AP CHeMISTRY

2018 Summer Packet

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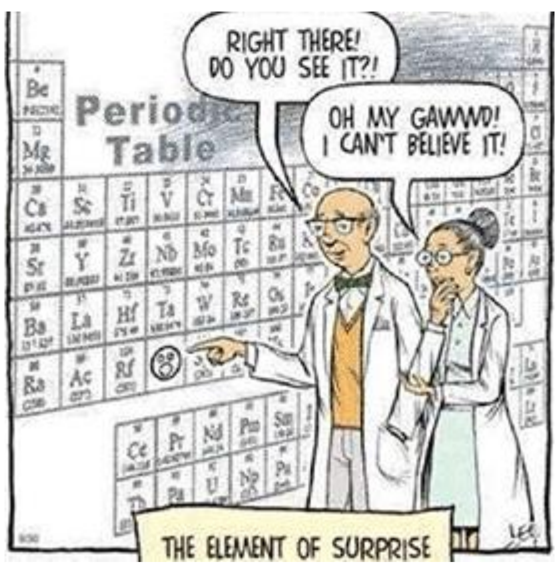
STUDENT NAME



AP CollegeBoard Advanced Placement Program

EAHS EASTWOOD ACADEMY

NATIONAL MATH + SCIENCE INITIATIVE

AP Chemistry Summer Packet

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TBD

Summer Overview

Welcome to your summer packet for AP Chemistry. It works to your advantage to make sure you are adept in numerical fluency (ratios, significant figures, dimension analysis, literal equations, simple algebra and measurements) before arriving to class the next school year.

The packet contains three chapters typical for a college freshman in the application of chemistry to the following topics:

1. Chemical Foundations
2. Atoms, Molecules, and Ions
3. Stoichiometry
4. Gases

Grading

Do not lose this packet. It is for a series of grades!

MCQ-grades and FRQ-grades

Course Materials

You are expected to show the work – the answers are given! Use the packet and annotate. Create your questions for clarifications.

The FRQs for stoichiometry must be written on a blank sheet of paper with your name. Box all final answers and you must show work – final answers are only worth 1 pt.

The packet also contains a safety contract that you should peruse and sign. **We have lab the first week of school.**

Checklist Planner

- □ June 01-15 Chemical Foundations
- □ June 16-30 Atoms, Molecules, and Ions
- □ July 01-15 Stoichiometry
- □ July 16-31 Stoichiometry
- □ August 1-15 Gases
- □ August 20-26 FRQs on Stoichiometry

AUGUST 28 – First Day of Class: Packets and FRQs are DUE!

FOR YOUR INFORMATION – Regarding the Class.

Grading Policy

Grading policies are based on these three (3) categories:

MCQs	35%
FRQs	35%
Lab./Test/Quiz	30%

Homework Policy

Unfinished work during class becomes homework. Homework may fall under the category of MCQ or FRQ. Demerits for late work are as follows:

Demerits for submission of late homework:

1 day – highest grade 89% (assuming all items are correct)
2 days – highest grade 79%
3 days – highest grade 69%

Laboratory work and summative assessments (30%)

Tests are 20-minute sit-down assessments and are cumulative in nature. And quizzes are 10 minute sit-downs. Both assessments for the semester simulate the actual AP exams and comprise two (2) sections: Section 1: Multiple Choice Questions (**MCQs**) and Section 2: Free Response Questions (**FRQs**), unless otherwise specified.

Free-Response Questions (FRQs) (35%)

Category assessing writing skills designed for the AP test. FRQs may come in the form of different activities (warps, exi-tickets, reports, quizzes, and tests)

Multiple Choice Questions (MCQs) (35%)

Category assessing content (i.e. what you know) designed for the AP test. MCQs may come in the form of different activities (warps, exi-tickets, quiz, and tests)

Reminder: Student Expectations

Please write your degree plan for college and expectations of the course:

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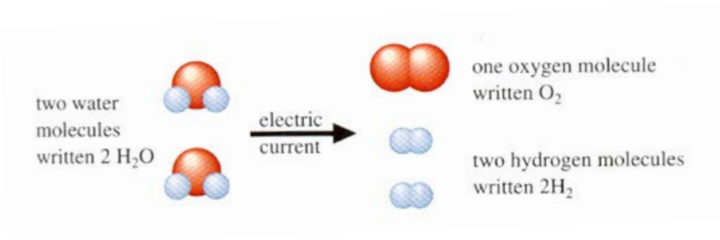
Additional Information

You have **15 days** to drop the course.

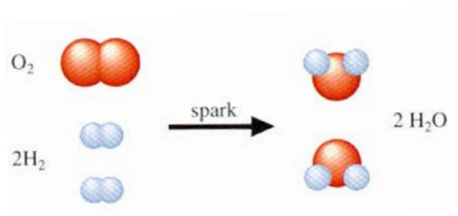
AP* Chemistry
Chemical Foundations

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

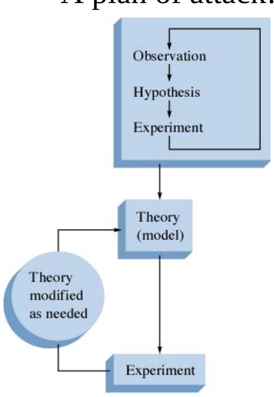


<u>Chemistry</u> – is defined as the study of matter and energy and more importantly, the changes between them

- **<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

The Scientific Method

- A plan of attack!



The fundamental steps of the scientific method

Steps in the Scientific Method

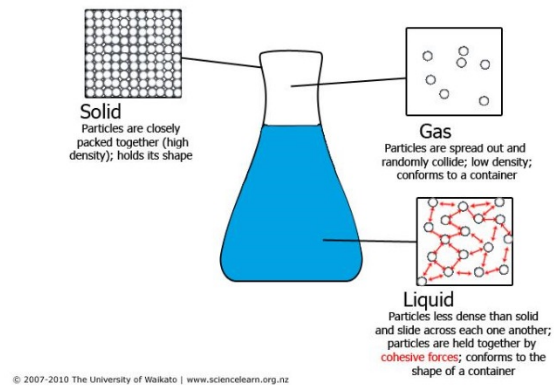
1. *Making observations.* Observations may be *qualitative* (the sky is blue; water is a liquid) or *quantitative* (water boils at 100°C; a certain chemistry book weighs 2 kilograms). A qualitative observation does not involve a number. A quantitative observation (called a **measurement**) involves both 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.

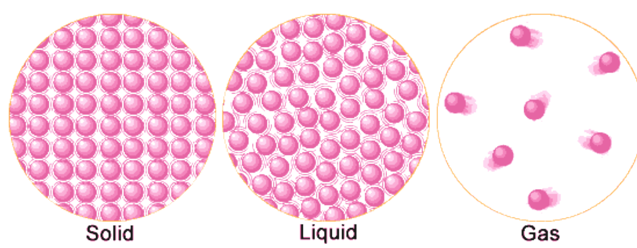
- Good experimental design coupled with repetition is key!
  <u>Theory</u> – hypotheses are assembled in an attempt at *explaining* “why” the “what” happened.
- **<u>Model</u>** – we use many models to explain natural phenomenon – when new evidence is found, the model changes!

- **Robert Boyle**
  - loved to experiment with air
  - created the first vacuum pump
  - coin and feather fell at the same rate due to gravity
  - in a vacuum there is no air resistance to impede the fall of either object!
  - Boyle defined elements as anything that cannot be broken into simpler substances.

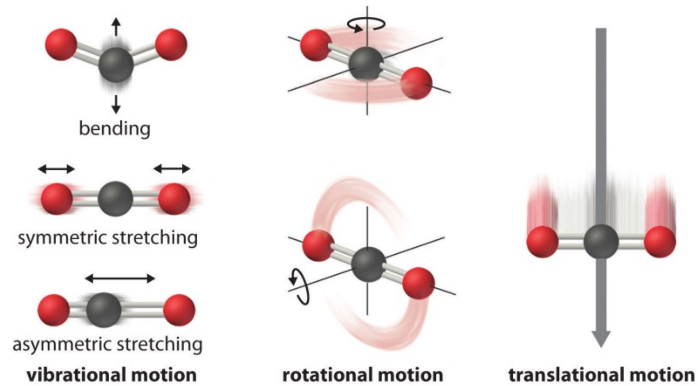
    Boyle’s Gas Law: $P_1V_1 = P_2V_2$

- **<u>Scientific Laws</u>** – 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.
  - **<u>Law of Conservation of Mass</u>** – mass reactants = mass products
  - **<u>Law of Conservation of Energy</u>** – (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





- **<u>Galileo</u>** – forced to recant his astronomical observations in the face of strong religious resistance
- **<u>Lavoisier</u>** – “father of modern chemistry”; beheaded due to political affiliations.
- The need for better <u>explosives</u>; (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



Units of Measure

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

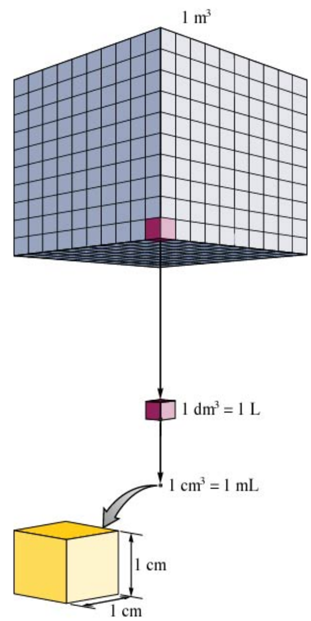
Length	A dime is 1 mm thick. A quarter is 2.5 cm in diameter. The average height of an adult man is 1.8 m.
Mass	A nickel has a mass of about 5 g. A 120-lb person has a mass of about 55 kg.
Volume	A 12-oz can of soda has a volume of about 360 mL.

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 THE UNITS AND PREFIXES shown in **BLUE**!!!



- Volume – derived from length; consider a cube 1m on each edge $\therefore$ 1.0 m^3
  - A decimeter is 1/10 of a meter so
    $(1m)^3 = (10\ dm)^3 = 10^3\ dm^3 = 1{,}000\ dm^3$
    $1dm^3$ = 1 liter (L) and is slightly larger than a quart also
    $1dm^3 = 1\ L = (10\ cm)^3 = 10^3 cm^3 = 1{,}000\ cm^3 = 1{,}000\ mL$
    AND $1\ cm^3 = 1\ mL = 1$ gram of H_2O (at 4ºC if you want to be picky)

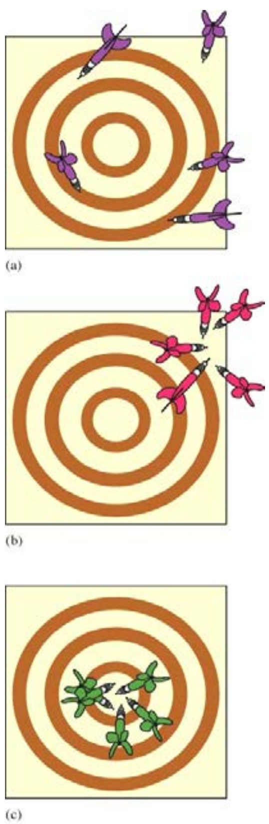
Mass vs. Weight – chemists are quite guilty of using these terms interchangeably.

- mass (g or kg) – a measure of the resistance of an object to a change in its state of motion (i.e. exhibits inertia); the quantity of matter present
- weight (a force $\therefore$ has units of Newtons) – the response of mass to gravity; since all of our measurements will be made here on Earth, we consider the acceleration due to gravity a constant so we'll use the terms interchangeably as well *although* it is technically incorrect! We "weigh" chemical quantities on a **balance NOT a scale!!**

<u>Gravity</u> – 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 $\therefore$ "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 $\therefore$ you would experience $\frac{1}{6}$ the gravitational field you experience on Earth and $\therefore$ you'd WEIGH $\frac{1}{6}$ of what you weigh on Earth.

Precision and Accuracy



The results of several dart throws show the difference between precise and accurate.

(a) Neither 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).

- **<u>Accuracy</u>** – correctness; agreement of a measurement with the true value
- **<u>Precision</u>** – reproducibility; degree of agreement among several measurements.
- **<u>Random or indeterminate error</u>** – equal probability of a measurement being high or low
- **<u>Systematic or determinate error</u>** – occurs in the same direction each time

Exercise 1 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).

Significant Figures and Calculations

Determining the Number of Significant Figures (or Digits) in a Measurement

- Nonzero digits are significant. (Easy enough to identify!)
- A zero is significant IF and ONLY IF it meets one of the conditions below:
  - The zero in question is "terminating AND right" of the decimal [must be both]
  - The zero in question is "sandwiched" between two significant figures
- Exact or counting numbers have an ∞ amount of significant figures as do fundamental constants (never to be confused with derived constants)

Exercise 2 Significant Figures (SF)

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

a. A student's extraction procedure on a sample of 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

Reporting the Result of a Calculation to the Proper Number of Significant Figures

- When × and ÷ , the term with the **least** number of *significant figures* (∴least accurate measurement) determines the number of **maximum** number of significant figures in the answer. (It's helpful to underline the digits in the least significant number as a reminder.)

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

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

$$\begin{array}{ll} 12.11 & \\ 18.0 & \leftarrow \text{limiting term (only 1 decimal place)} \\ \underline{1.013} & \\ 31.123 \xrightarrow{\text{corrected}} 31.1 & \text{(limits the overall answer to only one decimal place)} \end{array}$$

- pH – the *number of significant figures in least accurate measurement* determines *number decimal places* on the reported pH (usually explained in the appendix of your text)

Rounding Guidelines for the AP Exam and This Course:

- Round ONLY at the end of all calculations (keep the numbers in your calculator)
- Examine the significant figure one place beyond your desired number of significant figures.
  IF > 5 round up; < 5 drop the remaining digits.
- Don't "double round"! Example: The number 7.348 rounded to 2 SF is reported as 7.3
  In other words, DO NOT look beyond the 4 after the decimal and think that the 8 rounds the 4 up to a five which in turn makes the final answer 7.4.
  [Even though you may have conned a teacher into rounding your final average this way before!]

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

Example: Consider a straight pin measuring 2.85 cm in length. Calculate its length in inches.

Start with a conversion factor such as 2.54 cm = 1 inch $\therefore$ you can write TWO Conversion factors: $\frac{1 \text{ in}}{2.54 \text{ cm}}$ or $\frac{2.54 \text{ cm}}{1 \text{ in}}$. Why is this legal? Both quantities represent the exact same "thing" so the conversion factor is actually equal to "1".

To convert the length of the pin from cm to inches, simply multiply your given quantity by a conversion factor you engineer so that it "cancels" the undesirable unit and places the desired unit where you want it. For our example, we want inches in the numerator so our numerical answer is not reported in reciprocal inches! Thus,

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

Let's practice!

Exercise 3
A pencil is 7.00 in. long. Calculate the length in centimeters?

17.8 cm

Exercise 4
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 5
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 $\rightarrow$ meters $\rightarrow$ yards $\rightarrow$ miles

To proceed in this way, we need the following equivalence statements (conversion factors):

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

6.22 mi

Exercise 6
The speed limit on many highways in the United States is 55 mi/h. What number would be posted if expressed in kilometers per hour?

