vis machine.
correct: The absorbance of the solution was measured using a UV-vis spectrophotometer.

F. References

There are numerous styles for formatting references. Unless otherwise instructed, citations should be formatted in the *JACS* style and appear as endnotes. Alternatively, article titles can also be included. Most important is to prepare citations with a uniform style.

Last name, initials; Last name, initials *Journal Title* **year**, *volume (issue)*, starting page.
or
Last name, initials; Last name, initials "Article Title" *Journal Title* **year**, *volume (issue)*, starting page.

examples:

Schlabach, M.; Limbach, H.-H.; Shu, A.; Bunnenberg, E.; Tolf, B.; Djerassi, C. *J. Am. Chem. Soc.* **1993**, *115*, 4554.

Additional Materials for Writing Lab/Research Reports

Davis, Martha *Scientific papers and presentations* San Diego : Academic Press, **1997**

Dodd, Janet S. (ed.)*The ACS style guide : a manual for authors and editors* ACS, **1997**.

Eisenberg, Anne "Strategies five productive chemists use to handle the writing process." *J. Chem. Educ.* **1982,** *59,* 566.

Potera, Carol "The Basic Elements of Writing a Scientific Paper: The Art of Scientific Style" *J. Chem. Educ.* **1984,** *61,* 247.

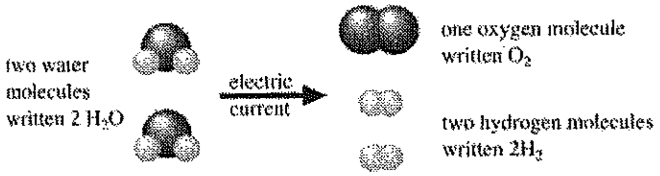
Spector, Thomas "Writing a Scientific Manuscript: Highlights for Success" *J. Chem. Educ.* **1994,** *71,* 47.

> "To avoid criticism, do nothing, say nothing, be nothing."
> *-Elbert Hubbard*

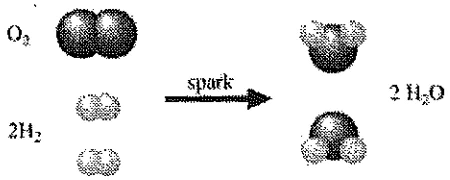
AP* Chemistry
CHEMICAL FOUNDATIONS

1.1 Chemistry: An Overview

- **<u>Matter</u>** – takes up space, has mass, exhibits inertia
  - composed of atoms only 100 or so different types
  - water made up of one oxygen and two hydrogen atoms
  - Pass an electric current through it to separate the two types of atoms and they rearrange to become two different types of molecules



- reactions are reversible

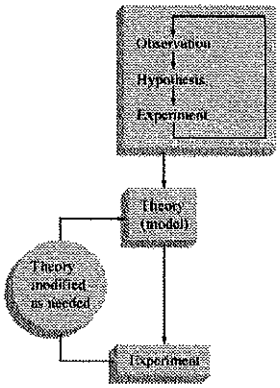


<u>Chemistry</u> – the study of matter and energy and more importantly, the changes between them

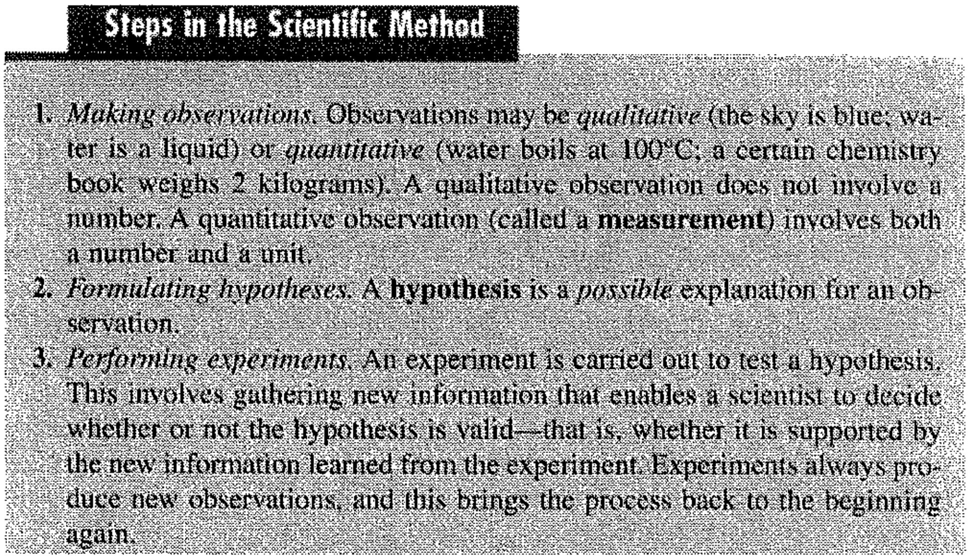
- **<u>Why study chemistry?</u>**
  - become a better problem solver in <u>all</u> areas of your life
  - safety – had the Roman's understood lead poisoning, their civilization would not have fallen
  - to better understand all areas of science

1.2 The Scientific Method

- A plan of attack!



The fundamental steps of the scientific method.



Steps in the Scientific Method

1. *Making observations.* Observations may be *qualitative* (the sky is blue; water is a liquid) or *quantitative* (water boils at 100°C; a certain chemistry book weighs 2 kilograms). A qualitative observation does not involve a number. A quantitative observation (called a **measurement**) involves both a number and a unit.
2. *Formulating hypotheses.* A **hypothesis** is a *possible* explanation for an observation.
3. *Performing experiments.* An experiment is carried out to test a hypothesis. This involves gathering new information that enables a scientist to decide whether or not the hypothesis is valid—that is, whether it is supported by the new information learned from the experiment. Experiments always produce new observations, and this brings the process back to the beginning again.

- Repetition of experiments is key

- **Theory** – hypotheses are assembled in an attempt at *explaining* "why" the "what" happened.
- **Model** – we use many models to explain natural phenomenon – when new evidence is found, the model changes!

Robert Boyle

- **Robert Boyle**
  - love to experiment with air
  - created the first vacuum pump
    - coin and feather fell at the same rate due to gravity in a vacuum since there is no air resistance.
  - $P_1V_1 = P_2V_2$
  - defined elements as anything that cannot be broken down into simpler substances

- **Scientific Laws** – a summary of observed (measurable) behavior [a theory is an explanation of behavior]

  A law summarizes what happens; a theory (model) is an attempt to explain WHY it happens.
  - **Law of Conservation of Mass** – mass reactants = mass products
  - **Law of Conservation of Energy** – (a.k.a. first law of thermodynamics)

    Energy CANNOT be created NOR destroyed; can only change forms.
  - scientists are human and subjected to
    - Data misinterpretations
    - Emotional attachments to theories
    - Loss of objectivity
    - Politics
    - Ego
    - Profit motives
    - Fads
    - Wars
    - Religious beliefs
- **Galileo** – forced to recant his astronomical observations in the face of strong religious resistance.
- **Lavoisier** – "father of modern chemistry"; beheaded due to political affiliations.
- The need for better explosives; (rapid change of solid or liquid to gas where molecules become ≈2,000 diameters farther apart and exert massive forces as a result) for wars have led to
  - fertilizers that utilizes nitrogen
  - nuclear devices

1.3 Units of Measurement

A quantitative observation, or measurement, ALWAYS consists of two parts: a *number* and a *unit*. Two major measurements systems exist: English (US and some of Africa) and Metric (the rest of the globe!)

- SI system – 1960 an international agreement was reached to set up a system of units so scientists everywhere could better communicate measurements. Le Système International in French; all based upon or derived from the metric system

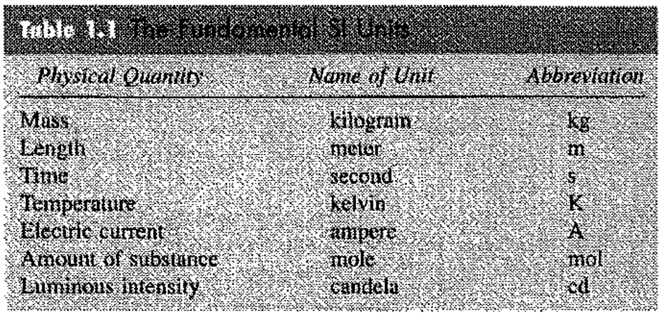


Table 1.1 The Fundamental SI Units

Physical Quantity	*Name of Unit*	*Abbreviation*
Mass	kilogram	kg
Length	meter	m
Time	second	s
Temperature	kelvin	K
Electric current	ampere	A
Amount of substance	mole	mol
Luminous intensity	candela	cd

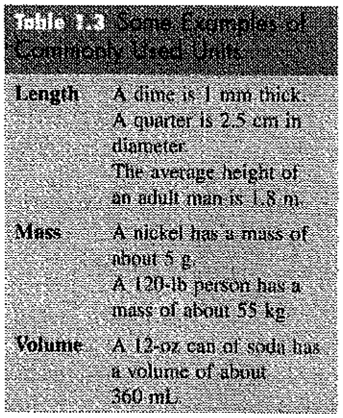


Table 1.3 Some Examples of Commonly Used Units

Length	A dime is 1 mm thick. A quarter is 2.5 cm in diameter. The average height of an adult man is 1.8 m.
Mass	A nickel has a mass of about 5 g. A 120-lb person has a mass of about 55 kg.
Volume	A 12-oz can of soda has a volume of about 360 mL.

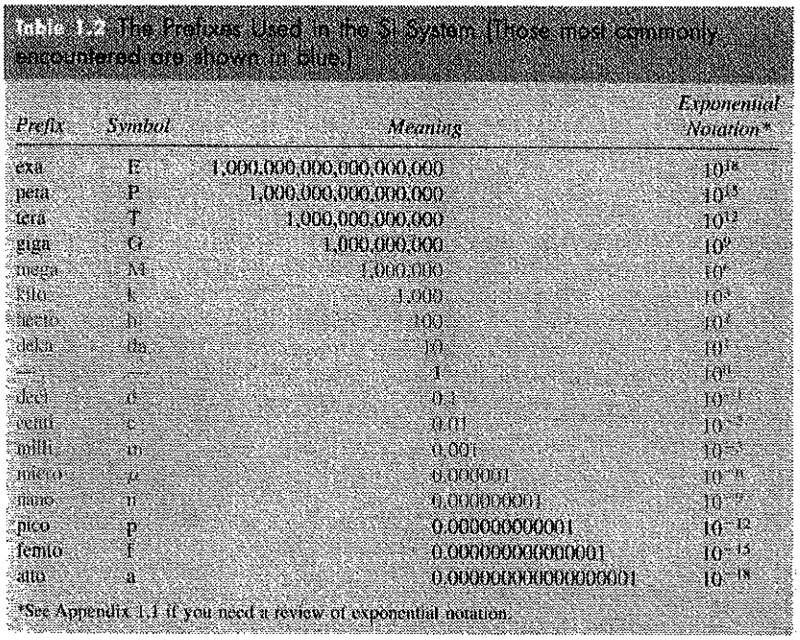
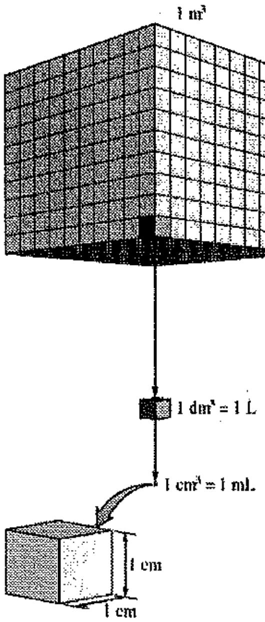


Table 1.2 The Prefixes Used in the SI System (Those most commonly encountered are shown in blue.)

Prefix	*Symbol*	*Meaning*	*Exponential Notation**
exa	E	1,000,000,000,000,000,000	10^{18}
peta	P	1,000,000,000,000,000	10^{15}
tera	T	1,000,000,000,000	10^{12}
giga	G	1,000,000,000	10^{9}
mega	M	1,000,000	10^{6}
kilo	k	1,000	10^{3}
hecto	h	100	10^{2}
deka	da	10	10^{1}
—	—	1	10^{0}
deci	d	0.1	10^{-1}
centi	c	0.01	10^{-2}
milli	m	0.001	10^{-3}
micro	μ	0.000001	10^{-6}
nano	n	0.000000001	10^{-9}
pico	p	0.000000000001	10^{-12}
femto	f	0.000000000000001	10^{-15}
atto	a	0.000000000000000001	10^{-18}

*See Appendix 1.1 if you need a review of exponential notation.

KNOW THESE UNITS AND PREFIXES!!!



- <u>Volume</u> – derived from length

  consider a cube 1m on each edge ∴ 1.0 m^3

  - a decimeter is 1/10 of a meter so

  $(1m)^3 = (10dm)^3 = 1{,}000\ dm^3$

  $1dm^3$ = 1 liter (L) and is slightly larger than a quart also

  $1dm^3 = 1\ L = (10\ cm)^3 = 1{,}000\ cm = 1{,}000\ mL$

  SINCE

  $1\ cm^3 = 1\ mL = 1$ gram of H_2O (at 4°C if you want to be picky)

- <u>Mass vs. Weight</u> – chemists are quite guilty of using these terms interchangeably.

  - <u>mass</u> (g or kg) – a measure of the resistance of an object to a change in its state of motion (ie. exhibits inertia); the quantity of matter present.