88 km/h

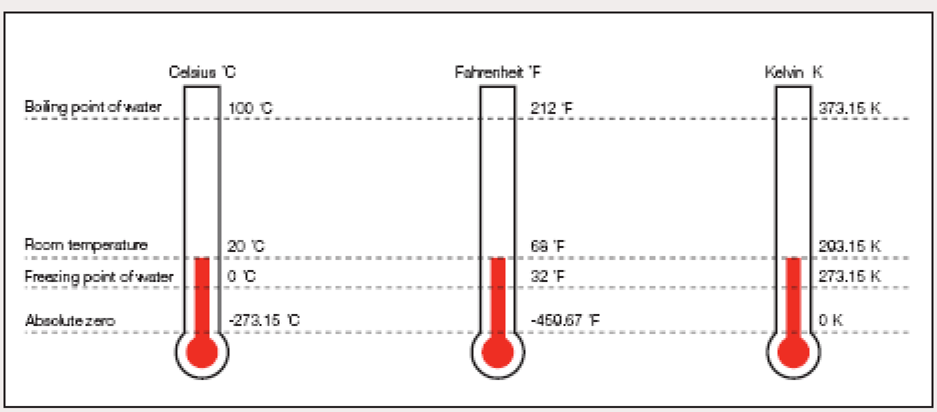
Exercise 7
A Japanese car is advertised as having a fuel economy of 15 km/L. Convert this rating to miles per gallon.

35 mi /gal

Temperature

I suspect you are aware there are **three** temperature scales commonly in use today. A comparison follows:

Known Temperature				Required Temperature	Formulae
Celsius	°C	to	°F	Fahrenheit	°F = (1.8 X °C) + 32
Celsius	°C	to	K	Felvin	K = °C + 273.15
Fahrenheit	°F	to	°C	Celsius	°C = (°F - 32)/1.8
Fahrenheit	°F	to	K	Kelvin	K = °F + 459.67)/1.8
Kelvin	°K	to	°C	Celsius	°C = K-273.15
Kelvin	°K	to	°F	Fahrenheit	°F = (1.8 X K) - 459.67



Notice a degree of temperature change on the Celsius scale represents the same quantity of change on the Kelvin scale.

Density

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

Exercise 8 Determining Density

A chemist, trying to identify the main component of a compact disc cleaning fluid, determines that 25.00 cm^3 of the substance has a mass of 19.625 g at 20°C. Use the information in the table below to identify which substance may serve as the main component of the cleaning fluid. Justify your answer with a calculation.

Compound	Density (g/cm^3) at 20°C
Chloroform	1.492
Diethyl ether	0.714
Ethanol	0.789
Isopropyl alcohol	0.785
Toluene	0.867

Density = 0.7850 g / cm^3 ∴ isopropyl alcohol

Classification of Matter

States of Matter (mostly a vocabulary lesson)

- **Be very, very clear that changes of state involve altering IMFs <u>not</u> altering actual chemical bonds!!**
- **<u>solid</u> –** rigid; definite shape and volume; *molecules close together vibrating about fixed points ∴ virtually incompressible*
- **<u>liquid</u>** – 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*
- **<u>gas</u>** – 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*
  - **<u>vapor</u>** – the gas phase of a substance that is normally a solid or liquid at room temperature
  - **<u>fluid</u>** – that which can flow; gases and liquids

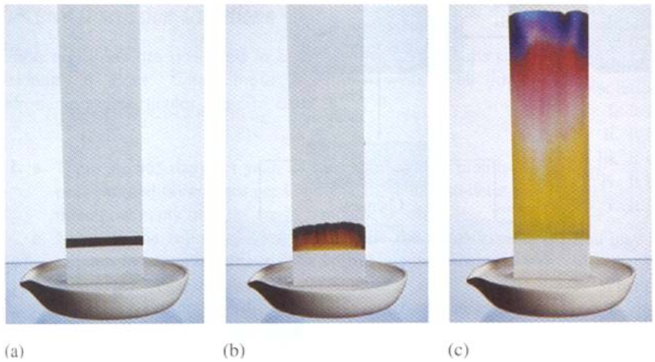
Table 1.5 Densities of Various Common Substances* at 20°C

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

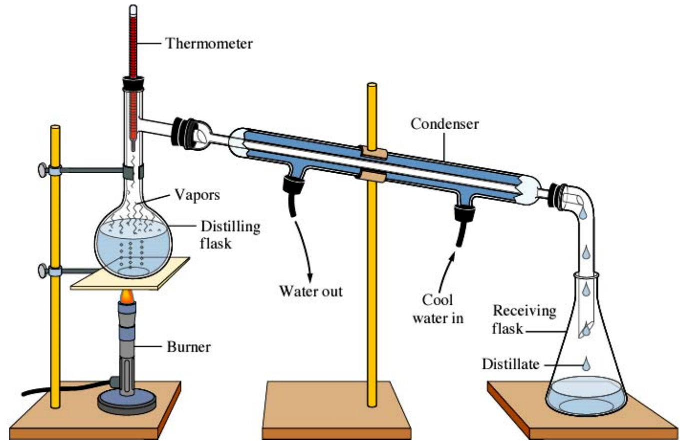
- **<u>Mixtures</u>** – can be **physically** separated
  - **<u>homogeneous</u>** – have visibly <u>in</u>distinguishable parts, solutions including air
  - **<u>heterogeneous</u>** – have visibly distinguishable parts
  - <u>means of physical separation include</u>: filtering, fractional crystallization,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

Paper Chromatography:

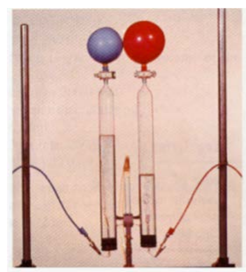


Paper chromatograph of ink. (a) A line of the mixture to be separate is placed at one end of a sheet of porous paper. (b) The paper acts as a wick to draw up the liquid. (c) The component with the weakest attraction for the paper travels faster than those that cling to the paper.

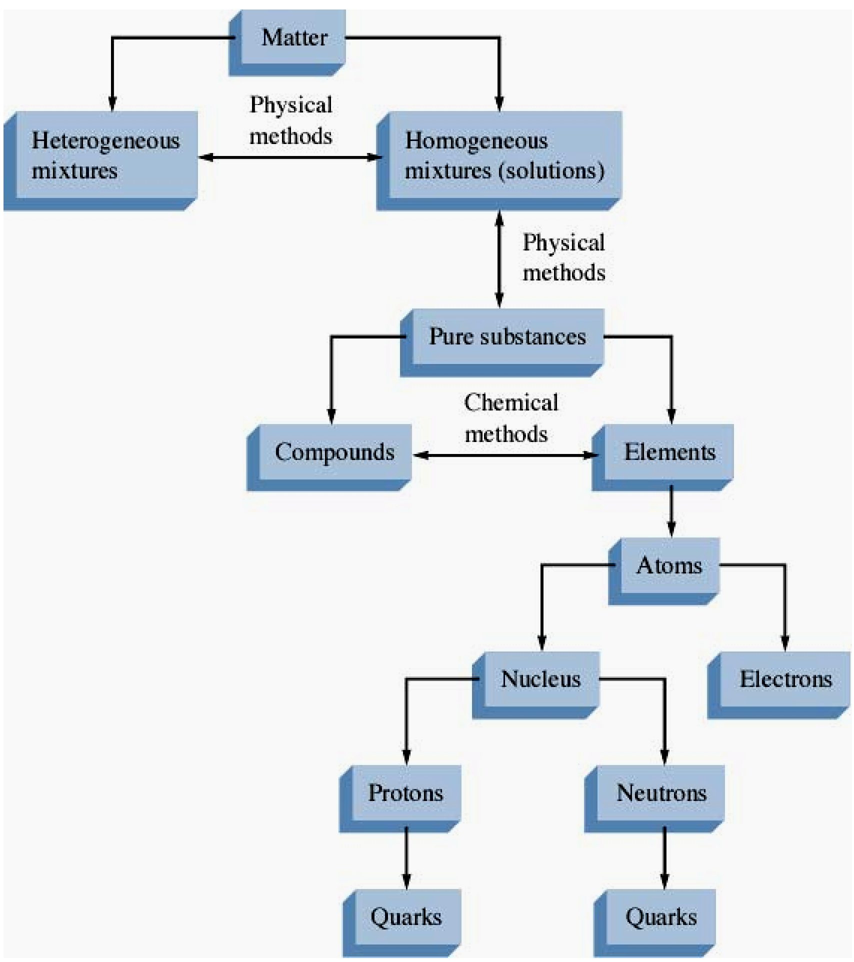
Distillation:



- **<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 further 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

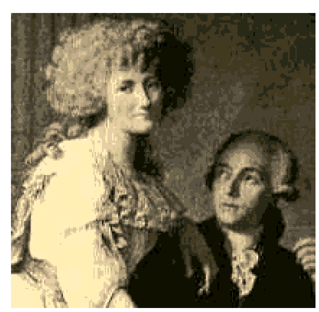
The highest honor given by the American Chemical Society. He discovered oxygen. Ben Franklin got him interested in electricity and he observed graphite conducts an electric current. Politics forced him out of England and he died in the US in 1804. The back side, pictured below was given to Linus Pauling 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.

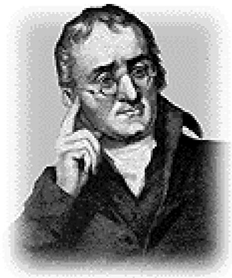


THE LAW OF CONSERVATION OF MASS:

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

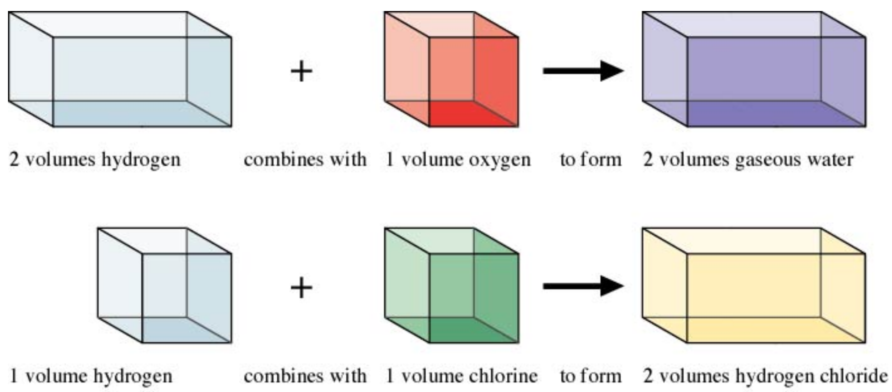
Dalton's ATOMIC THEORY OF MATTER: (based on knowledge *at that time):*

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.*

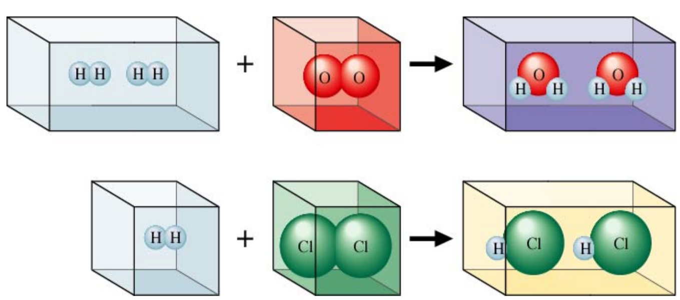
- 1809 Joseph Gay-Lussac, French—performed experiments [at constant temperature and pressure] and measured volumes of gases that reacted with each other.



- 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.

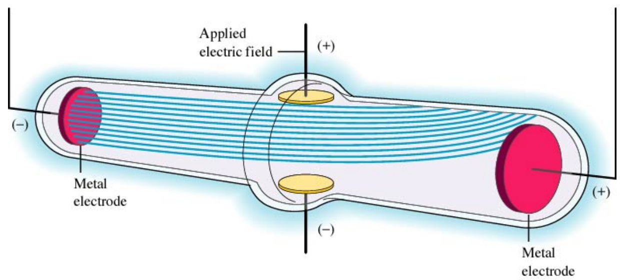


2.4 EARLY EXPERIMENTS TO CHARATERIZE THE ATOM

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

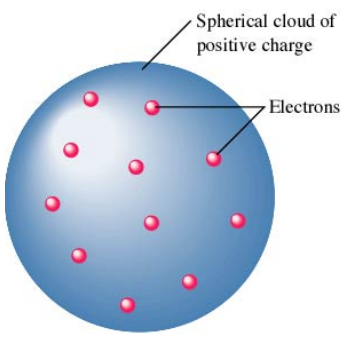
- J.J. Thomson, English (1898-1903)—found that when high voltage was applied to an evacuated tube, a "ray" he called a cathode ray [since it emanated from the (-) electrode or cathode when YOU apply a voltage across it] was produced.
  - The ray was produced at the (-) electrode
  - Repelled by the (-)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



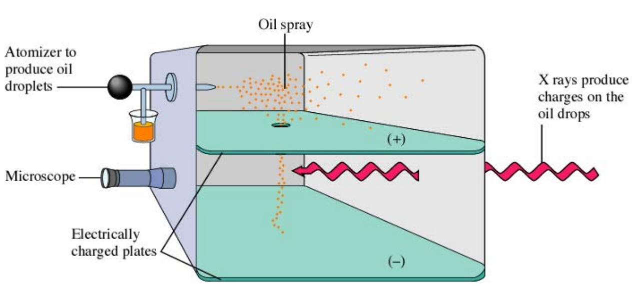
  - He then measured the deflection of beams of e^- to determine the **charge-to-mass ratio**

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

  - *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]



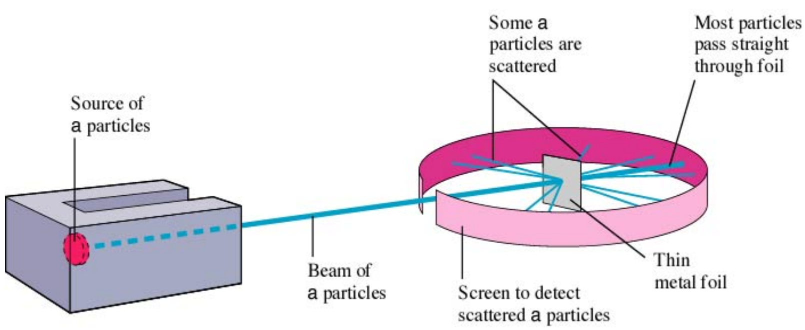
- 1909 Robert Millikan, American—University of Chicago, sprayed charged oil drops into a chamber. Next, he halted their fall due to gravity by adjusting the voltage across 2 charged plates. Now the voltage needed to halt the fall and the mass of the oil drop can be used to calculate the charge on the 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.

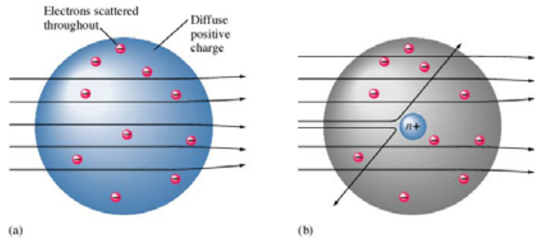
THE NUCLEAR ATOM

- 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!



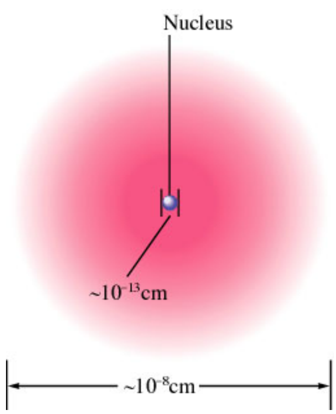
2.5 THE MODERN VIEW OF ATOMIC STRUCTURE: AN INTRODUCTION

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.



- **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.

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

mass number → $^{A}_{Z}\text{X}$ ← 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

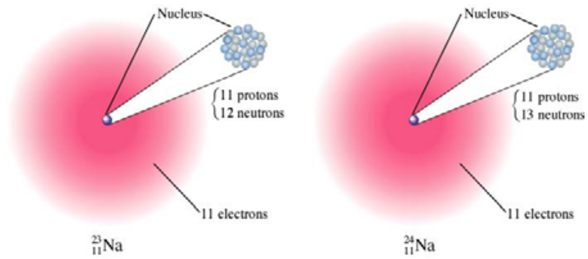
Exercise 2.2 Writing the Symbols for Atoms

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

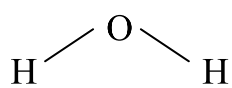


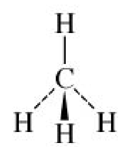
2.6 MOLECULES AND IONS

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

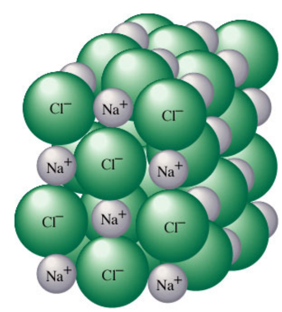
H — O — H





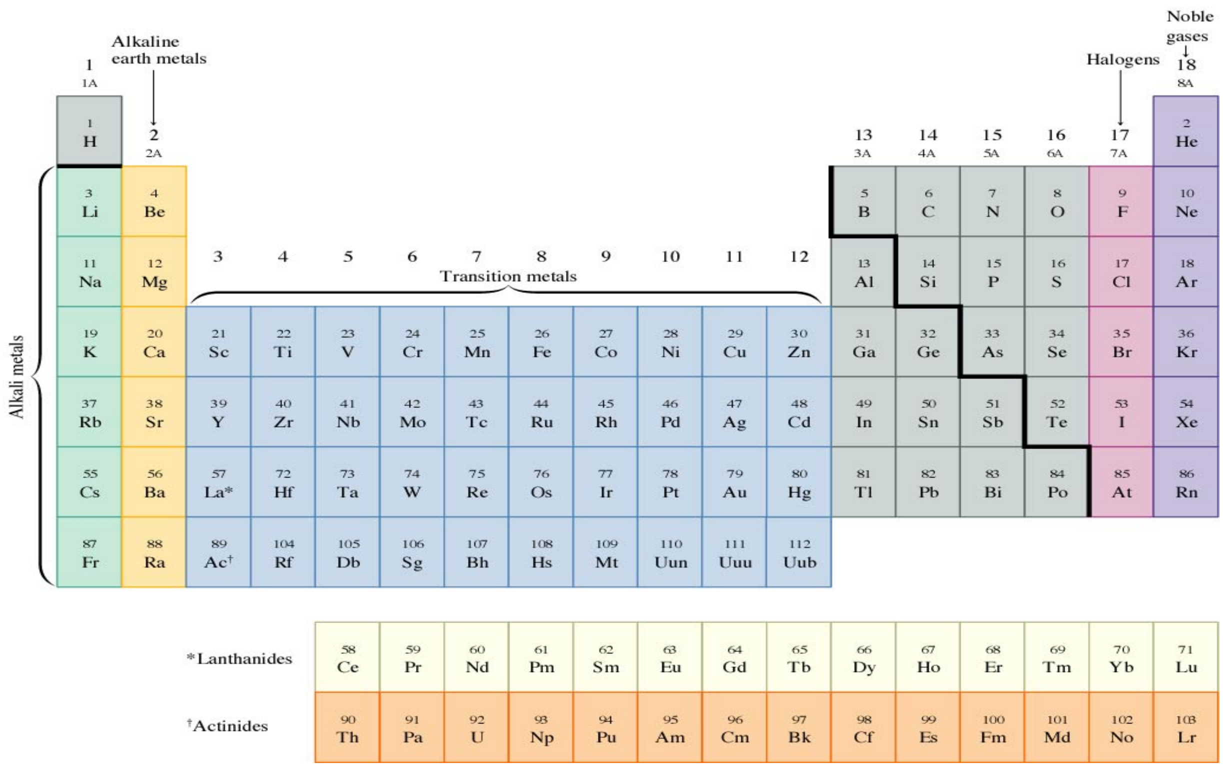
Methane

- **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.



(a)

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

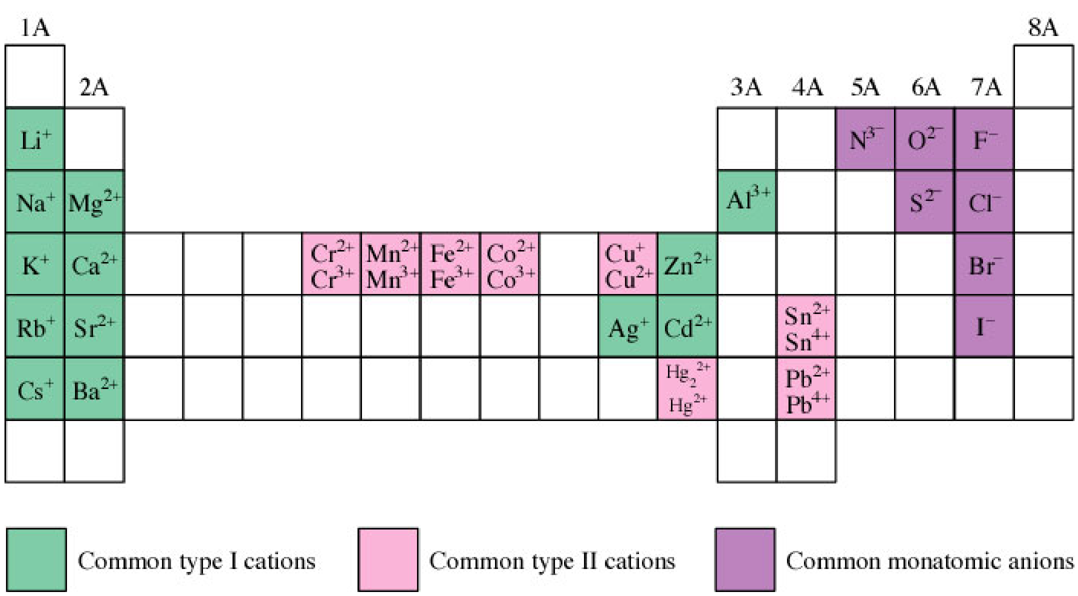
- monatomic, metal, cation → simply the name of the metal from which it is derived. **Al^{3+} is the aluminum ion**
- transition metals form *more than one* ion; Roman Numerals (in) follow the ion's name. **Cu^{2+} is copper(II)** (Yeah, the no space thing between the ion's name and (II) looks strange,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

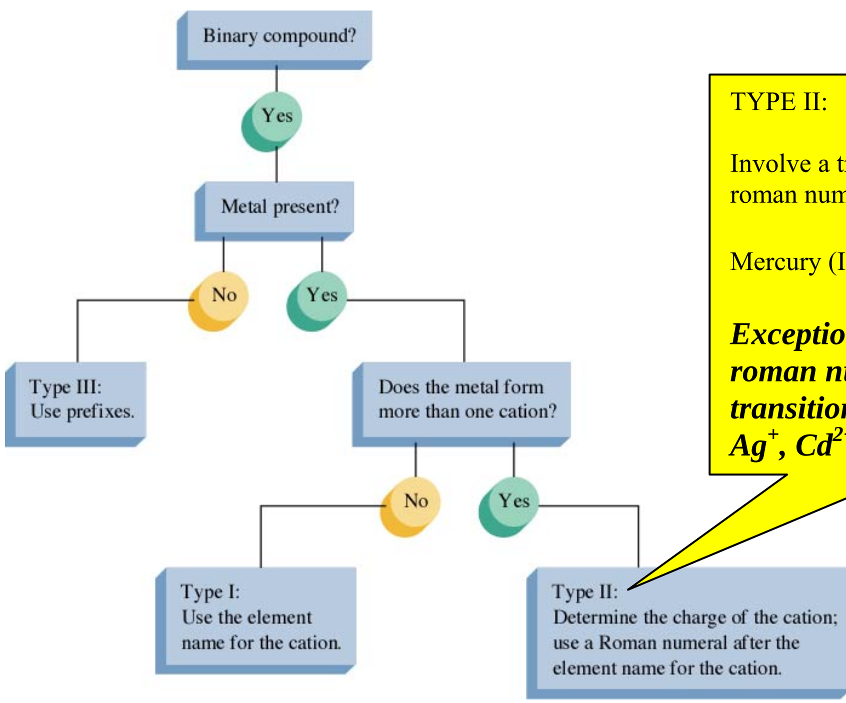
a. cesium fluoride
b. aluminum chloride
c. lithium hydride

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)

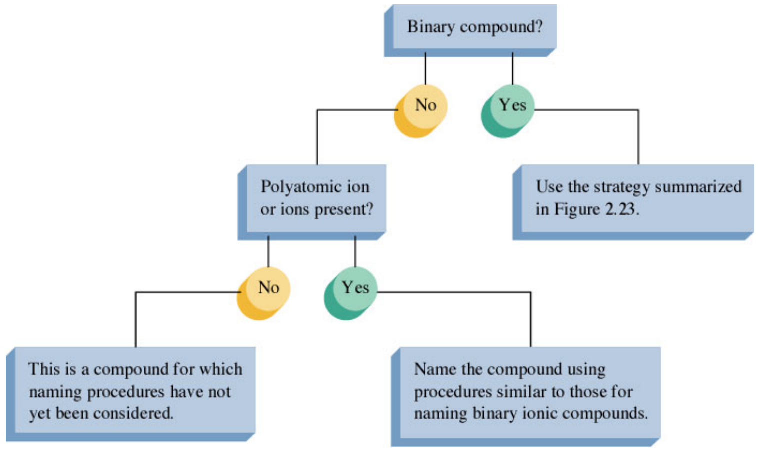
- use prefixes!!! Don't forget the –ide ending as well.

Exercise 2.7 Naming Type III Binary Compounds

Name each of the following compounds.

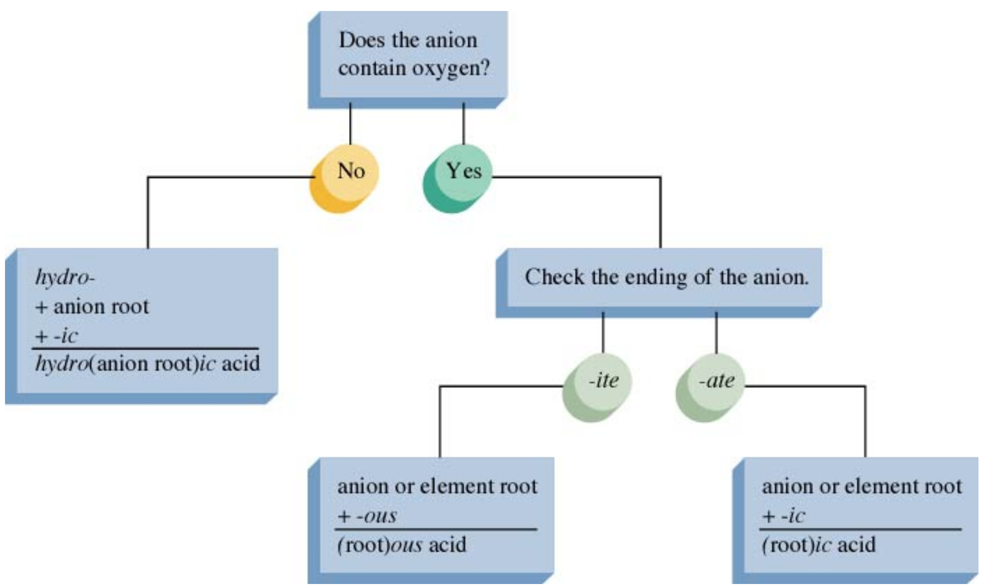
a. $PC1_5$ b. $PC1_3$ c. SF_6 d. SO_3 e. SO_2 f. CO_2

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

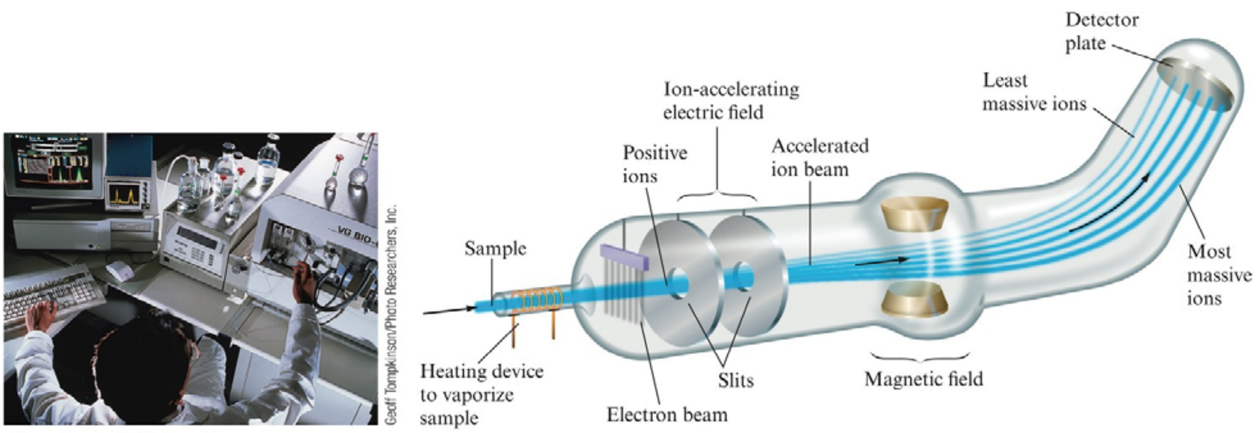
a. VF_5
b. O_2F_2
c. Rb_2O_2
d. Ga_2O_3

NATIONAL MATH + SCIENCE INITIATIVE

AP* Chemistry
Stoichiometry

ATOMIC MASSES

- **^{12}C—Carbon 12**—In 1961 it was agreed that this isotope of carbon would serve as the standard used to determine all other atomic masses 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.