  - <u>weight</u> (a force ∴ Newtons) – the response of mass to gravity; since all of our measurements will be made here on Earth, considered the acceleration due to gravity a constant so we'll use the terms interchangeably as well *although* it is incorrect.

We "weigh" chemical quantities on a **balance <u>NOT</u> a scale!!**

Physics moment:

$F_W = ma$
$F_W = mg$

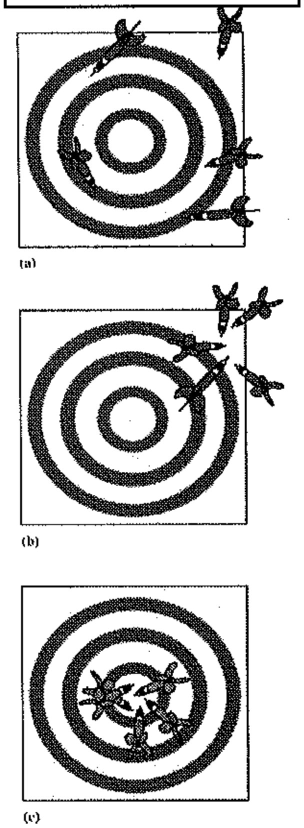
$$F_w = m\left(\frac{9.8m}{s^2}\right)$$

∴ its units are

$$N = (kg)\left(\frac{m}{s^2}\right)$$

- **gravity** – varies with altitude here on planet Earth
  - the closer you are to the center of the Earth, the stronger the gravitational field SINCE it originates from the center of the Earth.
  - Every object has a gravitational field – as long as you're on Earth, they are masked since the Earth's field is so HUGE compared to the object's.
  - The strength of the gravitational field $\propto$ mass
  - Ever seen astronauts in space that are "weightless" since they are very far removed from the center of Earth? Notice how they are constantly "drawn" to the sides of the ship and must push away?
  - The ships' mass is greater than the astronaut's mass ∴ "g" is greater for the ship and the astronaut is attracted to the ship just as you are attracted to Earth!
  - The moon has $\frac{1}{6}$ the mass of the Earth ∴ you would experience $\frac{1}{6}$ the gravitational field you experience on Earth and ∴ you'd WEIGH $\frac{1}{6}$ of what you weigh on Earth.

- **Precision and Accuracy**
  - **accuracy** – correctness; agreement of a measurement with the true value
  - **precision** – reproducibility; degree of agreement among several measurements.
  - **random or indeterminate error** – equal probability of a measurement being high or low
  - **systematic or determinate error** – occurs in the same direction each time



The results of several dart throws show the difference between precise and accurate. (a) Neither accurate nor precise (large random errors). (b) Precise but not accurate (small random errors, large systematic error). (c) Bull's-eye! Both precise and accurate (small random errors, no systematic error).

Exercise 1.2 Precision and Accuracy

To check the accuracy of a graduated cylinder, a student filled the cylinder to the 25-mL mark using water delivered from a buret and then read the volume delivered. Following are the results of five trials:

Trial	*Volume Shown by Graduated Cylinder*	*Volume Shown by the Buret*
1	25 mL	26.54 mL
2	25 mL	26.51 mL
3	25 mL	26.60 mL
4	25 mL	26.49 mL
5	25 mL	26.57 mL
Average	*25 mL*	*26.54 mL*

Is the graduated cylinder accurate?

Note that the average value measured using the buret is significantly different from 25 mL. Thus this graduated cylinder is not very accurate. It produces a systematic error (in this case, the indicated result is low for each measurement).

1.5 Significant Figures and Calculations

Rules

- Non zero digits are significant
- A zero is significant if it is
    - "terminating AND right" of the decimal [must be both]
    - "sandwiched" between significant figures
- Exact or counting numbers have an ∞ amount of significant figures as do fundamental constants

Exercise 1.3 **Significant Figures**

Give the number of significant figures for each of the following results.

a. A student's extraction procedure on tea yields 0.0105 g of caffeine.
b. A chemist records a mass of 0.050080 g in an analysis.
c. In an experiment, a span of time is determined to be 8.050×10^{-3} s .

a. three
b. five
c. four

Rules for calculating

- × and ÷ The term with the least number of *significant figures* (∴ least accurate measurement) determines the number of significant figures in the answer.

$$4.56 \times \underline{1}.\underline{4} = 6.38 \xrightarrow{\text{corrected}} \underline{6}.\underline{4}$$

- + and (−) The term with the least number of *decimal places* (∴ least accurate measurement) determines the number of significant figures in the answer.

$$\begin{array}{ll} 12.11 & \\ 18.0 & \leftarrow \text{limiting term} \\ \underline{1.013} & \\ 31.123 & \xrightarrow{\text{corrected}} 31.1 \end{array}$$

- pH – the *number of significant figures in least accurate measurement* determines *number decimal places* on the reported pH

Rounding Rules:

- Round at the end of all calculations
- Look at the significant figure one place beyond your desired number of significant figures if > 5 round up; < 5 drop the digit.
- Don't "double round" 4.348 to 2 SF = 4.3 NOT the 8 makes the 4 a 5 then 4.4. [Even though you may have conned an English teacher into this before!]

1.6 Dimensional Analysis

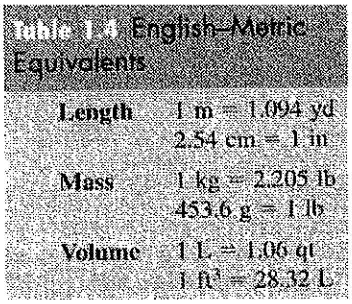


Table 1.4 English–Metric Equivalents

Length	1 m = 1.094 yd
	2.54 cm = 1 in
Mass	1 kg = 2.205 lb
	453.6 g = 1 lb
Volume	1 L = 1.06 qt
	1 ft^3 = 28.32 L

Consider a pin measuring 2.85 cm in length.
What is its length in inches?

2.54 cm = 1 inch $\therefore$ you can write 2 Conversion factors: $\frac{1\text{ in}}{2.54\text{ cm}}$ or $\frac{2.54\text{ cm}}{1\text{ in}}$

To convert multiply your quantity by a conversion factor that "cancels" the undesirable unit and puts the desired unit in the numerator.

$$2.85\ \cancel{\text{cm}} \times \frac{1\text{ in}}{2.54\ \cancel{\text{cm}}} = 1.12\text{ in}$$

Exercise 1.5 Unit Conversions I

A pencil is 7.00 in. long. What is its length in centimeters?

17.8 cm

Exercise 1.6 Unit Conversions II

You want to order a bicycle with a 25.5-in. frame, but the sizes in the catalog are given only in centimeters. What size should you order?

64.8 in

Exercise 1.7 Unit Conversions III

A student has entered a 10.0-km run. How long is the run in miles?

We have kilometers, which we want to change to miles. We can do this by the following route:

kilometers → meters → yards → miles

To proceed in this way, we need the following equivalence statements:

1 km = 1000 m
1 m = 1.094 yd
1760 yd = 1 mi

= 6.22 mi

Exercise 1.8 Unit Conversions IV
The speed limit on many highways in the United States is 55 mi/h. What number would be posted in kilometers per hour?

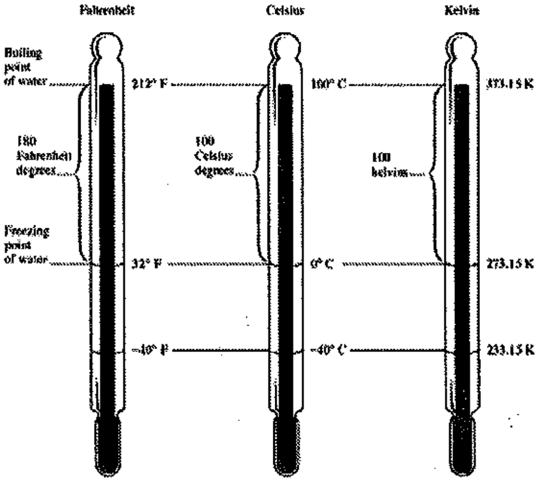
88 km / h

Exercise 1.9 Unit Conversions V
A Japanese car is advertised as having a gas mileage of 15 km/L. Convert this rating to miles per gallon.

35 mi / gal

1.7 Temperature

Three Scales



Notice a degree C = a degree K

$$T_F = T_C \times \frac{9°\text{F}}{5°\text{C}} + 32°\text{F}$$

$$T_K = T_C + 273.15\ \text{K}$$
$$T_C = T_K - 273.15\ °\text{C}$$

Exercise 1.10 Temperature Conversions I
Normal body temperature is 98.6°F. Convert this temperature to the Celsius and Kelvin scales.

98.6°F = 37.0°C
98.6°F = 310.2 K

Exercise 1.11 **Temperature Conversions II**
One interesting feature of the Celsius and Fahrenheit scales is that –40°C and –40°F represent the same temperature. Verify that this is true.

Exercise 1.12 **Temperature Conversions III**
Liquid nitrogen, which is often used as a coolant for low-temperature experiments, has a boiling point of 77 K. What is this temperature on the Fahrenheit scale?

T_F = –319 ° F

1.8 Density

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

Exercise 1.13 **Determining Density**
A chemist, trying to identify the main component of a compact disc cleaning fluid, finds that 25.00 cm^3 of the substance has a mass of 19.625 g at 20°C. The following are the names and densities of the compounds that might be the main component.

Compound	Density $\left(\frac{g}{cm^3}\right)$ at 20°C
Chloroform	1.492
Diethyl ether	0.714
Ethanol	0.789
Isopropyl alcohol	0.785
Toluene	0.867

Which of these compounds is the most likely to be the main component of the compact disc cleaner?

Density = 0.7850 g / cm^3 ∴ isopropyl alcohol

1.9 Classification of Matter

States of Matter

- **solid** – rigid; definite shape and volume; *molecules close together vibrating about fixed points ∴ virtually incompressible*
- **liquid** – definite volume but takes on the shape of the container; *molecules still vibrate but also have rotational and translational motion and can slide past one another BUT are still close together ∴ slightly compressible*
- **gas** – no definite volume and takes on the shape of the container; *molecules vibrate, rotate and translate and are independent of each other ∴ VERY far apart ∴ highly compressible*
  - **vapor** – the gas phase of a substance that is normally a solid or liquid at room temperature
  - **fluid** – that which can flow; gases and liquids

- **Mixtures** – can be **physically** separated
  - **homogeneous** – have visibly indistinguishable parts, solutions including air
  - **heterogeneous** – have visibly distinguishable parts
  - means of physical separation include: filtering, fractional crystallization, distillation, chromatography

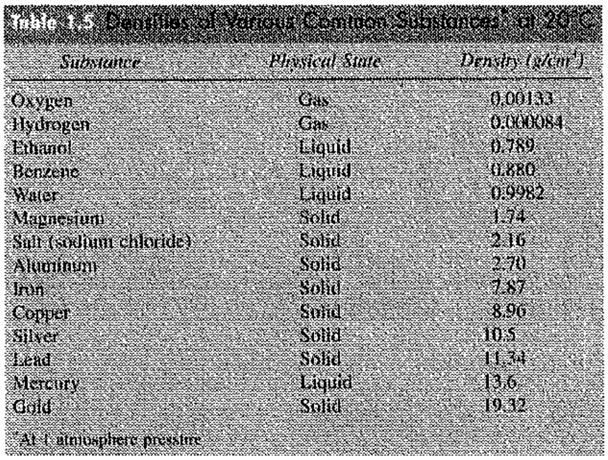
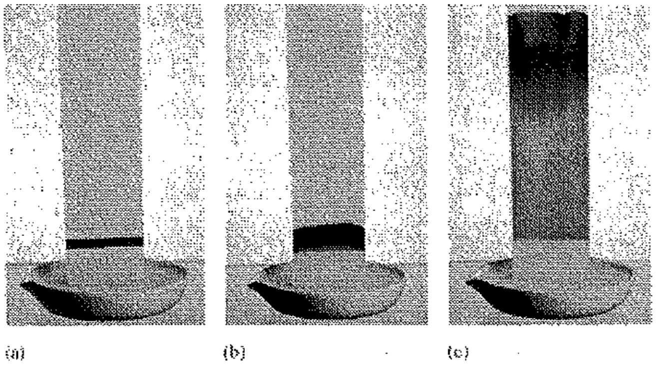


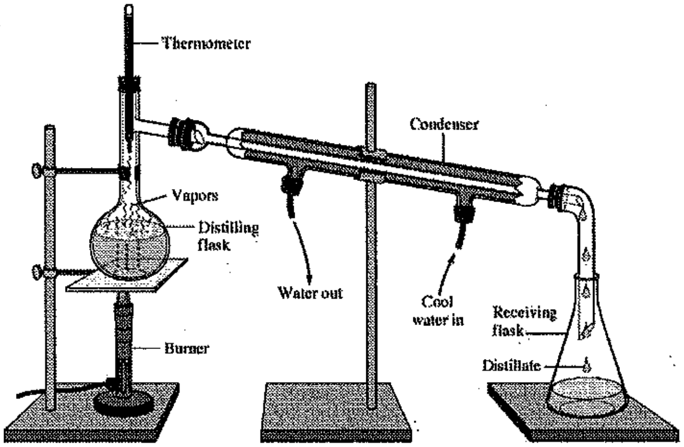
Table 1.5 Densities of Various Common Substances* at 20°C