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

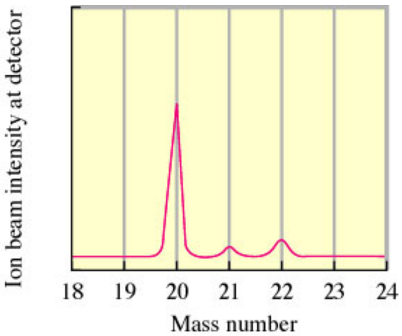
    Exact by definition

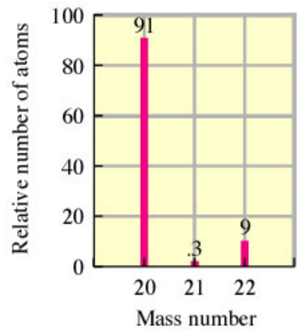
- **average atomic masses**—atoms have masses of whole numbers, HOWEVER samples of quadrillions of atoms have a few that are heavier or lighter [isotopes] due to different numbers of neutrons present
- **percent abundance**--percentage of atoms in a natural sample of the pure element represented by a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

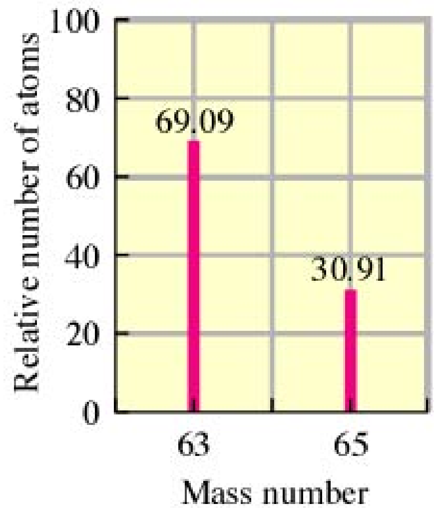
- **<u>mass spectrometer to determine isotopic composition</u>**—load in a pure sample of natural neon or other substance. The areas of the "peaks" or heights of the bars indicate the relative abundances of $^{20}_{10}Ne$, $^{21}_{10}Ne$, and $^{22}_{10}Ne$





Exercise 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 **calculate** 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

THE MOLE

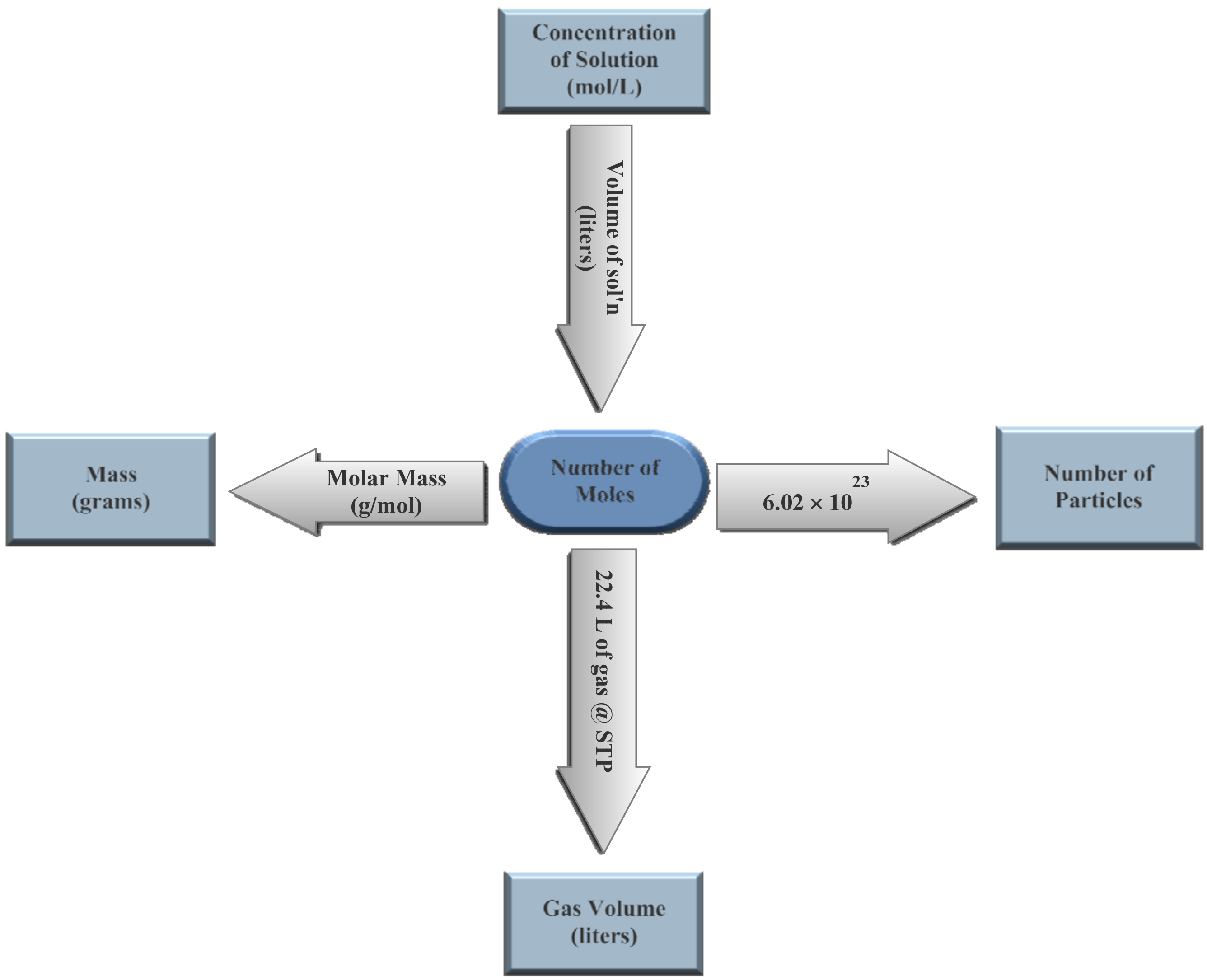
- **<u>mole</u>**—the number of C atoms in exactly 12.0 grams of ^{12}C; also a number, 6.02×10^{23} just as the word "dozen" means 12 and "couple" means 2.
- **<u>Avogadro's number</u>**—6.02×10^{23}, the number of particles in a mole of anything

DIMENSIONAL ANALYSIS DISCLAIMER: 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. 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 themselves did, once they got to college, and made lots of new friends once word got out they had this "easy 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.



Exercise 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*. **Calculate** the mass in grams of a sample of americium containing six atoms.

2.42×10^{-21} g

Exercise 3 **Determining Moles of Atoms**

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

0.371 mol Al
2.23×10^{23} atoms

Exercise 4 **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

MOLAR MASS AND FORMULA WEIGHT

- **molar mass, *MM***--the sum of all of the atomic masses in a given chemical formula in units of g/mol. It is also equal mass in grams of Avogadro's number of molecules; i.e. the mass of a mole
- **empirical formula**--the 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}. **Just use "molar mass" for all formula masses.**
- **A WORD ABOUT SIG. FIG.'s—**It is correct to "pull" from the periodic table the least number of sig. figs for your MM's as are in your problem—just stick with 2 decimal places for all MM's —much simpler!

Exercise 5 **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. **Calculate** the number of moles of juglone present in this sample.

a. 174.1 g
b. 8.96×10^{-5} mol juglone

Exercise 6 **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. **Calculate** the mass in grams of this sample. **Calculate** the mass of the CO_3^{2-} ions present.

a. 100 g/mol
b. 486 g; 292g CO_3^{2-}

Exercise 7 **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.

(a) **Calculate** the number of molecules of isopentyl acetate released in a typical bee sting.

(b) **Calculate** the number of carbon atoms present.

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

ELEMENTS THAT EXIST AS MOLECULES

Pure hydrogen, nitrogen, oxygen and the halogens exist as **DIATOMIC** molecules *under normal conditions.* MEMORIZE!!! Be sure you compute their molar masses as diatomics. We lovingly refer to them as the "gens", "Hydro**gen**, oxy**gen,** nitro**gen** & the halo**gen**s!"

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

PERCENT COMPOSITION OF COMPOUNDS

There are two common ways of describing the composition of a compound: 1) in terms of the number of its constituent atoms and 2) in terms of the percentages (by mass) of its elements.

Percent Composition (by mass): The 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.* Remember, all percent calculations are simply $\frac{\text{part}}{\text{whole}} \times 100\%$

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

Consider ethanol, C_2H_5OH

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

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

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

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

NEXT, THE MASS PERCENT CAN BE CALCULATED:

$$\text{Mass percent of C} = \frac{24.02 \text{ g C}}{46.08 \text{ g}} \times 100\% = 52.14\%$$

Repeat for the H and O present.

Exercise 8 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. **Calculate** the mass percent of each element in carvone.

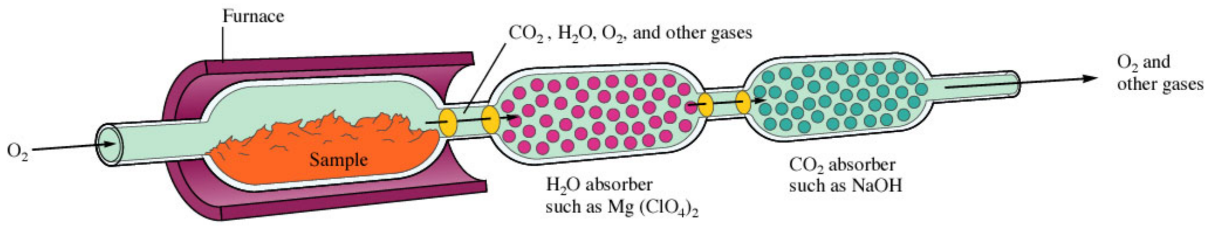
C = 79.96%
H = 9.394%
O = 10.65%

Exercise 9 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 $C_{14}H_{20}N_2SO_4$. Calculate the mass percent of each element.

C = 53.82%
H = 6.47%
N = 8.97%
S = 10.26%
O = 20.49%

DETERMINING THE FORMULA OF A COMPOUND



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

- **Calculating empirical and molecular formulas:** empirical formulas represent the *simplest or smallest ratio of elements within a compound* while molecular formulas represent the *actual numbers of elements within a compound*. The empirical mass is the **least common multiple** of the molar mass.

  Example: CH_2O is the empirical for a carbohydrate—get it? "carbon waters".

  Anyway, glucose is a perfect example of a carbohydrate (a sugar to be exact) with an empirical molar mass of 12 + 2(1) + 16 = 30 g/mol and since glucose is 6 units of CH_2O which is equivalent to $(CH_2O)_6$ or $C_6H_{12}O_6$; the empirical mass of 30 is also multiplied by 6. Thus the MM of glucose is 180 g/mol.

- Make your problem solving life easy and ***assume a 100 gram sample if given %'s***—that way you can convert the percents given directly into grams and subsequently into moles in order to simplify your life!

Other twists and turns occurring when calculating molar masses involve:

- **hydrates**—waters of hydration or "dot waters". They count in the calculation of molar masses for hydrates and used to "cement" crystal structures together

- **anhydrous**—means *without* water—just to complete the story—just calculate the molar masses of anhydrous substances as you would any other substance

Example:

A compound is composed of carbon, nitrogen and hydrogen. When **0.1156 g of this compound** is reacted with oxygen [a.k.a. "burned in air" or "combusted"], 0.1638 g of carbon dioxide and 0.1676 g of water are collected. Determine the empirical formula of the compound.

So, Compound + O_2 $\rightarrow$ oxides of what is burned. In this case Compound + O_2 $\rightarrow CO_2 + H_2O + N_2$
(clearly not balanced)

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

0.1638 g CO_2 ÷ 44.01 g/mol CO_2 = 0.003781 moles of CO_2 = 0.003781 moles of C (why?)

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

So, 0.1676 g H_2O ÷ 18.02 g/mol H_2O = 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 while there are 2 lungs per body, 2 atoms H in water and so on…] thus, DOUBLE THE NUMBER OF MOLES of H_2O GIVES THE NUMBER OF MOLES OF HYDROGEN!! moles H = 2 × 0 .009301 moles of H_2O = 0.01860 moles of H

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

$$\text{grams C} = 0.003781 \text{ moles C} \times 12.01 \frac{\text{g}}{\text{mol}} = 0.04540 \text{ grams C}$$

$$\text{grams H} = 0.01860 \text{ moles H} \times 1.01 \frac{\text{g}}{\text{mol}} = \underline{0.01879 \text{ grams H}}$$

Total grams : **0.06419 total grams accounted for thus far**

What to do next? SUBTRACT!

0.1156 g sample – **0.06419 total grams accounted for thus far** = grams N left = 0.05141 g N so….

$$0.05141 \text{ g N} \div 14.01 \frac{\text{g}}{\text{mol}} = \mathbf{0.003670 \text{ moles N}}$$

Next, realize that chemical formulas represent **mole to mole ratios**, so…divide the number of moles of each by the smallest # of moles for 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. Clear as mud? WATCH THE SCREENCAST!!

Element	# moles	ALL Divided by the smallest (0.003670 moles)
C	**0.003781**	**1**
H	**0.01860**	**5**
N	**0.003670**	**1**

Therefore, the correct EMPIRICAL formula based on the data given is CH_5N.

Finally (this is drumroll worthy), IF we are told that the MM of the original substance is 31.06 g/mol, then simply use this relationship:

(Empirical mass) × n = MM

(12.01 + 5.05 + 14.01) × n = 31.07 g/mol ∴ n = 0.999678

This is mighty close to 1.0! Thus, the empirical formula and the molecular formula are one and the same.

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 10

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_2 C1
Molecular formula = $C_2H_4C1_2$

Exercise 11

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 12
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 = $C_8H_{10}N_4O_2$

BALANCING CHEMICAL EQUATIONS

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*]. A chemical reaction transforms elements and compounds into new substances. A *balanced chemical equation* shows the relative amounts of reactants [on the left] and products [on the right] by molecule or by mole.

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

Subtle details:
- *s, l, g, aq*—state symbols that correspond to solid, liquid, gas, aqueous solution
- NO ENERGY or TIME is alluded to
- Antoine Lavoisier (1743-1794)—The Law of Conservation of Matter: *matter can be neither created nor destroyed* ☞ this means you having to "balance equations" is entirely his fault!!

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 $\rightarrow$ M_aX_b
- CH and/or O + O_2 $\rightarrow$# $CO_2(g) + H_2O(g)$
- H_2CO_3 [any time formed!] $\rightarrow$ $CO_2 + H_2O$; 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 $\rightarrow$ 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 13

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. 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.

$$(NH_4)_2Cr_2O_7(s) \rightarrow Cr_2O_3(s) + N_2(g) + 4H_2O(g)$$

(4 × 2) H **(4 ×2) H**

http://www.youtube.com/watch?v=CW4hN0dYnkM

Exercise 14

At 1000ºC, 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.

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

STOICHIOMETRIC CALCULATIONS: AMOUNTS OF REACTANTS AND PRODUCTS

Stoichiometry – The study of quantities of materials consumed and produced in chemical reactions. 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!

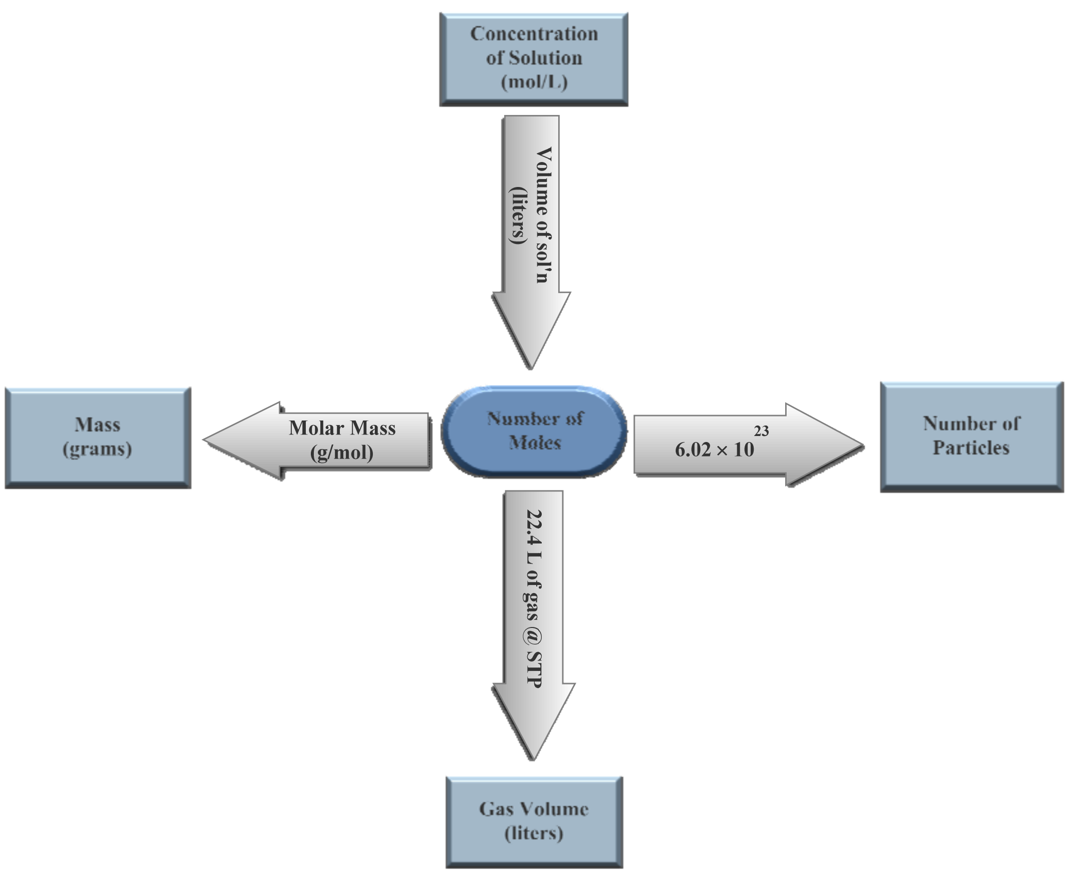
It's time to repeat my dimensional analysis disclaimer.

DIMENSIONAL ANALYSIS DISCLAIMER: 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:

Example: Calculate the mass of oxygen will react completely with 96.1 grams of propane?
[notice all words—you supply the chemical formulas!]

Molar Mass:	(44.11)	(32.00)		(44.01)	(18.02)
Balanced Eq'n	C_3H_8	+ 5 O_2	$\rightarrow$	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 in which you are interested.

3. Balance the equation! Examine the coefficients on the balanced equation, they ARE the mole:mole ratios! Isolating them helps you internalize the mol:mol until you get the hang of this.

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	$\rightarrow$	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. Calculate the number of moles of something, anything! Use the mole map. Start at 96.1 grams of C_3H_8, divide the 96.1 g [against the arrow on the mole map] by molar mass to calculate 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 mol amount you calculate is not a "1", just divide appropriately to make it "1" before moving on to calculate the moles of all the rest!] 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

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

Molar Mass:	(44.11)		(32.00)		(44.01)		(18.02)
Balanced Eq'n	C_3H_8	+	5 O_2	$\rightarrow$	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		

What if part (c) asked you to calculate how many water molecules are produced?

Use the mole map , start in the middle with 8.71 mol water × $6.02 \times 10^{23}\ \dfrac{\text{molecules}}{\text{mol}} = 5.24 \times 10^{24}$ molecules of water.

Try these two exercises with whichever method you like best!

Exercise 15
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?

920. g

Exercise 16
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 $Mg(OH)_2$? Justify your answer.

$Mg(OH)_2$

CALCULATIONS INVOLVING A LIMITING REACTANT

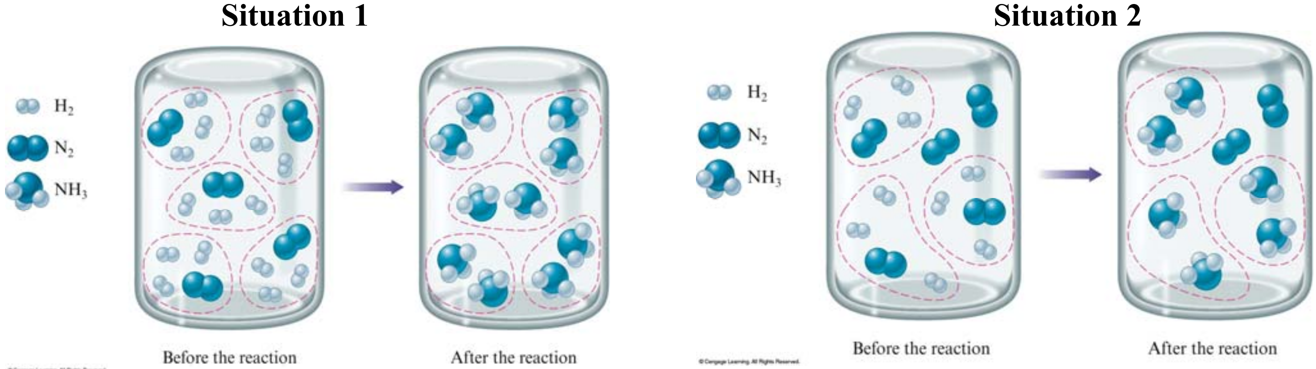
Ever notice how hot dogs are sold in packages of 10 while the buns come in packages of 8? What's up with that?! The bun is the limiting reactant and limits the hot dog production to 8 as well! The limiting reactant [or reagent] is the one consumed most entirely in the chemical reaction.

Let's use a famous process [meaning one the AP exam likes to ask questions about!], the Haber process. This reaction is essentially 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 shown below.

Exercise17
Examine the particle views and explain the differences between the two situations pictured below with regard to what is or is not reacting and total yield of ammonia.

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

Situation 1 **Situation 2**



<u>Plan of attack</u>: First, you must realize that you even ***need*** a plan of attack! **IF ever you are faced with TWO starting amounts of matter reacting, you have entered "The Land of Limiting Reactant".**

When faced with this situation …calculate the number of moles of everything you are given. Set up your table like before, only now you'll have TWO amounts and thus TWO # 's of moles to get you started.

Cover one set of moles up (pretending you only had one amount to work from) and ask yourself, "What if all of these moles reacted?" "How many moles of the other reactants would I need to use up all of these moles?" Next, do the calculation of how many moles of the "other" amount(s) you would need. Do you have enough? If so, the reactant you began with IS the limiting reactant. If not repeat this process with the "other" reactant amount you were given.

It doesn't matter where you start the "What if?" game….you get there either way.

Clear as mud? Read on…(and consider listening to the SCREENCAST!)

Let's revisit the Haber process:

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					
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	→	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.

Continued on next page.

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

In English, 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? 28,100 g produced = 28.1 kg
(It is always polite to respond in the unit given).

Which reactant is the limiting reactant? Hydrogen—once that's established, discard 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 18
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 19

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 6.82×10^4 g
Percent yield is 52.3%

AP* Chemistry
GASES

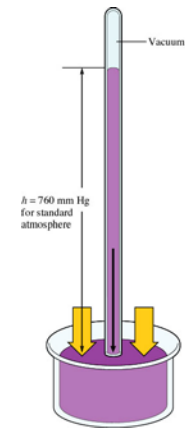
THE PROPERTIES OF GASES

Only 4 quantities are needed to define the **state of a gas**:

a) the *quantity* of the gas, n (in moles)
b) the *temperature* of the gas, T (**in KELVINS)**
c) the *volume* of the gas, V (in liters)
d) the *pressure* of the gas, P (in atmospheres)

A gas uniformly fills any container, is easily compressed & mixes completely with any other gas.

GAS PRESSURE:



Early barometer

A measure of the force that a gas exerts on its container. Force is the physical quantity that interferes with inertia. Gravity is the force responsible for weight.

Force = mass × acceleration; Newton's 2nd Law. The units of force follow:

$\text{N} = \text{kg} \times \text{m/s}^2$

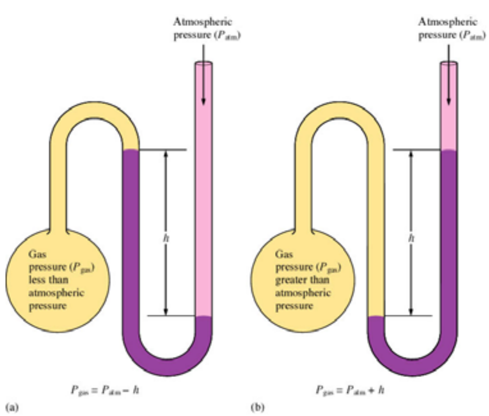
Pressure-- Force/ unit area; N/m^2

Barometer--invented by Evangelista Torricelli in 1643; uses the height of a column of mercury to measure gas pressure (especially atmospheric)

1 mm of Hg = 1 torr

760.00 mm Hg = 760.00 torr =1.00 atm = 101.325 kPa $\approx 10^5$ Pa

At sea level all of the above define STANDARD PRESSURE. The SI unit of pressure is the Pascal (Blaise Pascal); 1 Pa = 1 N / m^2



The **manometer**—a device for measuring the pressure of a gas in a container. The pressure of the gas is given by h [the difference in mercury levels] in units of torr (equivalent to mm Hg).

a) Gas pressure = atmospheric pressure – h

b) Gas pressure = atmospheric pressure + h

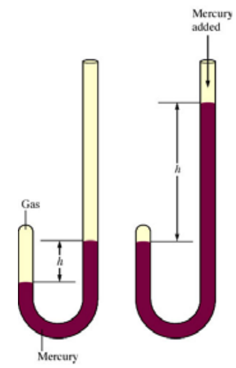
Exercise 1 Pressure Conversions

The pressure of a gas is measured as 49 torr. Represent this pressure in both atmospheres and pascals.

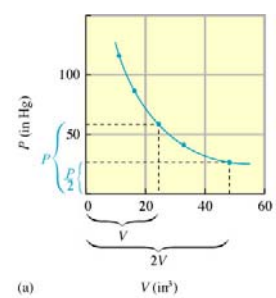
6.4×10^{-2} atm
6.5×10^{3} Pa

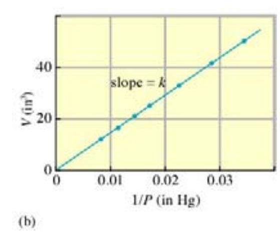
Exercise
Rank the following pressures in decreasing order of magnitude (largest first, smallest last): 75 kPa, 300. torr, 0.60 atm and 350. mm Hg.

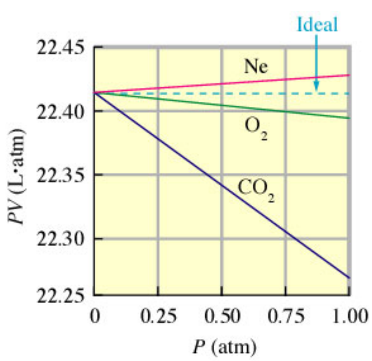
GAS LAWS: THE EXPERIMENTAL BASIS



- BOYLE'S LAW: "father of chemistry"--*the volume of a confined gas is inversely proportional to the pressure exerted on the gas.* ALL GASES BEHAVE IN THIS MANNER!
- Robert Boyle was an Irish chemist. He studied *PV* relationships using a J-tube set up in the multi-story entryway of his home.
  - o $P \propto 1/V$ plot = straight line
  - o pressure and volume are **inversely** proportional
  - o Volume ↑ pressure ↓ **at constant temperature,** the converse is also true
  - o *for a given quantity of a gas at constant temperature, the product of pressure and volume is a constant.*
    - ▪ $PV = k$
    - ▪ Therefore, $V = \frac{k}{P} = k\frac{1}{P}$
    - ▪ which is the equation for a straight line of the type
    - ▪ $y = mx + b$ where m = slope, and b the y-intercept
    - ▪ In this case, $y = V$, $x = 1/P$ and $b = 0$. Check out the plot on the right (b). The data Boyle collected is graphed on (a) above.
  - o $P_1V_1 = P_2V_2$ is the easiest form of Boyle's law to **memorize**
  - o Boyle's Law has been tested for over three centuries. It holds true ***<u>only at low pressures</u>***.







A plot of *PV* versus P for several gases at pressures below 1atm is pictured at left.

An **ideal** gas is expected to have a constant value of *PV*, as shown by the dotted line. CO_2 shows the largest change in PV, and this change is actually quite small: *PV* changes from about 22.39 L·atm at 0.25 atm to 22.26 L·atm at 1.00 atm. Thus Boyle's Law is a good approximation at these relatively low pressures.

Exercise 2 **Boyle's Law I**

Sulfur dioxide (SO_2), a gas that plays a central role in the formation of acid rain, is found in the exhaust of automobiles and power plants. Consider a 1.53- L sample of gaseous SO_2 at a pressure of 5.6×10^3 Pa. If the pressure is changed to 1.5×10^4 Pa at a constant temperature, what will be the new volume of the gas ?

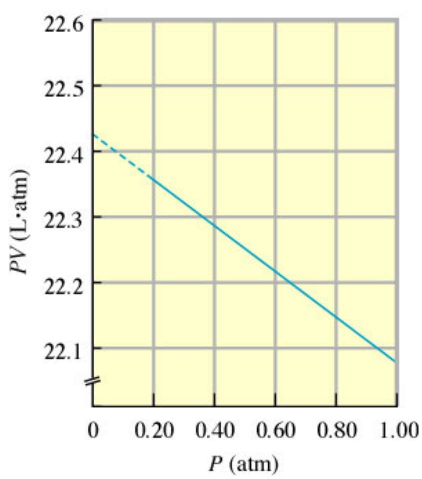
0.57 L

Exercise 3 **Boyle's Law II**

In a study to see how closely gaseous ammonia obeys Boyle's law, several volume measurements were made at various pressures, using 1.0 mol NH_3 gas at a temperature of 0ºC. Using the results listed below, calculate the Boyle's law constant for NH_3 at the various pressures.

Experiment	*Pressure (atm)*	*Volume (L)*
1	0.1300	172.1
2	0.2500	89.28
3	0.3000	74.35
4	0.5000	44.49
5	0.7500	29.55
6	1.000	22.08

experiment 1 is 22.37
experiment 2 is 22.32
experiment 3 is 22.31
experiment 4 is 22.25
experiment 5 is 22.16
experiment 6 is 22.08



PLOT the values of *PV* for the six experiments above.

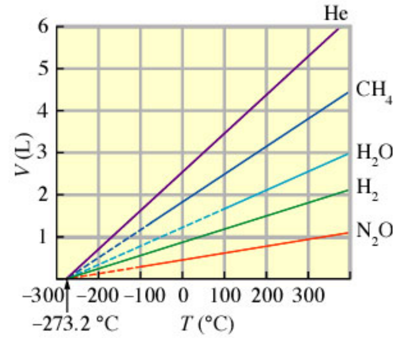
Extrapolate it back to see what PV equals at 0.00 atm pressure.

Compare it to the *PV* vs. *P* graph on page 2 of these notes.

What is the *y*-intercept for all of these gases?

Remember, gases behave most ideally at low pressures. You can't get a pressure lower than 0.00 atm!

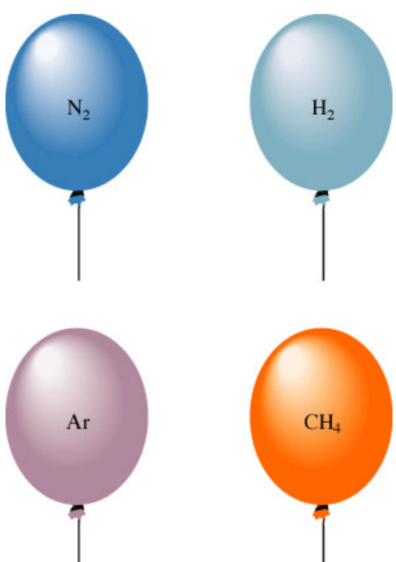
- **CHARLES' LAW**: *If a given quantity of gas is held at a constant pressure, then its volume is directly proportional to the absolute temperature.* **<u>Must use KELVIN</u>**
- Jacques Charles was a French physicist and the first person to fill a hot "air" balloon with hydrogen gas and made the first solo balloon flight!
    - o $V \propto T$ plot = straight line
    - o $V_1T_2 = V_2T_1$
    - o Temperature $\propto$ Volume **at constant pressure**
    - o This figure shows the plots of V vs. T (Celsius) for several gases. The solid lines represent experimental measurements on gases. The dashed lines represent extrapolation of the data into regions where these gases would become liquids or solids. Note that the samples of the various gases contain different numbers of moles.
    - o What is the temperature when the Volume extrapolates to zero?