Substance	*Physical State*	*Density* (g/cm^3)
Oxygen	Gas	0.00133
Hydrogen	Gas	0.000084
Ethanol	Liquid	0.789
Benzene	Liquid	0.880
Water	Liquid	0.9982
Magnesium	Solid	1.74
Salt (sodium chloride)	Solid	2.16
Aluminum	Solid	2.70
Iron	Solid	7.87
Copper	Solid	8.96
Silver	Solid	10.5
Lead	Solid	11.34
Mercury	Liquid	13.6
Gold	Solid	19.32

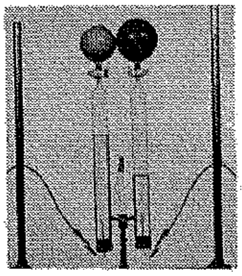
*At 1 atmosphere pressure



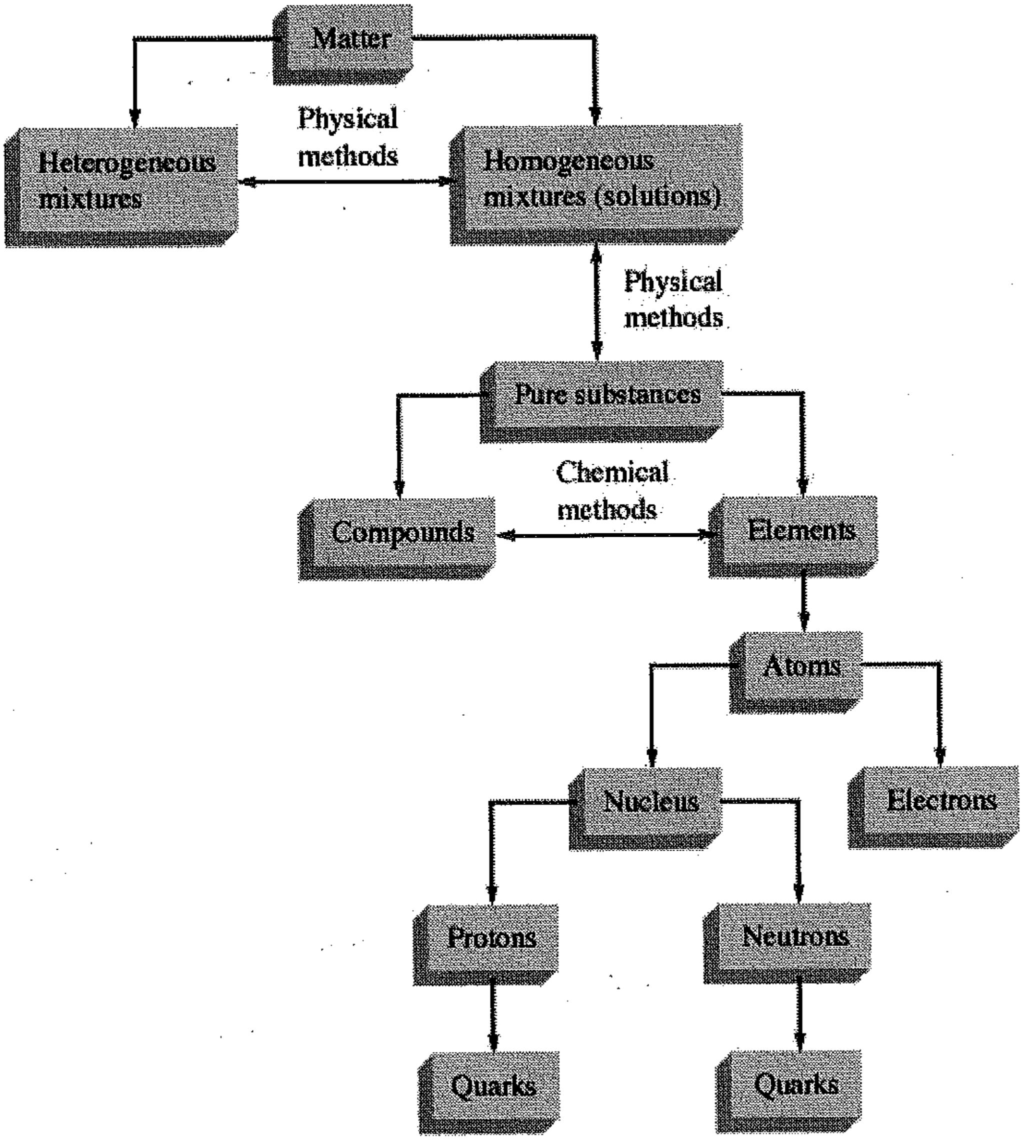
Paper chromatograph of ink. (a) A line of the mixture to be separate Is placed at one end of a sheet of porous paper. (b) The paper acts as a wick to draw up the liquid. (c) The component with the weakest attraction for the paper travels faster than those that cling to the paper.



- **<u>Pure substances</u>** – compounds like water, carbon dioxide etc. <u>and</u> elements. Compounds can be separated into elements by **chemical** means
  - electrolysis is a common chemical method for separating compounds into elements.
  - elements can be broken down into atoms
    - which can be broken down into
      - nuclei and electrons
      - p^+, n^0 and e^-
      - quarks



Electrolysis is an example of a chemical change. In this apparatus, water is decomposed to hydrogen gas (filling the red balloon) and Oxygen gas (filling the blue balloon).



AP* Chemistry
ATOMS, MOLECULES & IONS

The highest honor given by the American Chemical Society. He discovered oxygen. Ben Franklin got him interested in electricity and he observed graphite conducts an electric current. Politics forced him out of England and he died in the US in 1804. The back side, pictured below was given to Linus Pauling in 1984. Pauling was the only person to win Nobel Prizes in TWO Different fields: Chemistry and Peace.

2.1 THE EARLY HISTORY OF CHEMISTRY

- 1,000 B.C.—processing of ores to produce metals for weapons and ornaments; use of embalming fluids
- 400 B.C.—Greeks—proposed all matter was make up of 4 "elements" : fire, earth, water and air
- Democritus—first to use the term *atomos* to describe the ultimate, smallest particles of matter
- Next 2,000 years—*alchemy*—a pseudoscience where people thought they could turn metals into gold. Some good chemistry came from their efforts—lots of mistakes were made!
- 16th century—Georg Bauer, German , refined the process of extracting metals from ores & Paracelsus, Swiss, used minerals for medicinal applications
- Robert Boyle, English—first "chemist" to perform **quantitative** experiments of pressure versus volume. Developed a working definition for "elements".
- 17th & 18th Centuries—Georg Stahl, German—suggested "phlogiston" flowed OUT of burning material. An object stopped burning in a closed container since the air was "saturated with phlogiston"
- Joseph Priestley, English—discovered oxygen which was originally called "dephlogisticated air"

2.2 FUNDAMENTAL CHEMICAL LAWS

- late 18th Century—Combustion studied extensively
- CO_2, N_2, H_2 and O_2 discovered
- list of elements continued to grow
- Antione Lavoisier, French—explained the true nature of combustion—published the first modern chemistry textbook AND stated the Law of Conservation of Mass. The French Revolution broke out the same year his text was published. He once collected taxes for the government and was executed with a guillotine as an enemy of the people in 1794. He was the first to insist on *quantitative* experimentation.

THE LAW OF CONSERVATION OF MASS:

Mass is neither created nor destroyed.Joseph Proust, French—stated the Law of Definite Proportions

THE LAW OF DEFINITE PROPORTIONS:

A given compound always contains exactly the same proportions of elements by mass.

- 1808--John Dalton stated the Law of Definite proportions. He then went on to develop the Atomic Theory of Matter.

THE LAW OF MULTIPLE PROPORTIONS:

When two elements combine to form a series of compounds, the ratios of the masses of the second element that combine with 1 gram of the first element can always be reduced to small whole numbers.

Dalton considered compounds of carbon and oxygen and found:

	Mass of Oxygen that combines with 1 gram of C
Compound I	1.33 g
Compound II	2.66 g

Therefore Compound I may be CO while Compound II may be CO_2.

Exercise 2.1 Illustrating the Law of Multiple Proportions

The following data were collected for several compounds of nitrogen and oxygen:

Mass of Nitrogen That Combines With 1 g of Oxygen

Compound A	1.750 g
Compound B	0.8750 g
Compound C	0.4375 g

Show how these data illustrate the law of multiple proportions.

$$\frac{A}{B} = \frac{1.750}{0.875} = \frac{2}{1}$$

$$\frac{B}{C} = \frac{0.875}{0.4375} = \frac{2}{1}$$

$$\frac{A}{C} = \frac{1.750}{0.4375} = \frac{4}{1}$$

2.3 DALTON'S ATOMIC THEORY

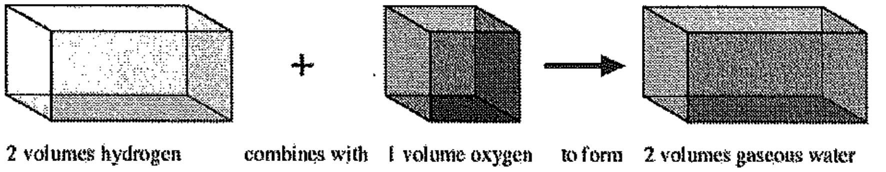
Dalton's ATOMIC THEORY OF MATTER: (based on knowledge *at that time):*

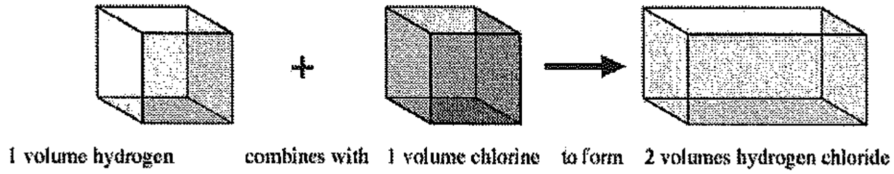
1. All matter is made of **atoms**. These *indivisible and indestructible objects* are the ultimate chemical particles.
2. All the atoms of a given element are identical, in both weight and chemical properties. However, atoms of different elements have different weights and different chemical properties.
3. **Compounds** are formed by the combination of different atoms in the ratio of small whole numbers.
4. A **chemical reaction** involves only the combination, separation, or rearrangement of atoms; atoms are neither created nor destroyed in the course of ordinary chemical reactions.

**TWO MODIFICATIONS HAVE BEEN MADE TO DALTON'S THEORY

1. *Subatomic particles were discovered.*
2. *Isotopes were discovered.*

- 1809 Joseph Gay-Lussac, French—performed experiments [at constant temperature and pressure] and measured volumes of gases that reacted with each other.

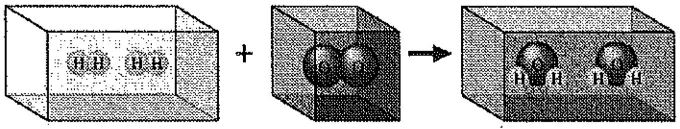


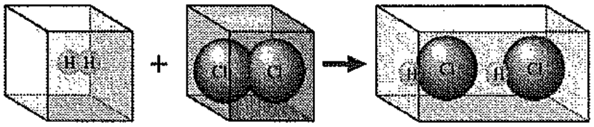


- 1811 Avogardro, Italian—proposed his hypothesis regarding Gay-Lussac's work [and you thought he was just famous for 6.02 x 10^{23}] He was basically ignored, so 50 years of confusion followed.

AVOGADRO'S HYPOTHESIS:

At the same temperature and pressure, equal volumes of <u>different</u> gases contain the same number of particles.



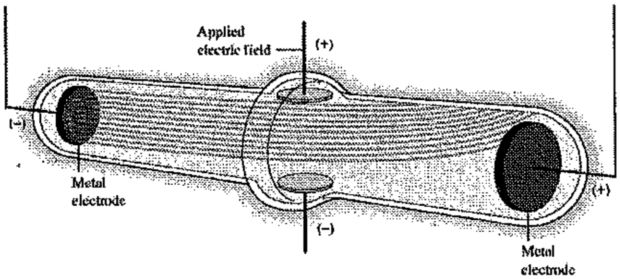


2.4 EARLY EXPERIMENTS TO CHARATERIZE THE ATOM

Based on the work of Dalton, Gay-Lussac, Avogadro, & others, chemistry was beginning to make sense [even if YOU disagree!] and the concept of the atom was clearly a good idea!

THE ELECTRON

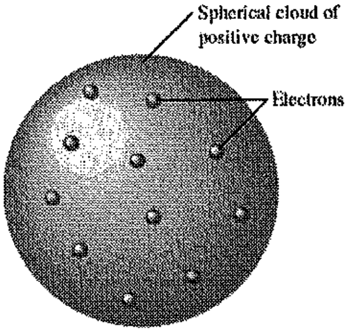
- J.J. Thomson, English (1898-1903)—found that when high voltage was applied to an evacuated tube, a "ray" he called a cathode ray [since it emanated from the (-) electrode or cathode when YOU apply a voltage across it] was produced.
  - The ray was produced at the (-) electrode
  - Repelled by the (-) pole of an applied electric field, E
  - He postulated the ray was a stream of NEGATIVE particles now called electrons, e^-



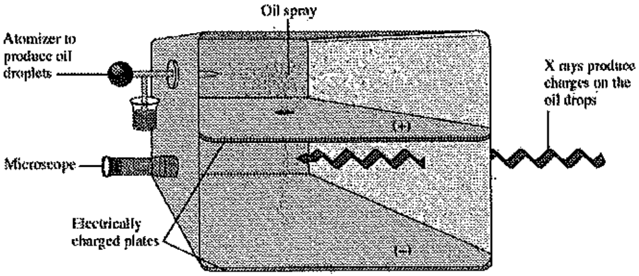
  - He then measured the deflection of beams of e^- to determine the **charge-to-mass ratio**

$$\frac{e}{m} = -1.76 \times 10^{8} \frac{C}{g}$$

  - *e* is charge on electron in Coulombs, (C) and m is its mass.