Exercise 4 **Charles's Law**

A sample of gas at 15ºC and 1 atm has a volume of 2.58 L. What volume will this gas occupy at 38ºC and 1 atm ?

2.79 L



These balloons each hold 1.0 L of gas at 25°C and 1 atm. Each balloon contains 0.041 mol of gas, or 2.5×10^{22} molecules.

- **GAY-LUSSAC'S LAW of combining volumes:** volumes of gases always combine with one another in the ratio of small whole numbers, as long as volumes are measured at the same T and P.
    - $P_1T_2 = P_2T_1$
    - Avogadro=s hypothesis: *equal volumes of gases under the same conditions of temperature and pressure contain equal numbers of molecules.*
- **AVOGADRO'S LAW:** *The volume of a gas, at a given temperature and pressure, is directly proportional to the quantity of gas.*
    - $V \propto n$
    - $n \propto$ Volume **at constant T & P**

HERE'S AN EASY WAY TO MEMORIZE ALL OF THIS!
Start with the combined gas law:

$$\boldsymbol{P_1V_1T_2 = P_2V_2T_1} \qquad \text{Memorize it.}$$

Next, put the fellas' names in alphabetical order.
Boyle's uses the first 2 variables, Charles' the second 2 variables & Gay-Lussac's the remaining combination of variables. What ever doesn't appear in the formula, is being held CONSTANT!

Exercise 5 **Avogadro's Law**

Suppose we have a 12.2-L sample containing 0.50 mol oxygen gas (O_2) at a pressure of 1 atm and a temperature of 25ºC. If all this O_2 were converted to ozone (O_3) at the same temperature and pressure, what would be the volume of the ozone ?

8.1 L

THE IDEAL GAS LAW

Four quantities describe the state of a gas: pressure, volume, temperature, and # of moles (quantity). Combine all 3 laws:

$$V \propto \frac{nT}{P}$$

Replace the $\propto$ with a constant, R, and you get:

$$PV = nRT$$

The ideal gas law! ***It is an equation of state***.

R = 0.08206 L• atm/mol • K also expressed as 0.08206 L atm mol^{-1} K^{-1}

Useful only at low Pressures and high temperatures! Guaranteed points on the AP Exam!

These next exercises can all be solved with the ideal gas law, BUT, you can use another if you like!

Exercise 6 **Ideal Gas Law I**

A sample of hydrogen gas (H_2) has a volume of 8.56 L at a temperature of 0ºC and a pressure of 1.5 atm. Calculate the moles of H_2 molecules present in this gas sample.

0.57 mol

Exercise 7 **Ideal Gas Law II**

Suppose we have a sample of ammonia gas with a volume of 3.5 L at a pressure of 1.68 atm. The gas is compressed to a volume of 1.35 L at a constant temperature. Use the ideal gas law to calculate the final pressure.

4.4 atm

Exercise 8 **Ideal Gas Law III**

A sample of methane gas that has a volume of 3.8 L at 5ºC is heated to 86ºC at constant pressure. Calculate its new volume.

4.9 L

Exercise 9 **Ideal Gas Law IV**

A sample of diborane gas (B_2H_6), a substance that bursts into flame when exposed to air, has a pressure of 345 torr at a temperature of -15ºC and a volume of 3.48 L. If conditions are changed so that the temperature is 36ºC and the pressure is 468 torr, what will be the volume of the sample?

3.07 L

Exercise 10 **Ideal Gas Law V**

A sample containing 0.35 mol argon gas at a temperature of 13ºC and a pressure of 568 torr is heated to 56ºC and a pressure of 897 torr. Calculate the change in volume that occurs.

decreases by 3 L

GAS STOICHIOMETRY

Use $PV = nRT$ to solve for the volume of one mole of gas at STP:

Look familiar? This is the **molar volume** of a gas at STP. Work stoichiometry problems using your favorite method, dimensional analysis, mole map, the table way…just work FAST! Use the ideal gas law to convert quantities that are NOT at STP.

Exercise 11 **Gas Stoichiometry I**

A sample of nitrogen gas has a volume of 1.75 L at STP. How many moles of N_2 are present?

7.81×10^{-2} mol N_2

Exercise 12 **Gas Stoichiometry II**

Quicklime (CaO) is produced by the thermal decomposition of calcium carbonate ($CaCO_3$). Calculate the volume of CO_2 at STP produced from the decomposition of 152 g $CaCO_3$ by the reaction

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

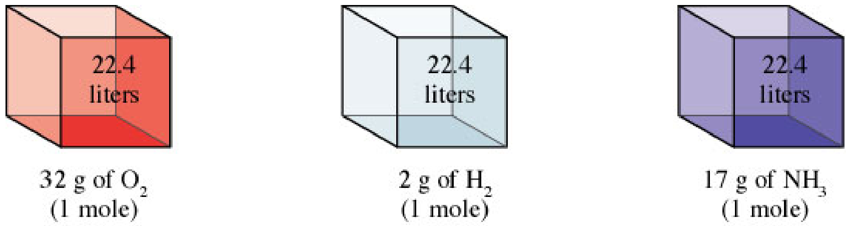
34.1 L CO_2 at STP

Exercise 13 **Gas Stoichiometry III**

A sample of methane gas having a volume of 2.80 L at 25ºC and 1.65 atm was mixed with a sample of oxygen gas having a volume of 35.0 L at 31ºC and 1.25 atm. The mixture was then ignited to form carbon dioxide and water. Calculate the volume of CO_2 formed at a pressure of 2.50 atm and a temperature of 125ºC.

2.47 L

THE DENSITY OF GASES:



$$d = \frac{m}{V} = \frac{P(MM)}{RT} \;\; \text{\{for ONE mole of gas\}} = \frac{MM}{22.4\ L} \quad \text{AND} \quad \text{Molar Mass} = MM = \frac{dRT}{P}$$

"Molecular Mass kitty cat"—all good cats put dirt [*dRT*] over their pee [*P*]. Corny? Yep! But, you'll thank me later!

Just remember that densities of gases are reported in g/L NOT g/mL.

What is the approximate molar mass of air? _________

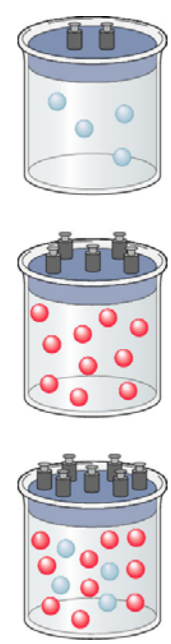
The density of air is approx. _______ g/L.
List 3 gases that float in air:
List 3 gases that sink in air:

Exercise 14 **Gas Density/Molar Mass**

The density of a gas was measured at 1.50 atm and 27ºC and found to be 1.95 g/L. Calculate the molar mass of the gas.

32.0 g/mol

GAS MIXTURES AND PARTIAL PRESSURES



The pressure of a mixture of gases is the sum of the pressures of the different components of the mixture:

$$P_{total} = P_1 + P_2 + \ldots P_n$$

John Dalton's Law of Partial Pressures also uses the concept of mole fraction, χ

$$\chi_A = \frac{\text{moles of A}}{\text{moles A + moles B + moles C + . . .}}$$

so now, $P_A = \chi_A P_{total}$

The partial pressure of each gas in a mixture of gases in a container depends on the number of moles of that gas. The total pressure is the SUM of the partial pressures and depends on the total moles of gas particles present, no matter what they are!

Exercise 15 **Dalton's Law I**

Mixtures of helium and oxygen are used in scuba diving tanks to help prevent "the bends." For a particular dive, 46 L He at 25ºC and 1.0 atm and 12 L O_2 at 25ºC and 1.0 atm were pumped into a tank with a volume of 5.0 L. Calculate the partial pressure of each gas and the total pressure in the tank at 25ºC.

P_{He} = 9.3 atm
P_{O2} = 2.4 atm
P_{TOTAL} = 11.7 atm

Exercise 16 **Dalton's Law II**

The partial pressure of oxygen was observed to be 156 torr in air with a total atmospheric pressure of 743 torr. Calculate the mole fraction of O_2 present.

0.210

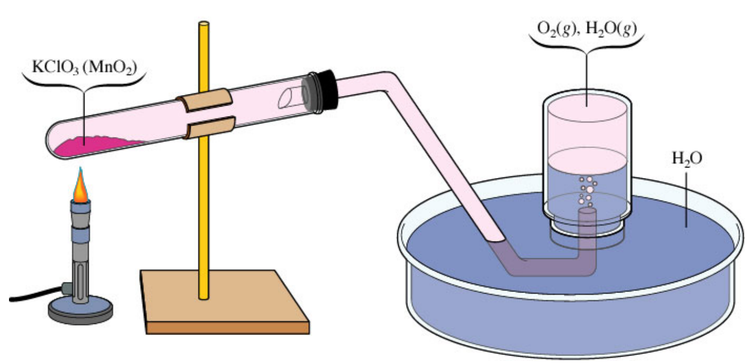
Exercise 17 **Dalton's Law III**

The mole fraction of nitrogen in the air is 0.7808. Calculate the partial pressure of N_2 in air when the atmospheric pressure is 760. torr.

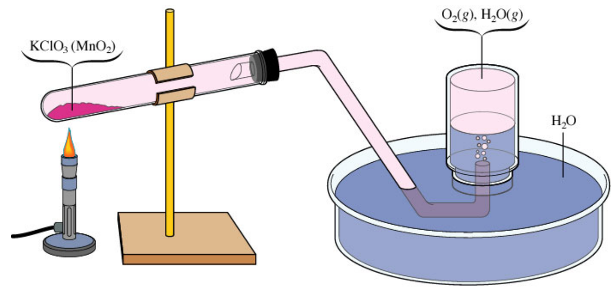
593 torr

WATER DISPLACEMENT

It is common to collect a gas by water displacement which means some of the pressure is due to water vapor collected as the gas was passing through! You must correct for this. You look up the partial pressure due to water vapor by knowing the ***temperature***.



Exercise 8 Gas Collection over Water



A sample of solid potassium chlorate ($KClO_3$) was heated in a test tube (see the figure above) and decomposed by the following reaction:

$$2\ KClO_3(s) \rightarrow 2\ KCl(s) + 3\ O_2(g)$$

The oxygen produced was collected by displacement of water at 22ºC at a total pressure of 754 torr. The volume of the gas collected was 0.650 L, and the vapor pressure of water at 22ºC is 21 torr. Calculate the partial pressure of O_2 in the gas collected and the mass of $KClO_3$ in the sample that was decomposed.

Partial pressure of O_2 = 733 torr
2.12 g $KClO_3$

KINETIC MOLECULAR THEORY OF GASES

Assumptions of the MODEL:

1. All particles are in constant, random, motion.
2. All collisions between particles are perfectly elastic.
3. The volume of the particles in a gas is negligible
4. The average kinetic energy of the molecules is its Kelvin temperature.

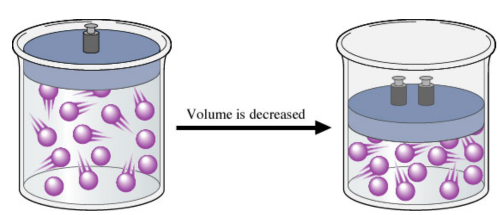
This neglects any intermolecular forces as well.
Gases expand to fill their container, solids/liquids do not.
Gases are compressible, solids/liquids are not appreciably compressible.

This helps explain Boyle's Law:

If the volume is decreased that means that the gas particles will hit the wall more often, thus increasing pressure

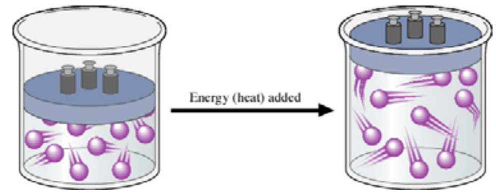
$$P = (nRT)\frac{1}{V}$$

Constant



And helps explain Charles' Law

When a gas is heated, the speed of its particles increase and thus hit the walls more often and with more force. The only way to keep the *P* constant is to increase the volume of the container.

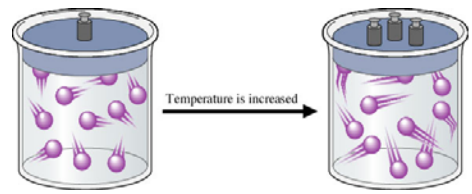


$$V = \left(\frac{nR}{P}\right)T$$

Constant

And also helps explain Gay-Lussac's Law

When the temperature of a gas increases, the speeds of its particles increase, the particles are hitting the wall with greater force and greater frequency.   Since the volume remains the same this would result in increased gas pressure.

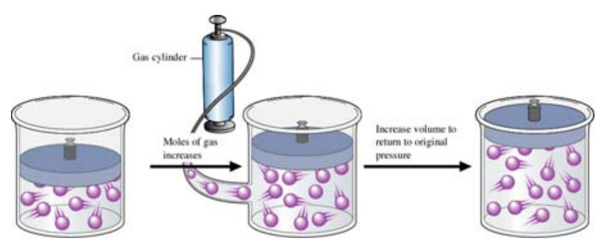


$$P = \left(\frac{nR}{V}\right)T$$

Constant

And even helps explain Avogadro's Law

An increase in the number of particles at the same temperature would cause the pressure to increase if the volume were held constant. The only way to keep constant *P* is to vary the *V*.



$$V = \left(\frac{RT}{P}\right)n$$

Constant

What about Dalton's Law? The *P* exerted by a *mixture* of gases is the SUM of the partial pressures since gas particles are acting *independent* of each other and the volumes of the individual particles DO NOT matter.

DISTRIBUTION OF MOLECULAR SPEEDS

Plot # of gas molecules having various speeds vs. the speed and you get a curve. Changing the temperature affects the *shape* of the curve NOT the area beneath it. Change the # of molecules and all bets are off!

Maxwell's equation:

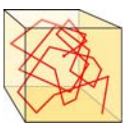
$$\sqrt{\overline{u^2}} = u_{rms} = \sqrt{\frac{3RT}{MM}}$$

Use the "energy R" or 8.314510 J/K• mol for this equation since kinetic energy is involved.

Exercise 19 Root Mean Square Velocity

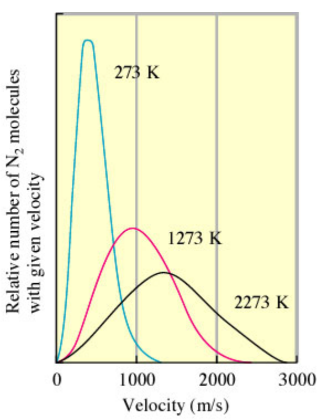
Calculate the root mean square velocity for the atoms in a sample of helium gas at 25ºC.

1.36×10^3 m/s



If we could monitor the path of a single molecule it would be very erratic.

Mean free path—the average distance a particle travels between collisions. It's on the order of a tenth of a micrometer. WAAAAY SMALL!



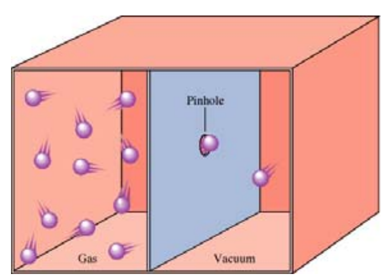
Examine the effect of temperature on the numbers of molecules with a given velocity as it relates to temperature.
HEAT 'EM UP, SPEED 'EM UP!!

Drop a vertical line from the peak of each of the three bell shaped curves—that point on the *x*-axis represents the AVERAGE velocity of the sample at that temperature. Note how the bells are "squashed" as the temperature increases. You may see graphs like this on the AP exam where you have to identify the highest temperature based on the shape of the graph!