  - Thomson discovered that he could repeat this deflection and calculation using electrodes of different metals ∴ all metals contained electrons and ALL ATOMS contained electrons
  - Furthermore, all atoms were neutral ∴ there must be some (+) charge within the atom and the "plum pudding" model was born. Lord Kelvin may have played a role in the development of this model. [the British call every dessert pudding—we'd call it raisin bread where the raisins were the electrons randomly distributed throughout the + bread]



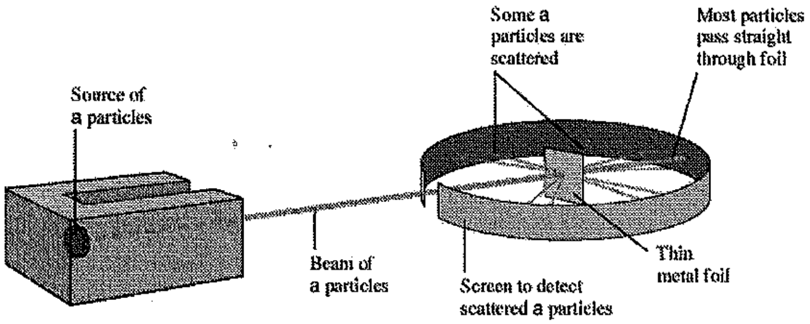
- 1909 Robert Millikan, American—University of Chicago, sprayed charged oil drops into a chamber. Next, he halted their fall due to gravity by adjusting the voltage across 2 charged plates. Now the voltage needed to halt the fall and the mass of the oil drop can be used to calculate the charge on the oil drop which is a whole number multiple of the electron charge. Mass of $e^- = 9.11 \times 10^{-31}$ kg.

RADIOACTIVITY

- Henri Becquerel, French—found out quite by accident [serendipity] that a piece of mineral containing uranium could produce its image on a photographic plate in the *absence* of light. He called this **radioactivity** and attributed it to a spontaneous emission of radiation by the uranium in the mineral sample.
- THREE types of radioactive emission:
  - **alpha, α--**equivalent to a helium nucleus; the largest particle radioactive particle emitted; 7300 times the mass of an electron. ${}^{4}_{2}He$ Since these are larger that the rest, early atomic studies often involved them.
  - **beta, β**--a high speed electron. ${}^{0}_{-1}\beta$ OR ${}^{0}_{-1}e$
  - **gamma, γ**--pure energy, no particles at all! Most penetrating, therefore, most dangerous.

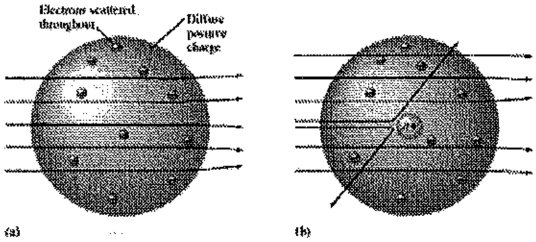
THE NUCLEAR ATOM

- 1911 Ernest Rutherford, England—A pioneer in radioactive studies, he carried out experiments to test Thomson's plum pudding model.
  - Directed α particles at a thin sheet of gold foil. He thought that if Thomson was correct, then the massive α particles would blast through the thin foil like "cannonballs through gauze". [He actually had a pair of graduate students Geiger & Marsden do the first rounds of experiments.] He expected the α particles to pass through with minor and occasional deflections.



  - The results were astounding [poor Geiger and Marsden first suffered Rutherford's wrath and were told to try again—this couldn't be!].
    - Most of the α particles did pass straight through, BUT many were deflected at LARGE angles and some even REFLECTED!
    - Rutherford stated that was like "shooting a howitzer at a piece of tissue paper and having the shell reflected back".
    - He knew the plumb pudding model could not be correct!
    - Those particles with large deflection angles had a "close encounter" with the dense positive center of the atom
    - Those that were reflected had a "direct hit"
    - He conceived the **nuclear atom**; that with a dense (+) core or nucleus

- This center contains most of the mass of the atom while the remainder of the atom is empty space!



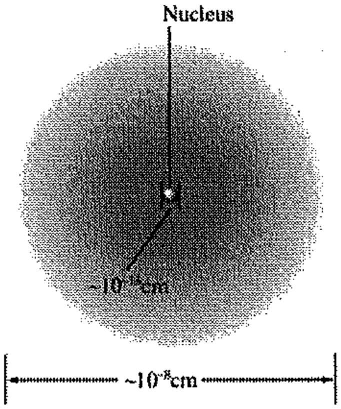
2.5 THE MODERN VIEW OF ATOMIC STRUCTURE: AN INTRODUCTION

ELEMENTS

All matter composed of only one type of atom is an element. There are 92 naturally occurring, all others are *man*made.

ATOMS

atom--the smallest particle of an element that retains the chemical properties of that element.



Particle	Mass	Charge
e^-	9.11×10^{-31}	1-
p^+	1.67×10^{-27}	1+
n^0	1.67×10^{-27}	None

- **nucleus**--contains the protons and the neutrons; the electrons are located outside the nucleus. Diameter = 10^{-13} cm. The electrons are located 10^{-8}cm from the nucleus. A mass of nuclear material the size of a pea would weigh 250 million tons! Very dense!
  - **proton**--positive charge, responsible for the identity of the element, defines *atomic number*
  - **neutron**--no charge, same size & mass as a proton, responsible for *isotopes*, alters *atomic mass number*
  - **electron**--negative charge, same size as a proton or neutron, BUT 1/2,000 the mass of a proton or neutron, responsible for bonding, hence reactions and ionizations, easily added or removed.
- **atomic number(Z)**--The number of p+ in an atom. All atoms of the *same* element have the *same number* of p+.
- **mass number(A)**--The sum of the number of neutrons and p+ for an atom. A different mass number *does not* mean a different element--just an isotope.

mass number → $^{A}_{Z}\mathrm{X}$ ← element symbol
atomic numbe r→

- actual mass is not an integral number! **mass defect**--causes this and is related to the energy binding the particles of the nucleus together

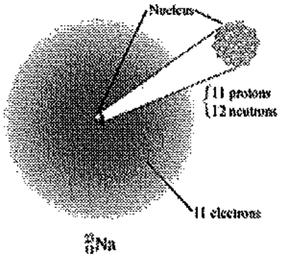
Exercise 2.2 Writing the Symbols for Atoms

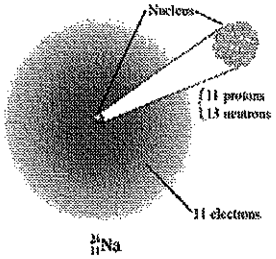
Write the symbol for the atom that has an atomic number of 9 and a mass number of 19. How many electrons and how many neutrons does this atom have?

F; **9 electrons and 10 neutrons**

ISOTOPES

- **isotopes**--atoms having the same atomic number (# of p+) but a different number of neutrons
    - most elements have at least two stable isotopes, there are very few with only one stable isotope (Al, F, P)
    - hydrogens isotopes are so important they have special names:
        - 0 neutrons ☞ hydrogen
        - 1 neutron ☞ deuterium
        - 2 neutrons ☞ tritium



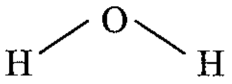


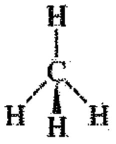
2.6 MOLECULES AND IONS

Electrons are responsible for bonding and chemical reactivity

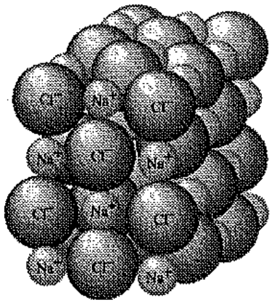
- **Chemical bonds**—forces that hold atoms together
- **Covalenet bonds**—atoms share electrons and make molecules [independent units]; H_2, CO_2, H_2O, NH_3, O_2, CH_4 to name a few.
- **molecule**--smallest unit of a compound that retains the chem. characteristics of the compound; characteristics of the constituent elements are lost.
- **molecular formula**--uses symbols and subscripts to represent the composition of the molecule. (Strictest sense--covalently bonded)
- **structural formula**—bonds are shown by lines [representing shared e^- pairs]; may NOT indicate shape

$H—O—H$





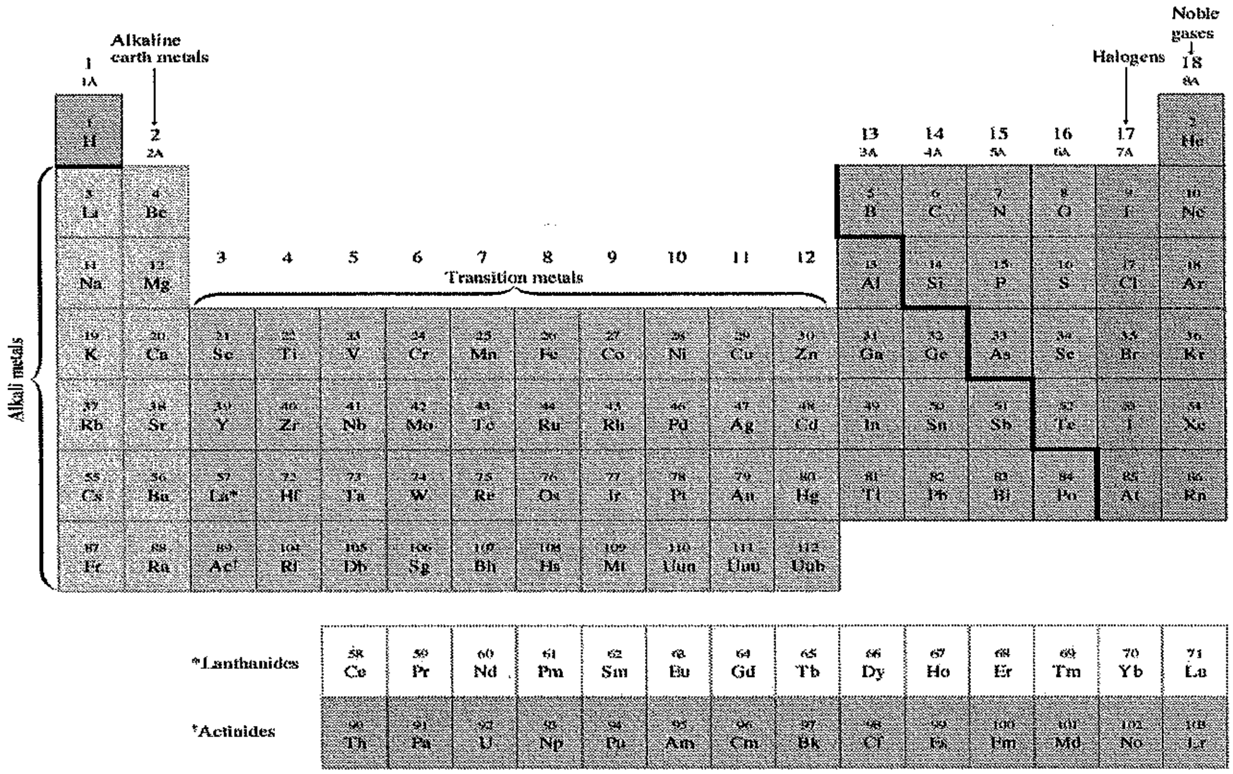
Methane



(a)

- **ions**--formed when electrons are lost or gained in ordinary chem. reactions; affect size of atom dramatically
    - **cations**--(+) ions; often metals since metals *lose* electrons to become + charged
    - **anions**--(-) ions; often nonmetals since nonmetals *gain* electrons to become - charged
    - **polyatomic ions**--units of atoms behaving as one entity ☞ MEMORIZE formula and charge!
    - **ionic solids**—Electrostatic forces hold ions together. Strong ∴ ions held close together ∴ solids.

2.7 AN INTRODUCTION TO THE PERIODIC TABLE



- Atomic number = number of protons and is written above each symbol
- **metals**—malleable, ductile & have luster; most of the elements are metals—exist as cations in a "sea of electrons" which accounts for their excellent conductive properties; form oxides [tarnish] readily and form POSITIVE ions [cations]. Why must some have such goofy symbols?

Current Name	Original Name	Symbol
Antimony	Stibium	Sb
Copper	Cuprum	Cu
Iron	Ferrum	Fe
Lead	Plumbum	Pb
Mercury	Hydrargyrum	Hg
Potassium	Kalium	K
Silver	Argentum	Ag
Sodium	Natrium	Na
Tin	Stannum	Sn
Tungsten	Wolfram	W

- **groups or families**--vertical columns; have similar physical and chemical properties (based on similar electron configurations!!)
  - **group A**—Representative elements
  - **group B**--transition elements; all metals; have numerous oxidation/valence states
- **periods**--horizonal rows; progress from metals to metalloids [either side of the black "stair step" line that separates metals from nonmetals] to nonmetals
- **MEMORIZE:**
  1. ALKALI METALS—1A
  2. ALKALINE EARTH METALS—2A
  3. HALOGENS—7A
  4. NOBLE (RARE) GASSES—8A

2.8 NAMING SIMPLE COMPOUNDS

BINARY IONIC COMPOUNDS

Naming + ions: usually metals

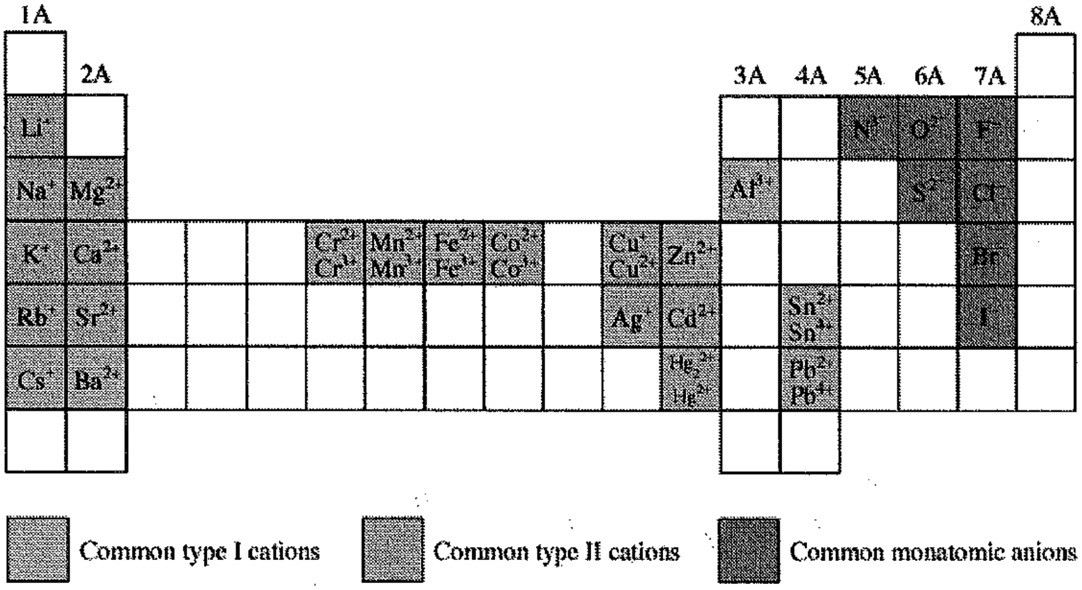
- monatomic, metal, cation → simply the name of the metal from which it is derived. **Al^{3+} is the aluminum ion**
- transition metals form *more than one* ion; Roman Numerals (in) follow the ion@ name. **Cu^{2+} is copper(II)** (Yeah, the no space thing between the ion's name and (II) looks strange, but it is the correct way to do it.)

  Mercury (I) is an exception ☞ it is Hg_2^{2+} ∴ two Hg^+ associated together.
- NH_4^+ is ammonium
- NO ROMAN NUMERAL WHEN USING silver, cadmium and zinc. [Arrange their SYMBOLS in alphabetical order—first one is 1+ and the other two are 2+]

Naming - ions: monatomic and polyatomic

- **MONATOMIC**--add the suffix ***-ide*** to the stem of the nonmetal's name. Halogens are called the ***halides***.
- **POLYATOMIC**--quite common; oxyanions are the PA ' s containing oxygen.
  - *hypo*--"ate" least oxygen
  - *-ite*--"ate" more oxygen than hypo-
  - *-ate*--"ate" more oxygen than -ite
  - ***hyper***---ate--"ate" the most oxygen

NAMING IONIC COMPOUNDS: The + ion name is given *first* followed by the name of the negative ion.



Exercise 2.3 **Naming Type I Binary Compounds**

Name each binary compound.

a. CsF b. $A1C1_3$ c. LiH

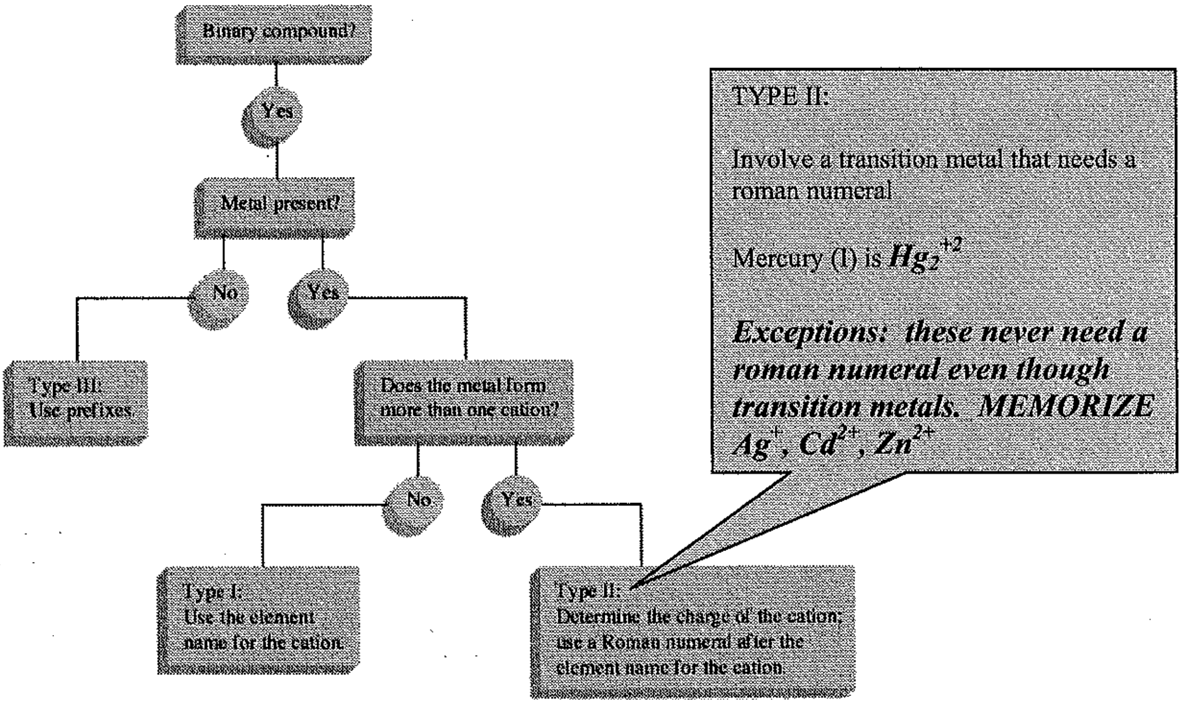
a. cesium fluoride
b. aluminum chloride
c. lithium hydride

Exercise 2.4 **Naming Type II Binary Compounds**

Give the systematic name of each of the following compounds.

a. $CuC1$ b. HgO c. Fe_2O_3 d. MnO_2 e. $PbC1_2$

a. copper(I) chloride,
b. mercury(II) oxide.
c. iron(III) oxide.
d. manganese(IV) oxide.
e. lead(II) chloride.



TYPE II:

Involve a transition metal that needs a roman numeral

Mercury (I) is Hg_2^{+2}

Exceptions: these never need a roman numeral even though transition metals. MEMORIZE Ag^+, Cd^{2+}, Zn^{2+}

Exercise 2.5 **Naming Binary Compounds**

Give the systematic name of each of the following compounds.

a. $CoBr_2$ b. $CaC1_2$ c. $A1_2O_3$ d. $CrC1_3$

a. Cobalt(II) bromide
b. Calcium chloride
c. Aluminume oxide
d. Chromium(III) chloride

Exercise 2.6 **Naming Compounds Containing Polyatomic Ions**

Give the systematic name of each of the following compounds.

a. Na_2SO_4 b. KH_2PO_4 c. $Fe(NO_3)_3$ d. $Mn(OH)_2$
e. Na_2SO_3 f. Na_2CO_3 g. $NaHCO_3$ h. $CsC1O_4$
i. NaOC1 j. Na_2SeO_4 k. $KBrO_3$

a. Sodium sulfate
b. Potassium dihydrogen phosphate
c. Iron(III) nitrate
d. Manganese(II) hydroxide
e. Sodium sulfite
f. Sodium carbonate
g. Sodium hydrogen carbonate
h. Cesium perchlorate
i. Sodium hypochlorite
j. Sodium selenate
k. Potassium bromate

NAMING BINARY COVALENT COMPOUNDS : (covalently bonded)

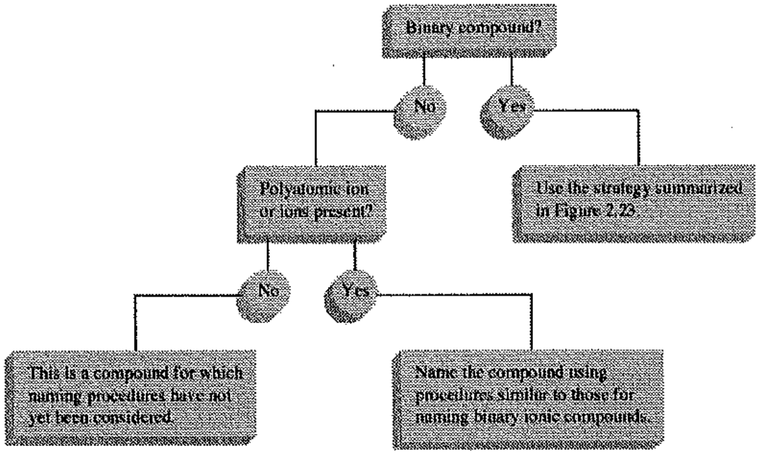
- use prefixes!!! Don't forget the –ide ending as well.

Exercise 2.7 **Naming Type III Binary Compounds**

Name each of the following compounds.

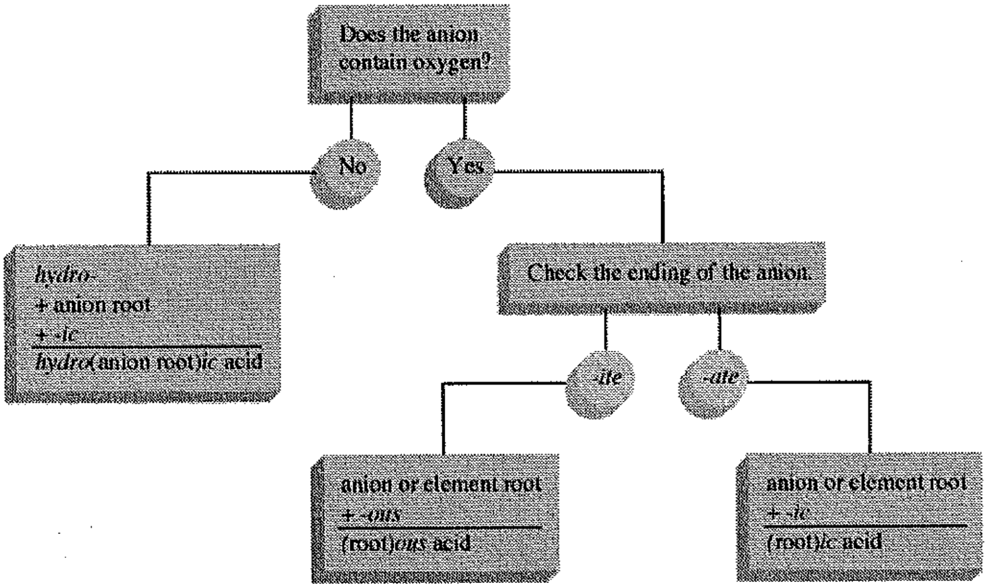
a. $PC1_5$ b. $PC1_3$ c. SF_6 d. SO_3 e. SO_2 f. CO_2

a. Phosphorus pentachloride
b. Phosphorus trichloride
c. Sulfur hexafluoride
d. Sulfur trioxide
e. Sulfur dioxide
f. Carbon dioxide



ACIDS

- hydrogen, if present, is listed first
  - naming acids ☞ *-ide* → hydro [negative ion root]ic ACID
    -ate → *-ic* *ACID*
    -ite → *-ous* *ACID*



- PAINS IN THE GLUTEUS MAXIMUS: these lovely creatures have been around longer than the naming system and no one wanted to adapt!!
  - water
  - ammonia
  - hydrazine
  - phosphine
  - nitric oxide
  - nitrous oxide ("laughing gas")

Exercise 2.8 **Naming Various Types of Compounds**

Give the systematic name for each of the following compounds.

a. P_4O_{10} b. Nb_2O_5 c. Li_2O_2 d. $Ti(NO_3)_4$

a. Tetraphosphorus decaoxide
b. Niobium(V) oxide
c. Lithium peroxide
d. Titanium(IV) nitrate

Exercise 2.9 **Writing Compound Formulas from Names**

Given the following systematic names, write the formula for each compound.