GRAHAM'S LAW OF DIFFUSION AND EFFUSION

Effusion is closely related to diffusion. **Diffusion** is the term used to describe the mixing of gases. The ***rate*** of diffusion is the ***rate*** of the mixing.

Effusion is the term used to describe the passage of a gas through a tiny orifice into an evacuated chamber as shown on the right. The rate of effusion measures the speed at which the gas is transferred into the chamber.



The rates of effusion of two gases are inversely proportional to the square roots of their molar masses at the same temperature and pressure.

$$\frac{\text{Rate of effusion of gas 1}}{\text{Rate of effusion of gas 2}} = \sqrt{\frac{MM_2}{MM_1}}$$

REMEMBER ***rate*** is a change in a quantity over time, NOT just the time!

Exercise 20 **Effusion Rates**

Calculate the ratio of the effusion rates of hydrogen gas (H_2) and uranium hexafluoride (UF_6), a gas used in the enrichment process to produce fuel for nuclear reactors.

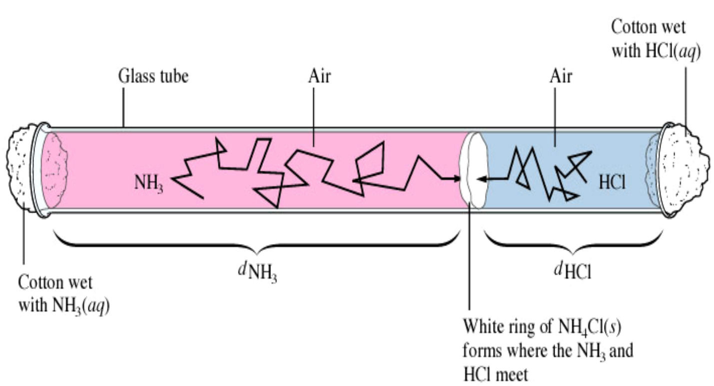
13.2

Exercise

A pure sample of methane is found to effuse through a porous barrier in 1.50 minutes. Under the same conditions, an equal number of molecules of an unknown gas effuses through the barrier in 4.73 minutes. What is the molar mass of the unknown gas?

Diffusion

This is a classic!



$$\frac{\text{Distance traveled by } NH_3}{\text{Distance traveled by HCl}} = \frac{u_{rms} \text{ for } NH_3}{u_{rms} \text{ for HCl}} = \sqrt{\frac{MM_{HCl}}{MM_{NH_3}}} = \sqrt{\frac{36.5}{17}} = 1.5$$

The observed ratio is LESS than a 1.5 distance ratio—why?

This diffusion is slow considering the molecular velocities are 450 and 660 meters per second—which one is which?

This tube contains air and all those collisions slow the process down in the real world. Speaking of real world….

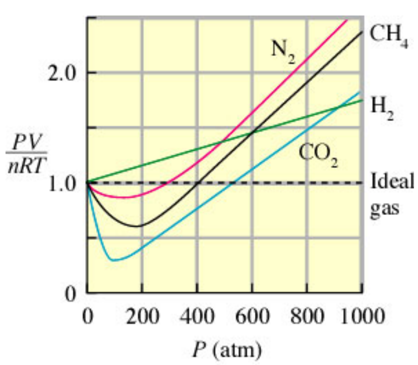
REAL, thus *NON*IDEAL GASES
Most gases behave ideally until you reach high pressure and low temperature. (By the way, either of these can cause a gas to liquefy, go figure!)

van der Waals Equation--corrects for negligible volume of molecules and accounts for inelastic collisions leading to intermolecular forces (his real claim to fame).

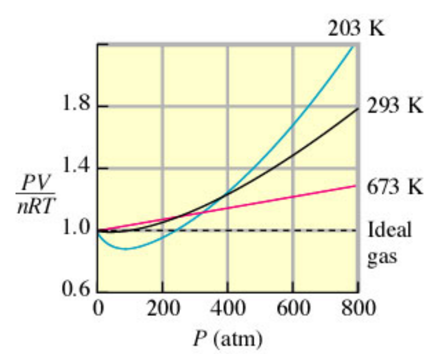
$$[P + a(\frac{n}{V})^2][V - bn] = nRT$$

a and *b* are van der Waals constants; no need to work problems, it's the concepts that are important! Notice pressure is increased (intermolecular forces lower real pressure, you're correcting for this) and volume is decreased (corrects the container to a smaller "free" volume).

These graphs are classics and make great multiple choice questions on the AP exam.



When PV/nRT = 1.0, the gas is ideal
All of these are at 200K.
Note the P's where the curves cross the dashed line [ideality].



This graph is just for nitrogen gas.
Note that although nonideal behavior is evident at each temperature, the deviations are smaller at the higher Ts.

Don't underestimate the power of understanding these graphs. We love to ask question comparing the behavior of ideal and real gases. It's not likely you'll be asked an entire free-response gas problem on the real exam in May. Gas Laws are tested extensively in the multiple choice since it's easy to write questions involving them! You will most likely see $PV = nRT$ as one part of a problem in the free-response, just not a whole problem!

GO FORTH AND RACK UP THOSE MULTIPLE CHOICE POINTS!!

1982

Water is added to 4.267 grams of UF_6. The only products are 3.730 grams of a solid containing only uranium, oxygen and fluorine and 0.970 gram of a gas. The gas is 95.0% fluorine, and the remainder is hydrogen.

(a) From these data, determine the empirical formula of the gas.
(b) What fraction of the fluorine of the original compound is in the solid and what fraction in the gas after the reaction?
(c) What is the formula of the solid product?
(d) Write a balanced equation for the reaction between UF_6 and H_2O. Assume that the empirical formula of the gas is the true formula.

1986

Three volatile compounds X, Y, and Z each contain element Q. The percent by weight of element Q in each compound was determined. Some of the data obtained are given below.

Compound	Percent by weight of Element Q	Molecular Weight
X	64.8%	?
Y	73.0%	104.
Z	59.3%	64.0

(a) The vapor density of compound X at 27°C and 750. mm Hg was determined to be 3.53 grams per liter. Calculate the molecular weight of compound X.
(b) Determine the mass of element Q contained in 1.00 mole of each of the three compounds.
(c) Calculate the most probable value of the atomic weight of element Q.
(d) Compound Z contains carbon, hydrogen, and element Q. When 1.00 gram of compound Z is oxidized and all of the carbon and hydrogen are converted to oxides, 1.37 grams of CO_2 and 0.281 gram of water are produced. Determine the most probable molecular formula of compound Z.

1991

The molecular formula of a hydrocarbon is to be determined by analyzing its combustion products and investigating its colligative properties.

(a) The hydrocarbon burns completely, producing 7.2 grams of water and 7.2 liters of CO_2 at standard conditions. What is the empirical formula of the hydrocarbon?
(b) Calculate the mass in grams of O_2 required for the complete combustion of the sample of the hydrocarbon described in (a).
(c) The hydrocarbon dissolves readily in $CHCl_3$. The freezing point of a solution prepared by mixing 100. grams of $CHCl_3$ and 0.600 gram of the hydrocarbon is −64.0°C. The molal freezing-point depression constant of $CHCl_3$ is 4.68°C/molal and its normal freezing point is −63.5°C. Calculate the molecular weight of the hydrocarbon.
(d) What is the molecular formula of the hydrocarbon?

1993

I. $2\ Mn^{2+} + 4\ OH^- + O_2(g) \rightarrow 2\ MnO_2(s) + 2\ H_2O$
II. $MnO_2(s) + 2\ I^- + 4\ H^+ \rightarrow Mn^{2+} + I_2(aq) + 2\ H_2O$
III. $2\ S_2O_3^{2-} + I_2(aq) \rightarrow S_4O_6^{2-} + 2\ I^-$

The amount of oxygen, O_2, dissolved in water can be determined by titration. First, $MnSO_4$ and NaOH are added to a sample of water to convert all of the dissolved O_2 to MnO_2, as shown in equation I above. Then H_2SO_4 and KI are added and the reaction represented by equation II proceeds. Finally, the I_2 that is formed is titrated with standard sodium thiosulfate, $Na_2S_2O_3$, according to equation III.

(a) According to the equation above, how many moles of $S_2O_3^{2-}$ are required for analyzing 1.00 mole of O_2 dissolved in water?
(b) A student found that a 50.0-milliliter sample of water required 4.86 milliliters of 0.0112-molar $Na_2S_2O_3$ to reach the equivalence point. Calculate the number of moles of O_2 dissolved in this sample.
(c) How would the results in (b) be affected if some I_2 were lost before the $S_2O_3^{2-}$ was added? Explain.
(d) What volume of dry O_2 measured at 25°C and 1.00 atmosphere of pressure would have to be dissolved in 1.00 liter of pure water in order to prepare a solution of the same concentration as that obtained in (b)? (cont.)
(e) Name an appropriate indicator for the reaction shown in equation III and describe the change you would observe at the end point of the titration.

1998

An unknown compound contains only the three elements C,H, and O. A pure sample of the compound is analyzed and found to be 65.60 percent C and 9.44 percent H by mass.

(a) Determine the empirical formula of the compound.

(b) A solution of 1.570 grams of the compound in 16.08 grams of camphor is observed to freeze at a temperature 15.2 Celsius degrees below the normal freezing point of pure camphor. Determine the molar mass and apparent molecular formula of the compound. (The molal freezing-point depression constant, K_f, for camphor is 40.0 kg·K·mol^{-1}.)

(c) When 1.570 grams of the compound is vaporized at 300 °C and 1.00 atmosphere, the gas occupies a volume of 577 milliliters. What is the molar mass of the compound based on this result?

(d) Briefly describe what occurs in solution that accounts for the difference between the results obtained in parts (b) and (c).

2000

Answer the following questions about $BeC_2O_4(s)$ and its hydrate.

(a) Calculate the mass percent of carbon in the hydrated form of the solid that has the formula $BeC_2O_4 \bullet 3H_2O$

(b) When heated to 220.°C, $BeC_2O_4 \bullet 3\ H_2O(s)$ dehydrates completely as represented below.

$$BeC_2O_4 \bullet 3\ H_2O(s) \rightarrow BeC_2O_4(s) + 3\ H_2O(g)$$

If 3.21 g of $BeC_2O_4 \bullet 3\ H_2O(s)$ is heated to 220.°C, calculate

(i) the mass of $BeC_2O_4(s)$ formed, and,

(ii) the volume of the $H_2O(g)$ released, measured at 220.°C and 735 mm Hg.

(c) A 0.345 g sample of anhydrous BeC_2O_4, which contains an inert impurity, was dissolved in sufficient water to produce 100. mL of solution. A 20.0 mL portion of the solution was titrated with $KMnO_4(aq)$. The balanced equation for the reaction that occurred is as follows.

$$16\ H^+(aq) + 2\ MnO_4^-(aq) + 5\ C_2O_4^{2-}(aq) \rightarrow 2\ Mn^{2+}(aq) + 10\ CO_2(g) + 8\ H_2O(l).$$

The volume of 0.0150 *M* $KMnO_4(aq)$ required to reach the equivalence point was 17.80 mL.

(i) Identify the reducing agent in the titration reaction.

(ii) For the titration at the equivalence point, calculate the number of moles of each of the following that reacted.

- $MnO_4^-(aq)$
- $C_2O_4^{2-}(aq)$

(iii) Calculate the total number of moles of $C_2O_4^{2-}(aq)$ that were present in the 100. mL of prepared solution.

(iv) Calculate the mass percent of $BeC_2O_4(s)$ in the impure 0.345 g sample.

2003B

Answer the following questions that relate to chemical reactions.

(a) Iron(III) oxide can be reduced with carbon monoxide according to the following equation.

$$Fe_2O_3(s) + 3\ CO(g) \rightarrow 2\ Fe(s) + 3\ CO_2(g)$$

A 16.2 L sample of $CO(g)$ at 1.50 atm and 200.°C is combined with 15.39 g of $Fe_2O_3(s)$.

(i) How many moles of CO(g) are available for the reaction?

(ii) What is the limiting reactant for the reaction? Justify your answer with calculations.

(iii) How many moles of Fe(s) are formed in the reaction?

(b) In a reaction vessel, 0.600 mol of $Ba(NO_3)_2(s)$ and 0.300 mol of $H_3PO_4(aq)$ are combined with deionized water to a final volume of 2.00 L. The reaction represented below occurs.

$$3\ Ba(NO_3)_2(aq) + 2\ H_3PO_4(aq) \rightarrow Ba_3(PO_4)_2(s) + 6\ HNO_3(aq)$$

(i) Calculate the mass of $Ba_3(PO_4)_2(s)$ formed.

(ii) Calculate the pH of the resulting solution.

(iii) What is the concentration, in mol L^{-1}, of the nitrate ion, $NO_3^-(aq)$, after the reaction reaches completion?

1971

$$2\ HCOONa + H_2SO_4 \rightarrow 2\ CO + 2\ H_2O + Na_2SO_4$$

A 0.964 gram sample of a mixture of sodium formate and sodium chloride is analyzed by adding sulfuric acid. The equation for the reaction for sodium formate with sulfuric acid is shown above. The carbon monoxide formed measures 242 milliliters when collected over water at 752 torr and 22.0°C. Calculate the percentage of sodium formate in the original mixture.

1971

At 20°C the vapor pressure of benzene is 75 torr, and the vapor pressure of toluene is 22 torr. Solutions in both parts of this question are to be considered ideal.

(a) A solution is prepared from 1.0 mole of biphenyl, a nonvolatile solute, and 49.0 moles of benzene. Calculate the vapor pressure of the solution at 20°C.

(b) A second solution is prepared from 3.0 moles of toluene and 1.0 mole of benzene. Determine the vapor pressure of this solution and the mole fraction of benzene in the vapor.

1972

A 5.00 gram sample of a dry mixture of potassium hydroxide, potassium carbonate, and potassium chloride is reacted with 0.100 liter of 2.0 molar HCl solution.

(a) A 249 milliliter sample of dry CO_2 gas, measured at 22°C and 740 torr, is obtained from the reaction. What is the percentage of potassium carbonate in the mixture?

(b) The excess HCl is found by titration to be chemically equivalent to 86.6 milliliters of 1.50 molar NaOH. Calculate the percentages of potassium hydroxide and of potassium chloride in the original mixture.

1973

A 6.19 gram sample of PCl_5 is placed in an evacuated 2.00 liter flask and is completely vaporized at 252°C.

(a) Calculate the pressure in the flask if no chemical reaction were to occur.

(b) Actually at 252°C the PCl_5 is partially dissociated according to the following equation:

$$PCl_5(g) \rightarrow PCl_3(g) + Cl_2(g)$$

The observed pressure is found to be 1.00 atmosphere. In view of this observation, calculate the partial pressure of PCl_5 and PCl_3 in the flask at 252°C.

1976

When the molecular weight of a volatile liquid is calculated from the weight, volume, temperature, and pressure of a sample of that liquid when vaporized, the assumption is usually made that the gas behaves ideally. In fact at a temperature not far above the boiling point of the liquid, the gas is not ideal. Explain how this would affect the results of the molecular weight determination.

1982

(a) From the standpoint of the kinetic-molecular theory, discuss briefly the properties of gas molecules that cause deviations from ideal behavior.

(b) At 25°C and 1 atmosphere pressure, which of the following gases shows the greatest deviation from ideal behavior? Give two reasons for your choice.

CH_4 SO_2 O_2 H_2

(c) Real gases approach ideality at low pressure, high temperature, or both. Explain these observations.