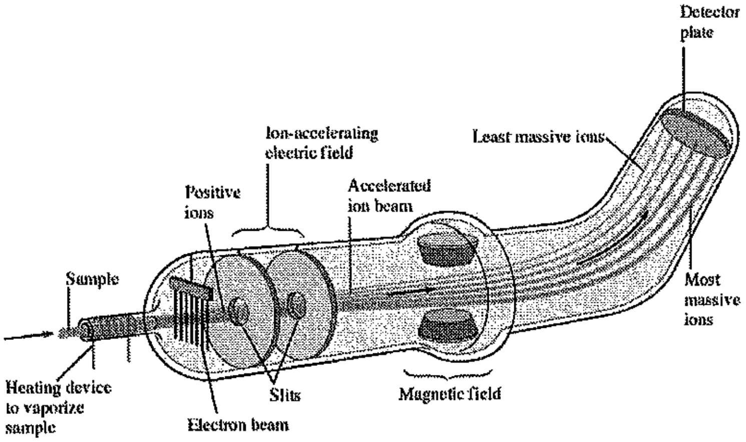
a. Vanadium(V) fluoride b. Dioxygen difluoride
c. Rubidium peroxide d. Gallium oxide

a. VF_5
b. O_2F_2
c. Rb_2O_2
d. Ga_2O_3

AP* Chemistry
STOICHIOMETRY

3.1 ATOMIC MASSES

- **^{12}C—Carbon 12**—In 1961 it was agreed that this would serve as the standard and would be defined to have a mass of EXACTLY 12 atomic mass units (amu). All other atomic masses are measured ***relative*** to this.
- **mass spectrometer**—a device for measuring the mass of atoms or molecules
  - atoms or molecules are passed into a beam of high-speed electrons
  - this knocks electrons OFF the atoms or molecules transforming them into cations
  - apply an electric field
  - this accelerates the cations since they are repelled from the (+) pole and attracted toward the (−) pole
  - send the accelerated cations into a magnetic field
  - an accelerated cation creates it's OWN magnetic field which perturbs the original magnetic field
  - this perturbation changes the path of the cation
  - the amount of deflection is proportional to the mass; heavy cations deflect little
  - ions hit a detector plate where measurements can be obtained.



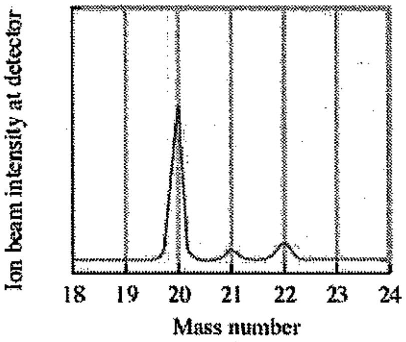
- $$\frac{Mass\,^{13}C}{Mass\,^{12}C} = 1.0836129 \therefore Mass\,^{13}C = (1.0836129)(12amu) = 13.003355amu$$

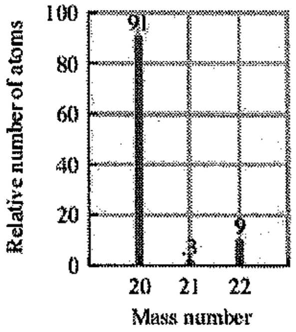
  Exact by definition

- **average atomic masses**—atoms have masses of whole numbers, HOWEVER samples of quadrillions of atoms have a few that are heavier or lighter [isotopes] due to different numbers of neutrons present
- **percent abundance**--percentage of atoms in a natural sample of the pure element represented by a particular isotope

$$\text{percent abundance} = \frac{\text{number of atoms of a given isotope}}{\text{Total number of atoms of all isotopes of that element}} \times 100$$

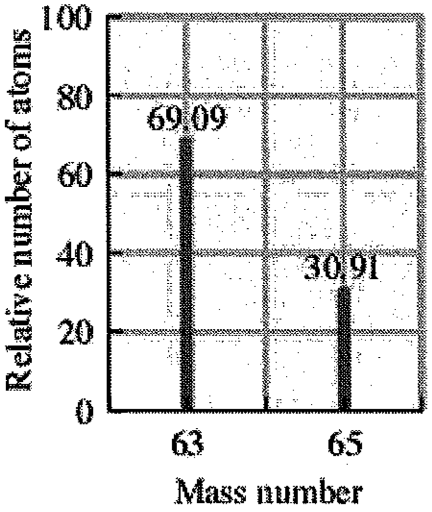
- **counting by mass**—when particles are small this is a matter of convenience. Just as you buy 5 lbs of sugar rather than a number of sugar crystals, or a pound of peanuts rather than counting the individual peanuts….this concept works very well if your know an *average* mass.
- **mass spectrometer to determine isotopic composition**—load in a pure sample of natural neon or other substance. The areas of the "peaks" or heights of the bars indicate the relative abundances of ${}^{20}_{10}Ne$, ${}^{21}_{10}Ne$, and ${}^{22}_{10}Ne$





Exercise 3.1 **The Average Mass of an Element**

When a sample of natural copper is vaporized and injected into a mass spectrometer, the results shown in the figure are obtained. Use these data to compute the average mass of natural copper. (The mass values for ${}^{63}Cu$ and ${}^{65}Cu$ are 62.93 amu and 64.93 amu, respectively.)



63.55 amu/atom

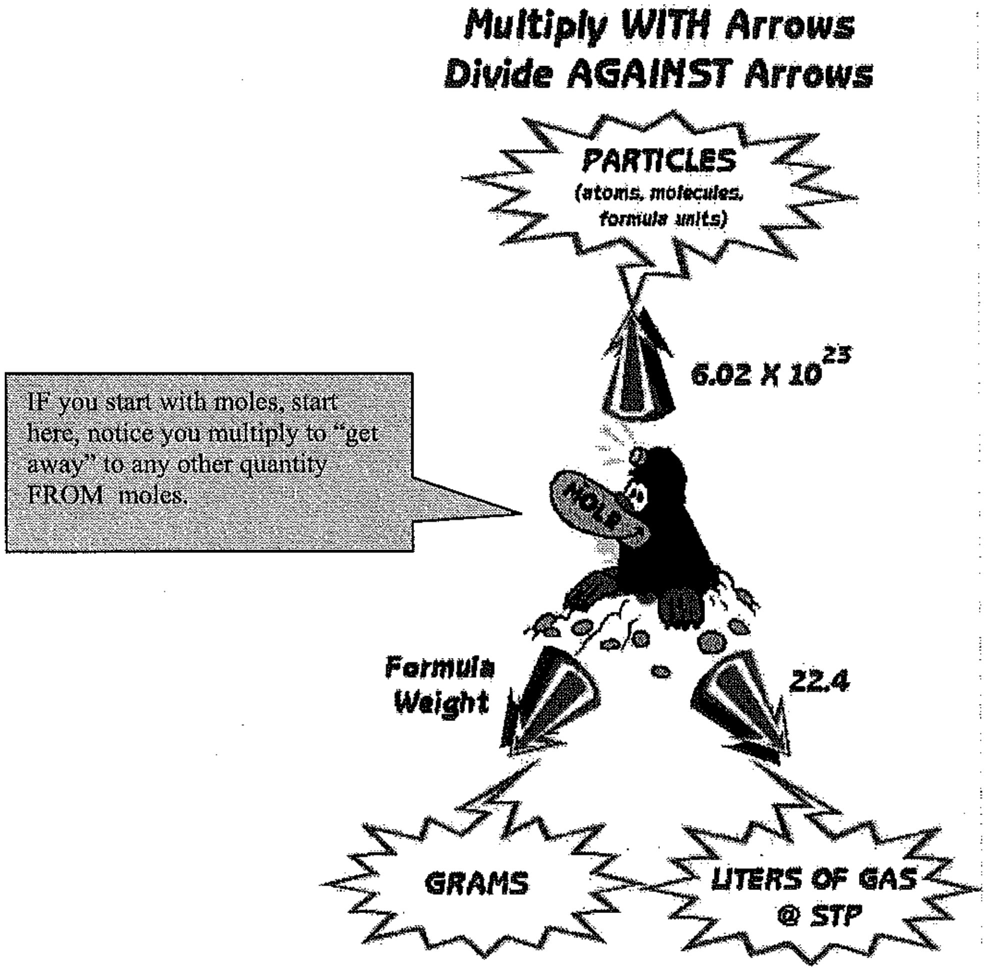
3.2 THE MOLE

- **mole**—the number of C atoms in exactly 12.0 grams of ${}^{12}C$; also a number, 6.02×10^{23} just as the word "dozen" means 12 and "couple" means 2.
- **Avogadro's number**—6.02×10^{23}, the number of particles in a mole of anything

DIMENSIONAL ANALYSIS DISCLAIMER: Beginning on page 84 of the Chapter 3 text files you can find on this CD, you can find all of the remaining exercises worked out with dimensional analysis. This is most likely the way you were taught in Chemistry I. I will show you some alternatives to dimensional analysis. WHY? First, some of these techniques are faster and well-suited to the multi-step problems you will face on the AP Exam. Secondly, these techniques better prepare you to work the complex equilibrium problems you will face later in this course. The first problem you must solve in the free response section of the AP Exam will be an equilibrium problem and you will need to be able to work them quickly. Lastly, I used to teach both methods. Generations of successful students have encouraged me to share these techniques with as many students as possible. They did, once they got to college, and made lots of new friends once word got out they had this "cool way" to solve stoichiometry problems—not to mention their good grades! Give this a try. It doesn't matter which method you use, I encourage you to use the method that works best for you and lets you solve problems *accurately and quickly*!

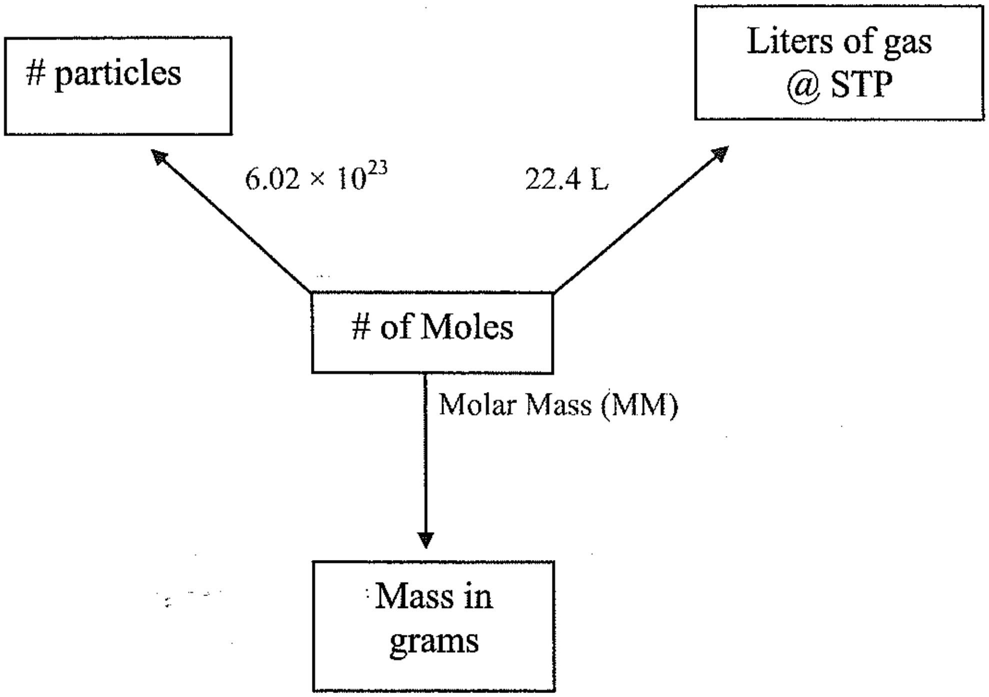
ALTERNATE TECHNIQUE #1—USING THE MOLE MAP:
Simply reproduce this map on your scratch paper until you no longer need to since the image will be burned into your brain!

MULTIPLY [by the conversion factor on the arrow] when traveling IN THE DIRECTION OF THE ARROW and obviously, divide when "traveling" against an arrow.



IF you start with moles, start here, notice you multiply to "get away" to any other quantity FROM moles.

When you draw this it will look more like this:



Exercise 3.2 Determining the Mass of a Sample of Atoms

Americium is an element that does not occur naturally. It can be made in very small amounts in a device known as a *particle accelerator*. Compute the mass in grams of a sample of americium containing six atoms.
[There's a copy of the AP Chem periodic table at apchemistrynmsi.wikispaces.com PRINT—print it!]

2.42×10^{-21} g

Exercise 3.4 Determining Moles of Atoms

Aluminum (A1) is a metal with a high strength-to-mass ratio and a high resistance to corrosion; thus it is often used for structural purposes. Compute both the number of moles of atoms and the number of atoms in a 10.0-g sample of aluminum.

2.23×10^{23} atoms

Exercise 3.5 **Calculating the Number of Moles and Mass**

Cobalt (Co) is a metal that is added to steel to improve its resistance to corrosion. Calculate both the number of moles in a sample of cobalt containing 5.00×10^{20} atoms and the mass of the sample.

8.31×10^{-4} mol Co
4.89×10^{-2} g Co

3.3 MOLAR MASS, MOLECULAR WEIGHT, AND FORMULA WEIGHT

- **molar mass, *MM***--the mass in grams of Avogadro's number of molecules; i.e. the mass of a mole!
- **molecular weight, MW**--sum of all the atomic weights of all the atoms in the formula (it is *essential you have a correct formula as you'll painfully discover!)*
- **empirical formula**--that ratio in the network for an ionic substance.
- **formula weight**--same as molecular weight, just a language problem ☞ "molecular" implies covalent bonding while "formula" implies ionic bonding {just consider this to be a giant conspiracy designed to keep the uneducated from *ever* understanding chemistry—kind of like the scoring scheme in tennis}. **The AP Exam uses MM for all formula masses.**
- **A WORD ABOUT SIG. FIG.'s**—It is correct to "pull" from the periodic table as many sig. figs for your MM's as are in your problem—just stick with 2 decimal places for all—much simpler!

Exercise 3.6 **Calculating Molar Mass I**

Juglone, a dye known for centuries, is produced from the husks of black walnuts. It is also a natural herbicide (weed killer) that kills off competitive plants around the black walnut tree but does not affect grass and other noncompetitive plants [a concept called *allelopathy*]. The formula for juglone is $C_{10}H_6O_3$.

a. Calculate the molar mass of juglone.

b. A sample of 1.56×10^{-2} g of pure juglone was extracted from black walnut husks. How many moles of juglone does this sample represent?

a. 174.16 g/mol

b. 8.96×10^{-5} mol juglone

Exercise 3.7 Calculating Molar Mass II

Calcium carbonate ($CaCO_3$), also called *calcite*, is the principal mineral found in limestone, marble, chalk, pearls, and the shells of marine animals such as clams.

a. Calculate the molar mass of calcium carbonate.

b. A certain sample of calcium carbonate contains 4.86 moles. What is the mass in grams of this sample? What is the mass of the CO_3^{2-} ions present?

a. 100.09 g/mol
b. 486 g; 292g CO_3^{2-}

Exercise 3.8 Molar Mass and Numbers of Molecules

Isopentyl acetate ($C_7H_{14}O_2$), the compound responsible for the scent of bananas, can be produced commercially. Interestingly, bees release about 1μg (1×10^{-6} g) of this compound when they sting. The resulting scent attracts other bees to join the attack. How many molecules of isopentyl acetate are released in a typical bee sting?

How many atoms of carbon are present?

5×10^{15} molecules
4×10^{16} carbon atoms

ELEMENTS THAT EXIST AS MOLECULES

Pure hydrogen, nitrogen, oxygen and the halogens [I call them the "gens" collectively—easier to remember!] exist as **<u>DIATOMIC</u>** molecules *under normal conditions.* MEMORIZE!!! Be sure you compute their molar masses as diatomics. Others to be aware of, but not memorize:

- P_4--tetratomic form of elemental phosphorous; an allotrope
- S_8—sulfur's elemental form; also an allotrope
- Carbon--diamond and graphite ☞ covalent networks of atoms

3.4 PERCENT COMPOSITION OF COMPOUNDS

Two common ways of describing the composition of a compound: in terms of the number of its constituent atoms and in terms of the percentages (by mass) of its elements.

Percent (by mass) Composition: law of constant composition states that *any sample of a pure compound always consists of the same elements combined in the same proportions by mass.*

$$\% \text{ comp} = \frac{\text{mass of desired element}}{\text{Total mass of compound}} \times 100$$

Consider ethanol, C_2H_5OH

$$\text{Mass \% of C} = 2 \cancel{\text{mol}} \times 12.01 \frac{g}{\cancel{\text{mol}}} = 24.02 \text{ g}$$

$$\text{Mass \% of H} = 6 \cancel{\text{mol}} \times 1.01 \frac{g}{\cancel{\text{mol}}} = 6.06 \text{ g}$$

$$\text{Mass \% of O} = \cancel{1 \text{ mol}} \times 16.00 \frac{g}{\cancel{\text{mol}}} = \underline{16.00\text{g}}$$

$$\text{Mass of 1 mol of } C_2H_5OH = 46.08 \text{ g/mol}$$

NEXT THE MASS PERCENT CAN BE CALCULATED:

$$\text{Mass percent of C} = \frac{24.02 \text{ g C}}{46.08 \text{ g/mol}} \times 100\% = 52.13\%$$

Repeat for the H and O present.

Exercise 3.9 **Calculating Mass Percent I**

Carvone is a substance that occurs in two forms having different arrangements of the atoms but the same molecular formula ($C_{10}H_{14}O$) and mass. One type of carvone gives caraway seeds their characteristic smell, and the other type is responsible for the smell of spearmint oil. Compute the mass percent of each element in carvone.

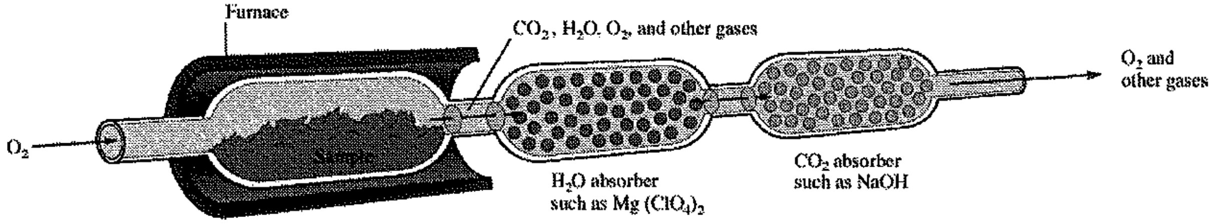
C = 79.94%
H = 9.41%
O = 10.65%

Exercise 3.10 Calculating Mass Percent II

Penicillin, the first of a now large number of antibiotics (antibacterial agents), was discovered accidentally by the Scottish bacteriologist Alexander Fleming in 1928, but he was never able to isolate it as a pure compound. This and similar antibiotics have saved millions of lives that might have been lost to infections. Penicillin F has the formula $C_{14}H_{20}N_2SO_4$. Compute the mass percent of each element.

C = 53.82%
H = 6.466%
N = 8.969%
S = 10.27%
O = 20.49%

3.5 DETERMINING THE FORMULA OF A COMPOUND



When faced with a compound of "unknown" formula, one of the most common techniques is to combust it with oxygen to produce CO_2, H_2O, and N_2 which are then collected and weighed.

- **empirical and molecular formulas:** assume a 100 gram sample if given %
  - empirical gives smallest ratio
  - need to know molar mass to establish molecular formula which is (empirical formula)$_n$, where n is an integer

- **determining empirical and molecular formulas**
  - **<u>hydrates</u>**—"dot waters" used to cement crystal structures.
  - **<u>anhydrous</u>**--*without* water

Example: A compound is composed of carbon, nitrogen and hydrogen. When 0.1156 g of this compound is reacted with oxygen [burned, combusted], 0.1638 g or carbon dioxide and 0.1676 g of water are collected. What is the empirical formula of the compound?

Compound + $O_2 \rightarrow CO_2 + H_2O + N_2$ but NOT balanced!!

You can see that all of the carbon ended up in CO_2 so…when in doubt, FIND THE NUMBER OF MOLES!!

0.1638 g ÷ 44.01 g/mol = 0.003781 moles of CO_2 = 0.003781 moles of C

Next, you can see that all of the hydrogen ended up in H_2O, so….FIND THE NUMBER OF MOLES!!

0.1676 ÷ 18.02 g/mol = 0 .009301 moles of H_2O, BUT there are 2 moles of H for each mole of water [think "organ bank" one heart per body, one C per molecule of carbon dioxide—2 lungs per body, 2 atoms H in water and so on…] so DOUBLE THE NUMBER OF MOLES TO GET THE NUMBER OF MOLES OF HYDROGEN!! moles H = 0.01860 moles of H

The rest must be nitrogen, BUT we only have mass data for the sample so convert your moles of C and H to grams:

g C = 0.003781 moles C × 12.01 = 0.04540 grams C

g H = 0.01860 moles H × 1.01 = 0.01879 grams H
0.06419 grams total thus far

SUBTRACT!

0.1156 g sample – 0.06419 g thus far = grams N left = 0.05141 g N so….

0.05141 g N ÷ 14.01 = 0.003670 moles N

Chemical formulas represent mole to mole ratios, so…divide the number of moles of each by the smallest # of moles of any one of them to get a guaranteed ONE in your ratios…multiply by 2, then 3, etc to get to a ratio of small whole numbers!!

Element	# moles	ALL Divided by 0.003670
C	0.003781	1
H	0.01860	5
N	0.003670	1

Therefore the correct EMPIRICAL formula is CH_5N.

Next, if we are told that the MM is 31.06 g/mol, then simply use this relationship:

$$(\text{Empirical mass}) \times n = MM$$

$$(12.01 + 5.05 + 14.01) \times n = 31.06$$

Solve for n

n = 0.999678… or essentially one, so the empirical formula and the molecular formula are one in the same.

One last trick of the trade: When you don't know the mass of your sample, assume 100 grams so that any percents become grams….proceed by finding the number of moles!

Empirical Formula Determination

- Since mass percentage gives the number of grams of a particular element per 100 grams of compound, base the calculation on 100 grams of compound. Each percent will then represent the mass in grams of that element.
- Determine the number of moles of each element present in 100 grams of compound using the atomic masses of the elements present.
- Divide each value of the number of moles by the smallest of the values. If each resulting number is a whole number (after appropriate rounding), these numbers represent the subscripts of the elements in the empirical formula.
- If the numbers obtained in the previous step are not whole numbers, multiply each number by an integer so that the results are all whole numbers.

Exercise 3.11 Determining Empirical and Molecular Formulas I

Determine the empirical and molecular formulas for a compound that gives the following analysis (in mass percents):

71.65% C1 24.27% C 4.07% H

The molar mass is known to be 98.96 g/mol.

Empirical formula = CH_2Cl
Molecular formula = $C_2H_4C1_2$

Exercise 3.12 **Determining Empirical and Molecular Formulas II**

A white powder is analyzed and found to contain 43.64% phosphorus and 56.36% oxygen by mass. The compound has a molar mass of 283.88 g/mol. What are the compound's empirical and molecular formulas?

Empirical formula = P_2O_5

Molecular formula = $(P_2O_5)_2$ or P_4O_{10}

Exercise 3.13 **Determining a Molecular Formula**

Caffeine, a stimulant found in coffee, tea, and chocolate, contains 49.48% carbon, 5.15% hydrogen, 28.87% nitrogen, and 16.49% oxygen by mass and has a molar mass of 194.2 g/mol. Determine the molecular formula of caffeine.

Molecular formula = $C_8H_{10}N_4O_2$

3.6 & 3.7 BALANCING CHEMICAL EQUATIONS

CHEMICAL REACTIONS

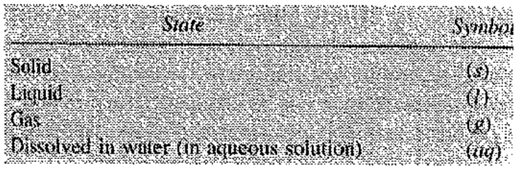
Chemical reactions are the result of a chemical change where atoms are reorganized into one or more new arrangements. Bonds are broken [requires energy] and new ones are formed [releases energy].

CHEMICAL EQUATIONS

<u>chemical reaction</u>--transforms elements and compounds into new substances

<u>balanced chemical equation</u>--shows the relative amounts of reactants [left] and products [right] by molecule or by mole.

- *s, l, g, aq*--solid, liquid, gas, aqueous solution
- NO ENERGY or TIME is alluded to
- Antoine Lavoisier (1743-1794)--law of conservation of matter: *matter can neither be created nor destroyed*
  - ☞ this means "balancing equations" is all his fault!!



State	*Symbo*
Solid	(*s*)
Liquid	(*l*)
Gas	(*g*)
Dissolved in water (in aqueous solution)	(*aq*)

BALANCING CHEMICAL EQUATIONS

- Begin with the most complicated-looking thing (often the scariest, too).
- Save the elemental thing for last.
- If you get stuck, double the most complicated-looking thing.
- MEMORIZE THE FOLLOWING:
- metals + halogens → M_aX_b
- CH and/or O + O_2 → $CO_2(g) + H_2O(g)$
- H_2CO_3 [any time formed!] → $CO_2 + H_2O$; in other words, never write carbonic acid as a product, it spontaneously decomposes [in an open container] to become carbon dioxide and water.
- metal carbonates → metal OXIDES + CO_2

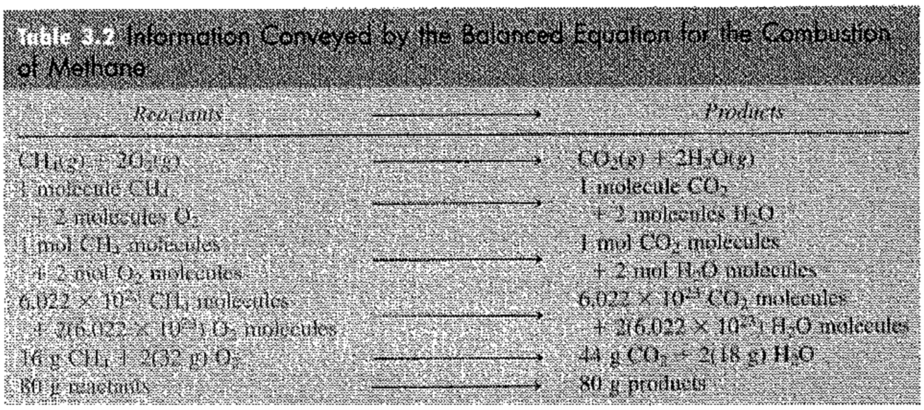


Table 3.2 Information Conveyed by the Balanced Equation for the Combustion of Methane

Reactants	→	*Products*
$CH_4(g) + 2O_2(g)$	→	$CO_2(g) + 2H_2O(g)$
1 molecule CH_4 + 2 molecules O_2	→	1 molecule CO_2 + 2 molecules H_2O
1 mol CH_4 molecules + 2 mol O_2 molecules	→	1 mol CO_2 molecules + 2 mol H_2O molecules
6.022×10^{23} CH_4 molecules + $2(6.022 \times 10^{23})$ O_2 molecules	→	6.022×10^{23} CO_2 molecules + $2(6.022 \times 10^{23})$ H_2O molecules
16 g CH_4 + 2(32 g) O_2	→	44 g CO_2 + 2(18 g) H_2O
80 g reactants	→	80 g products

Exercise 3.14 **Balancing a Chemical Equation I**

Chromium compounds exhibit a variety of bright colors. When solid ammonium dichromate, $(NH_4)_2Cr_2O_7$, a vivid orange compound, is ignited, a spectacular reaction occurs, as shown in the two photographs on page 105. Although the reaction is actually somewhat more complex, let's assume here that the products are solid chromium(III) oxide, nitrogen gas (consisting of N_2 molecules), and water vapor. Balance the equation for this reaction.

$$\underset{(4 \times 2)\ \mathrm{H}}{(NH_4)_2Cr_2O_7(s)} \rightarrow Cr_2O_3(s) + N_2(g) + \underset{(4 \times 2)\ \mathrm{H}}{4H_2O(g)}$$

Exercise 3.15 **Balancing a Chemical Equation II**

At 1000°C, ammonia gas, $NH_3(g)$, reacts with oxygen gas to form gaseous nitric oxide, $NO(g)$, and water vapor. This reaction is the first step in the commercial production of nitric acid by the Ostwald process. Balance the equation for this reaction.

$$4\,NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$$

3.8 STOICHIOMETRIC CALCULATIONS: AMOUNTS OF REACTANTS AND PRODUCTS

Stoichiometry is the most important thing you can learn as you embark upon AP Chemistry! Get good at this and you will do well all year. This NEVER goes away!