1984

The van der Waals equation of state for one mole of a real gas is as follows:

$$(P + {}^{a}/_{V^2})(V - b) = RT$$

For any given gas, the values of the constants *a* and *b* can be determined experimentally. Indicate which physical properties of a molecule determine the magnitudes of the constants a and b. Which of the two molecules, H_2 or H_2S, has the higher value for *a* and which has the higher value for *b*? Explain.

One of the van der Waals constants can be correlated with the boiling point of a substance. Specify which constant and how it is related to the boiling point.

1986

Three volatile compounds X, Y, and Z each contain element Q. The percent by weight of element Q in each compound was determined. Some of the data obtained are given below.

Compound	Percent by Weight of Element Q	Molecular Weight
X	64.8%	?
Y	73.0%	104.
Z	59.3%	64.0

(a) The vapor density of compound X at 27 degrees Celsius and 750. mm Hg was determined to be 3.53 grams per liter. Calculate the molecular weight of compound X.

(b) Determine the mass of element Q contained in 1.00 mole of each of the three compounds.

(c) Calculate the most probable value of the atomic weight of element Q.

(d) Compound Z contains carbon, hydrogen, and element Q. When 1.00 gram of compound Z is oxidized and all of the carbon and hydrogen are converted to oxides, 1.37 grams of CO_2 and 0.281 gram of water are produced. Determine the most probable molecular formula.

1990

A mixture of $H_2(g)$, $O_2(g)$, and 2 millilitres of $H_2O(l)$ is present in a 0.500 litre rigid container at 25°C. The number of moles of H_2 and the number of moles of O_2 are equal. The total pressure is 1,146 millimetres mercury. (The equilibrium vapor pressure of pure water at 25°C is 24 millimetres mercury.)

The mixture is sparked, and H_2 and O_2 react until one reactant is completely consumed.

(a) Identify the reactant remaining and calculate the number of moles of the reactant remaining.

(b) Calculate the total pressure in the container at the conclusion of the reaction if the final temperature is 90°C. (The equilibrium vapor pressure of water at 90°C is 526 millimetres mercury.)

(c) Calculate the number of moles of water present <u>as vapor</u> in the container at 90°C.

1993

Observations about real gases can be explained at the molecular level according to the kinetic molecular theory of gases and ideas about intermolecular forces. Explain how each of the following observations can be interpreted according to these concepts, including how the observation supports the correctness of these theories.

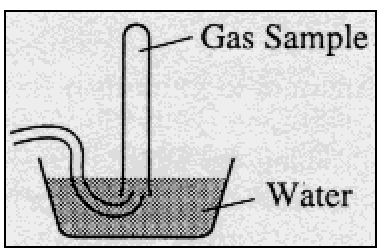
(a) When a gas-filled balloon is cooled, it shrinks in volume; this occurs no matter what gas is originally placed in the balloon.

(b) When the balloon described in (a) is cooled further, the volume does not become zero; rather, the gas becomes a liquid or solid.

(c) When NH_3 gas is introduced at one end of a long tube while HCl gas is introduced simultaneously at the other end, a ring of white ammonium chloride is observed to form in the tube after a few minutes. This ring is closer to the HCl end of the tube than the NH_3 end.

(d) A flag waves in the wind.

1994



A student collected a sample of hydrogen gas by the displacement of water as shown by the diagram above. The relevant data are given in the following table.

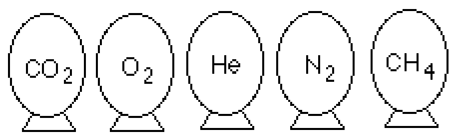
GAS SAMPLE DATA	
Volume of sample	90.0 mL
Temperature	25°C
Atmospheric Pressure	745 mm Hg
Equilibrium Vapor Pressure of H_2O (25°C)	23.8 mm Hg

(a) Calculate the number of moles of hydrogen gas collected.
(b) Calculate the number of molecules of water vapor in the sample of gas.
(c) Calculate the ratio of the average speed of the hydrogen molecules to the average speed of the water vapor molecules in the sample.
(d) Which of the two gases, H_2 or H_2O, deviates more from ideal behavior? Explain your answer.

1995

Propane, C_3H_8, is a hydrocarbon that is commonly used as fuel for cooking.
(a) Write a balanced equation for the complete combustion of propane gas, which yields $CO_2(g)$ and $H_2O(l)$.
(b) Calculate the volume of air at 30°C and 1.00 atmosphere that is needed to burn completely 10.0 grams of propane. Assume that air is 21.0 percent O_2 by volume.
(c) The heat of combustion of propane is −2,220.1 kJ/mol. Calculate the heat of formation, ΔH_f°, of propane given that ΔH_f° of $H_2O(l)$ = −285.3 kJ/mol and ΔH_f° of $CO_2(g)$ = −393.5 kJ/mol.
(d) Assuming that all of the heat evolved in burning 30.0 grams of propane is transferred to 8.00 kilograms of water (specific heat = 4.18 J/g·K), calculate the increase in temperature of water.

1996



Represented above are five identical balloons, each filled to the same volume at 25°C and 1.0 atmosphere pressure with the pure gas indicated.

(a) Which balloon contains the greatest mass of gas? Explain.

(b) Compare the average kinetic energies of the gas molecules in the balloons. Explain.

(c) Which balloon contains the gas that would be expected to deviate most from the behavior of an ideal gas? Explain.

(d) Twelve hours after being filled, all the balloons have decreased in size. Predict which balloon will be the smallest. Explain your reasoning.

2002B

A rigid 8.20 L flask contains a mixture of 2.50 moles of H_2, 0.500 mole of O_2, and sufficient Ar so that the partial pressure of Ar in the flask is 2.00 atm. The temperature is 127ºC.

(a) Calculate the total pressure in the flask.
(b) Calculate the mole fraction of H_2 in the flask.
(c) Calculate the density (in g L^{-1}) of the mixture in the flask.

The mixture in the flask is ignited by a spark, and the reaction represented below occurs until one of the reactants is entirely consumed.

$$2\,H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$

(d) Give the mole fraction of all species present in the flask at the end of the reaction.

Student Safety Contract

School Name ______________________________ **Teacher** ____________________

PURPOSE

Science is a hands-on laboratory class. You will be doing many laboratory activities which require the use of hazardous chemicals. Safety in the science classroom is the #1 priority for students, teachers, and parents. To ensure a safe science classroom, a list of rules has been developed and provided to you in this student safety contract. These rules must be followed at all times. Two copies of the contract are provided. One copy must be signed by both you and a parent or guardian before you can participate in the laboratory. The second copy is to be kept in your science notebook as a constant reminder of the safety rules.

GENERAL RULES

1. Conduct yourself in a responsible manner at all times in the laboratory.
2. Follow all written and verbal instructions carefully. If you do not understand a direction or part of a procedure, ask the instructor before proceeding.
3. Never work alone. No student may work in the laboratory without an instructor present.
4. When first entering a science room, do not touch any equipment, chemicals, or other materials in the laboratory area until you are instructed to do so.
5. Do not eat food, drink beverages, or chew gum in the laboratory. Do not use laboratory glassware as containers for food or beverages.
6. Perform only those experiments authorized by the instructor. Never do anything in the laboratory that is not called for in the laboratory procedures or by your instructor. Carefully follow all instructions, both written and oral. Unauthorized experiments are prohibited.
7. Be prepared for your work in the laboratory. Read all procedures thoroughly before entering the laboratory.
8. Never fool around in the laboratory. Horseplay, practical jokes, and pranks are dangerous and prohibited.
9. Observe good housekeeping practices. Work areas should be kept clean and tidy at all times. Bring only your laboratory instructions, worksheets, and/or reports to the work area. Other materials (books, purses, backpacks, etc.) should be stored in the classroom area.
10. Keep aisles clear. Push your chair under the desk when not in use.
11. Know the locations and operating procedures of all safety equipment including the first aid kit, eyewash station, safety shower, fire extinguisher, and fire blanket. Know where the fire alarm and the exits are located.
12. Always work in a well-ventilated area. Use the fume hood when working with volatile substances or poisonous vapors. Never place your head into the fume hood.
13. Be alert and proceed with caution at all times in the laboratory. Notify the instructor immediately of any unsafe conditions you observe.
14. Dispose of all chemical waste properly. Never mix chemicals in sink drains. Sinks are to be used only for water and those solutions designated by the instructor. Solid chemicals, metals, matches, filter paper, and all other insoluble materials are to be disposed of in the proper waste containers, not in the sink. Check the label of all waste containers twice before adding your chemical waste to the container.
15. Labels and equipment instructions must be read carefully before use. Set up and use the prescribed apparatus as directed in the laboratory instructions or by your instructor.
16. Keep hands away from face, eyes, mouth and body while using chemicals or preserved specimens. Wash your hands with soap and water after performing all experiments. Clean all work surfaces and apparatus at the end of the experiment. Return all equipment clean and in working order to the proper storage area.
17. Experiments must be personally monitored at all times. You will be assigned a laboratory station at which to work. Do not wander around the room, distract other students, or interfere with the laboratory experiments of others.
18. Students are never permitted in the science storage rooms or preparation areas unless given specific permission by their instructor.
19. Know what to do if there is a fire drill during a laboratory period; containers must be closed, gas valves turned off, fume hoods turned off, and any electrical equipment turned off.
20. Handle all living organisms used in a laboratory activity in a humane manner. Preserved biological materials are to be treated with respect and disposed of properly.
21. When using knives and other sharp instruments, always carry with tips and points pointing down and away. Always cut away from your body. Never try to catch falling sharp instruments. Grasp sharp instruments only by the handles.
22. If you have a medical condition (e.g., allergies, pregnancy, etc.), check with your physician prior to working in lab.

CLOTHING

23. Any time chemicals, heat, or glassware are used, students will wear laboratory goggles. There will be no exceptions to this rule!
24. Contact lenses should not be worn in the laboratory unless you have permission from your instructor.
25. Dress properly during a laboratory activity. Long hair, dangling jewelry, and loose or baggy clothing are a hazard in the laboratory. Long hair must be tied back and dangling jewelry and loose or baggy clothing must be secured. Shoes must completely cover the foot. No sandals allowed.
26. Lab aprons have been provided for your use and should be worn during laboratory activities.

ACCIDENTS AND INJURIES

27. Report any accident (spill, breakage, etc.) or injury (cut, burn, etc.) to the instructor immediately, no matter how trivial it may appear.
28. If you or your lab partner are hurt, immediately yell out "Code one, Code one" to get the instructor's attention.
29. If a chemical splashes in your eye(s) or on your skin, immediately flush with running water from the eyewash station or safety shower for at least 20 minutes. Notify the instructor immediately.
30. When mercury thermometers are broken, mercury must not be touched. Notify the instructor immediately.

HANDLING CHEMICALS

31. All chemicals in the laboratory are to be considered dangerous. Do not touch, taste, or smell any chemicals unless specifically instructed to do so. The proper technique for smelling chemical fumes will be demonstrated to you.
32. Check the label on chemical bottles twice before removing any of the contents. Take only as much chemical as you need.
33. Never return unused chemicals to their original containers.

Student Safety Contract
Continued

34. Never use mouth suction to fill a pipet. Use a rubber bulb or pipet pump.
35. When transferring reagents from one container to another, hold the containers away from your body.
36. Acids must be handled with extreme care. You will be shown the proper method for diluting strong acids. Always add acid to water, swirl or stir the solution and be careful of the heat produced, particularly with sulfuric acid.
37. Handle flammable hazardous liquids over a pan to contain spills. Never dispense flammable liquids anywhere near an open flame or source of heat.
38. Never remove chemicals or other materials from the laboratory area.
39. Take great care when transporting acids and other chemicals from one part of the laboratory to another. Hold them securely and walk carefully.

HANDLING GLASSWARE AND EQUIPMENT

40. Carry glass tubing, especially long pieces, in a vertical position to minimize the likelihood of breakage and injury.
41. Never handle broken glass with your bare hands. Use a brush and dustpan to clean up broken glass. Place broken or waste glassware in the designated glass disposal container.
42. Inserting and removing glass tubing from rubber stoppers can be dangerous. Always lubricate glassware (tubing, thistle tubes, thermometers, etc.) before attempting to insert it in a stopper. Always protect your hands with towels or cotton gloves when inserting glass tubing into, or removing it from, a rubber stopper. If a piece of glassware becomes "frozen" in a stopper, take it to your instructor for removal.
43. Fill wash bottles only with distilled water and use only as intended, e.g., rinsing glassware and equipment, or adding water to a container.
44. When removing an electrical plug from its socket, grasp the plug, not the electrical cord. Hands must be completely dry before touching an electrical switch, plug, or outlet.
45. Examine glassware before each use. Never use chipped or cracked glassware. Never use dirty glassware.
46. Report damaged electrical equipment immediately. Look for things such as frayed cords, exposed wires, and loose connections. Do not use damaged electrical equipment.
47. If you do not understand how to use a piece of equipment, ask the instructor for help.
48. Do not immerse hot glassware in cold water; it may shatter.

HEATING SUBSTANCES

49. Exercise extreme caution when using a gas burner. Take care that hair, clothing and hands are a safe distance from the flame at all times. Do not put any substance into the flame unless specifically instructed to do so. Never reach over an exposed flame. Light gas (or alcohol) burners only as instructed by the teacher.
50. Never leave a lit burner unattended. Never leave anything that is being heated or is visibly reacting unattended. Always turn the burner or hot plate off when not in use.
51. You will be instructed in the proper method of heating and boiling liquids in test tubes. Do not point the open end of a test tube being heated at yourself or anyone else.
52. Heated metals and glass remain very hot for a long time. They should be set aside to cool and picked up with caution. Use tongs or heat-protective gloves if necessary.
53. Never look into a container that is being heated.
54. Do not place hot apparatus directly on the laboratory desk. Always use an insulating pad. Allow plenty of time for hot apparatus to cool before touching it.
55. When bending glass, allow time for the glass to cool before further handling. Hot and cold glass have the same visual appearance. Determine if an object is hot by bringing the back of your hand close to it prior to grasping it.

QUESTIONS

56. Do you wear contact lenses?
    ☐ YES ☐ NO
57. Are you color blind?
    ☐ YES ☐ NO
58. Do you have allergies?
    ☐ YES ☐ NO

    If so, list specific allergies __________

    ____

    ____

AGREEMENT

I, ______________________________
________________(student's name) have read and agree to follow all of the safety rules set forth in this contract. I realize that I must obey these rules to ensure my own safety, and that of my fellow students and instructors. I will cooperate to the fullest extent with my instructor and fellow students to maintain a safe lab environment. I will also closely follow the oral and written instructions provided by the instructor. I am aware that any violation of this safety contract that results in unsafe conduct in the laboratory or misbehavior on my part, may result in being removed from the laboratory, detention, receiving a failing grade, and/or dismissal from the course.

____
Student Signature

____
Date

Dear Parent or Guardian:

We feel that you should be informed regarding the school's effort to create and maintain a safe science classroom/laboratory environment.

With the cooperation of the instructors, parents, and students, a safety instruction program can eliminate, prevent, and correct possible hazards.

You should be aware of the safety instructions your son/daughter will receive before engaging in any laboratory work. Please read the list of safety rules above. No student will be permitted to perform laboratory activities unless this contract is signed by both the student and parent/guardian and is on file with the teacher.

Your signature on this contract indicates that you have read this Student Safety Contract, are aware of the measures taken to ensure the safety of your son/daughter in the science laboratory, and will instruct your son/daughter to uphold his/her agreement to follow these rules and procedures in the laboratory.

____
Parent/Guardian Signature

____
Date