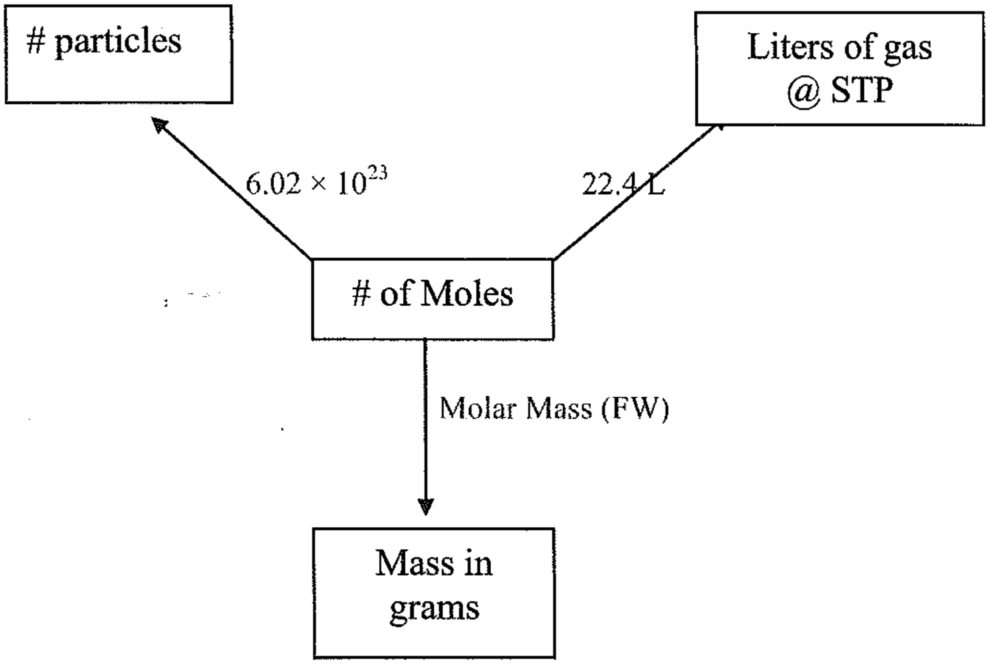
It's time to repeat my dimensional analysis disclaimer.

DIMENSIONAL ANALYSIS DISCLAIMER: Beginning on page 84 of the Chapter 3 text files you can find on this CD, you can find all of the remaining exercises worked out with dimensional analysis. This is most likely the way you were taught in Chemistry I. I will show you some alternatives to dimensional analysis. WHY? First, some of these techniques are faster and well-suited to the multi-step problems you will face on the AP Exam. Secondly, these techniques better prepare you to work the complex equilibrium problems you will face later in this course. The first problem you must solve in the free response section of the AP Exam will be an equilibrium problem and you will need to be able to work them quickly. Lastly, I used to teach both methods. Generations of successful students have encouraged me to share these techniques with as many students as possible. They did, once they got to college, and made lots of new friends once word got out they had this "cool way" to solve stoichiometry problems—not to mention their good grades! Give this a try. It doesn't matter which method you use, I encourage you to use the method that works best for you and lets you solve problems *accurately and quickly*!

First you have to be proficient at the following no matter which method you choose!:

- Writing CORRECT formulas—this requires knowledge of your polyatomic ions and being able to use the periodic table to deduce what you have not had to memorize. Review section 2.8 in your Chapter 2 notes or your text.
- Calculate CORRECT molar masses from a correctly written formula
- Balance a chemical equation
- Use the mole map to calculate the number of moles or anything else!

Remember the mole map? It will come in mighty handy as well!



Here's the "template" for solving the problems…you'll create a chart. Here's a typical example:

<u>What mass of oxygen will react with 96.1 grams of propane?</u>

[notice all words—you supply chemical formulas!]

Molar Mass:	(44.11)		(32.00)		(44.01)		(18.02)
Balanced Eq'n	C_3H_8	+	5 O_2	→	3 CO_2	+	4 H_2O
mole:mole	**1**		**5**		**3**		**4**
# moles							
amount							

1. Write a chemical equation paying special attention to writing correct chemical formulas!
2. Calculate the molar masses and put in parentheses above the formulas—soon you'll figure out you don't have to do this for every reactant and product, just those you're interested in.
3. Look at the coefficients on the balanced equation, they ARE the mole:mole ratios!

4. Next, re-read the problem and put in an amount—in this example it's 96.1 g of propane.

Molar Mass:	(44.11)		(32.00)		(44.01)		(18.02)
Balanced Eq'n	C_3H_8	+	5 O_2	→	3 CO_2	+	4 H_2O
mole:mole	**1**		**5**		**3**		**4**
# moles	2.18		10.9		6.53		8.71
amount	96.1 grams						

5. Find the number of moles of something, anything! Use the mole map. Start at 96.1 grams, divide [against the arrow] by molar mass to get the # moles of propane.

6. USE the mole: mole to find moles of EVERYTHING! If 1 = 2.18 then oxygen is 5(2.18) etc.... [IF the first you find is not a "1", just divide to make it "1" and then it's easy greasy!] Leave everything in your calculator—I only rounded to save space!

7. Re-read the problem to determine which amount was asked for...here's the payoff....AP problems ask for several amounts! First, we'll find the mass of oxygen required since that's what the problem asked. 10.9 moles × 32.00 g/mol = 349 g of oxygen

Molar Mass:	(44.11)		(32.00)		(44.01)		(18.02)
Balanced Eq'n	C_3H_8	+	5 O_2	→	3 CO_2	+	4 H_2O
mole:mole	**1**		**5**		**3**		**4**
# moles	2.18		10.9		6.53		8.71
amount	96.1 grams		349 g		146 L		

Now, humor me...What if part b asked for liters of CO_2 at STP [1 atm, 273K]? Use the mole map. Start in the middle with 6.53 moles × [in direction of arrow] 22.4 L/mol = 146 L

AND...how many water molecules are produced? Use the mole map , start in the middle with 8.71 mol (6.02×10^{23}) = 5.24×10^{24} molecules of water.

Try these two exercises with whichever method you like best!

Exercise 3.16 **Chemical Stoichiometry I**

Solid lithium hydroxide is used in space vehicles to remove exhaled carbon dioxide from the living environment by forming solid lithium carbonate and liquid water. What mass of gaseous carbon dioxide can be absorbed by 1.00 kg of lithium hydroxide?

919 g

Exercise 3.17 **Chemical Stoichiometry II**

Baking soda ($NaHCO_3$) is often used as an antacid. It neutralizes excess hydrochloric acid secreted by the stomach:

$$NaHCO_3(s) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l) + CO_2(aq)$$

Milk of magnesia, which is an aqueous suspension of magnesium hydroxide, is also used as an antacid:

$$Mg(OH)_2(s) + 2HCl(aq) \rightarrow 2H_2O(l) + MgCl_2(aq)$$

Which is the more effective antacid per gram, $NaHCO_3$ or $Mg(OH)_2$?

$Mg(OH)_2$

3.9 CALCULATIONS INVOLVING A LIMITING REACTANT

Ever notice how hot dogs are sold in packages of 10 while the buns come in packages of 8?? The bun is the limiting reactant and limits the hot dog production to 8 as well! The limiting reactant [or reagent] is the one consumed most entirely in the chemical reaction.

Plan of attack: First, you'll know you ***need*** a plan if you are given TWO amounts of matter that react.

Next, when in doubt…find the number of moles. Set up your table like before, only now you'll have TWO amounts and thus TWO # 's of moles to get you started. I cover one up and "What if?" More to follow! It doesn't matter where you start the "What if?" game….you get there either way.

Let's use a famous process [meaning one the AP exam likes to ask questions about!], the Haber process. This is basically making ammonia for fertilizer production from the nitrogen in the air reacted with hydrogen gas. The hydrogen gas is obtained from the reaction of methane with water vapor. This process has saved millions from starvation!! The reaction is below:

Molar Mass:	(28.04)		(2.02)		(17.04)
Balanced Eq'n	N_2	+	$3\ H_2$	$\rightarrow$	$2\ NH_3$
mole:mole	**1**		**3**		**2**
# moles					
amount					

Suppose 25.0 kg of nitrogen reacts with 5.00 kg of hydrogen to form ammonia. What mass of ammonia can be produced? Which reactant is the limiting reactant? What is the mass of the reactant that is in excess?

Insert the masses in the amount row and find the number of moles of BOTH!

Molar Mass:	(28.02)		(2.02)		(17.04)
Balanced Eq'n	N_2	+	$3\ H_2$	$\rightarrow$	$2\ NH_3$
mole:mole	**1**		**3**		**2**
# moles	892 moles		2,475 moles		
amount	25,000 g		5,000 g		

WHAT IF I used up all the moles of hydrogen? I'd need **1/3** × 2,475 moles = 825 moles of nitrogen. Clearly I have EXCESS moles of nitrogen!! Therefore, hydrogen limits me.

OR

WHAT IF I used up all the moles of nitrogen? I'd need **3** × 892 moles = 2,676 moles of hydrogen. Clearly I don't have enough hydrogen, so it limits me!! Therefore nitrogen is in excess.

Either way, I've established that hydrogen is the limiting reactant so I modify the table:

That means I'll use up all the hydrogen but not all the nitrogen!

Molar Mass:	(28.02)	(2.02)	(17.04)
Balanced Eq'n	N_2 +	$3\ H_2$ →	$2\ NH_3$
mole:mole	**1**	**3**	**2**
# moles	825 mol used ~~892~~ moles	2,475 moles	1650 mol produced
amount	825 mol (28.02) = 23,116 g used ~~25,000~~ g 1,884 g excess!!	5,000 g	1650 mol (17.04) = 28,116 g produced

Here's the question again, let's clean up any sig.fig issues:
Suppose 25.0 kg of nitrogen reacts with 5.00 kg of hydrogen to form ammonia. (3 sig. fig. limit)

What mass of ammonia can be produced? 23,100 g produced = 23.1 kg (always polite to respond in the unit given).

Which reactant is the limiting reactant? hydrogen—once that's established, chunk the nitrogen amounts and let hydrogen be your guide!

What is the mass of the reactant that is in excess? 1,884 g = 1.88 kg excess nitrogen!!

Exercise 3.18 Stoichiometry: Limiting Reactant

Nitrogen gas can be prepared by passing gaseous ammonia over solid copper(II) oxide at high temperatures. The other products of the reaction are solid copper and water vapor. If a sample containing 18.1 g of NH_3 is reacted with 90.4 g of CuO, which is the limiting reactant ? How many grams of N_2 will be formed?

CuO is limiting; 10.6 g N_2

Theoretical Yield: The amount of product formed when a limiting reactant is completely consumed. This assumes perfect conditions and gives a maximum amount!! Not likely!

Actual yield: That which is realistic!

Percent yield: The ratio of actual to theoretical yield.

$$\frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\% = \text{Percent yield}$$

Exercise 3.19 **Calculating Percent Yield**

Methanol (CH_3OH), also called *methyl alcohol*, is the simplest alcohol. It is used as a fuel in race cars and is a potential replacement for gasoline. Methanol can be manufactured by combination of gaseous carbon monoxide and hydrogen. Suppose 68.5 kg $CO(g)$ is reacted with 8.60 kg $H_2(g)$. Calculate the theoretical yield of methanol. If 3.57×10^4 g CH_3OH is actually produced, what is the percent yield of methanol ?

Theoretical yield is 6.86×10^4 g
Percent yield is 52.3%