Chemistry 103 *(Rohde)*

Chapters 1, 2A, & 25A – Atoms and Elements

Class Notes

<u>*Suggested Reading*</u>*: All of Chapter 1* *Chapter 2 (Sections 2.1 – 2.5)* *Chapter 25 (Sections 25.1 – 25.3, & 25.5)*

I. Highlights from Chapter 1

<u>NOTE</u>: It is strongly recommended that you review the material in Chapter 1 of the text. This is basic material that should you should be familiar with from previous studies in chemistry. Although some of these topics will be reviewed during Chem 103, it will be to your advantage to be fluent in the vocabulary and skills presented in Chapter 1.

A. Basic Vocabulary

element = one type of atom, listed on the Periodic Table.

compound = two or more elements chemically combined.

molecule = group of atoms held together by chemical bonds.

molecular elements = elements that exist in nature as molecules

Diatomic: H_2, O_2, N_2, etc. ("HONBrICl + FAt")

Others: P_4 and S_8

ion = electrically charged atom or molecule.

monatomic ion = charged atom (Na^+, Cl^-, Ca^{2+}, O^{2-}, etc.)

polyatomic ion = charged molecule (NH_4^+, H_3O^+, SO_4^{2-}, NO_3^-, etc.)

cation = positive ion

anion = negative ion

ionic bond = strong force of attraction between oppositely charged ions

ionic compounds = molecules (substances) formed by ionic bonds

physical properties = observed and measured without changing the composition of the substance (color, odor, melting and boiling points, density, solubility, etc.)

chemical properties = observed and measured during a chemical change (combustibility, reaction rate, heat of reaction, etc.)

extensive properties = depend on the amount of substance present (mass, volume)

intensive properties = do not depend on the amount of substance present (density, (melting and boiling points, combustibility, etc.)

B. Dimensional Analysis

$$\boxed{[Given\ Value]\ \ x\ \ [Conversion\ Factor(s)]\ \ =\ \ [Answer]}$$

Ex #1: In 1618, the Pharmacopoeia of London defined the following mass units to be used in the preparation of drugs:

20 grains = 1 scruple
3 scruples = 1 drachm (dram)
8 drachm (drams) = 1 ounce
12 ounces = 1 pound *(Note the archaic difference)*

a) How many grains are present in 16 ounces?

b) How many pounds were there in 78.6 drams?

Ex. #2: At sea, distances are measured in "nautical miles" and speeds are expressed in units of "knots".

1 nautical mile = 6076.12 ft
1 knot = 1 nautical mile/hour *(exactly)*

a) How many miles in one nautical mile?

b) How many meters in one nautical mile?

c) A ship is traveling at a rate of 22 knots. What is the ship's speed in units of miles per hour?

d) Convert the ship's speed to units of meters per second.

C. Uncertainty in Measurement

Important: There is a certain amount of doubt (uncertainty) in any measurement due to:

(1) ______________________________ } "Determinate"

(2) ______________________________ }

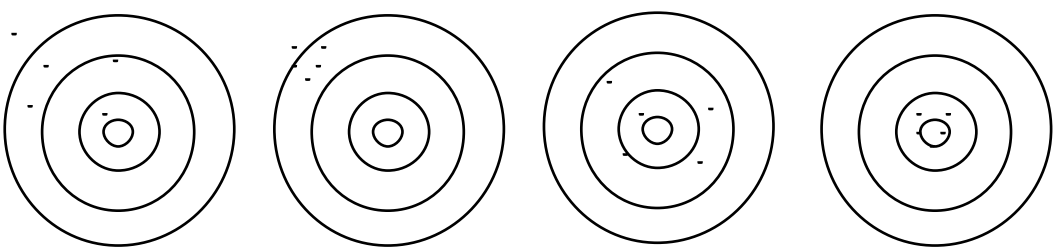
(3) ______________________________ "Indeterminate"

Note: **Standard Deviation** can be used to evaluate Indeterminate Errors *(See Text)*

D. Accuracy and Precision

Accuracy = ____

Precision = ____



Not Precise Not Accurate | Precise Not Accurate | Not Precise Accurate | Precise Accurate

- How are accuracy and precision reported?

E. Uncertainty (Significant Figures)

Reminder: 2 and 2.0 are NOT the same!

- Additional zeros to the right of the decimal point indicate greater accuracy.

2 2.0 2.00

- **Significant Figures** = "numbers that have meaning"

- **Simplified Rules for Determining Significant Figures**

  a) Non-zero digits are significant

  b) Zeros used as numbers are significant (“placeholders” and “cosmetic zeros” are not significant)

Ex. #3: Label each of the zeros in the following number and indicate significance.

$$0.02050$$

Ex. #4: Give the number of significant figures in each of the following

a) 250	_______	f) 2000	_______
b) 250.0	_______	g) 0.00200	_______
c) 0.025	_______	h) 2.0005	_______
d) 0.2000	_______	i) 2,000,000	_______
e) 0.2	_______	j) 2.00×10^6	_______

- **Exceptions to General “Rules”**

  a) **Defined Quantities** (There is no uncertainty here!)

  1 liter = 1000 mL *(Not necessary to write as 1.000×10^3 mL)*

  b) **Exact Values** (There is an assumed certainty here!)

  “There are only 10 eggs in this carton.” *(Not necessary to write as 1.0×10^1)*

F. Operations Involving Significant Figures

- **Addition and Subtraction**: In these operations, the number of places to the right of of the decimal point (or the accuracy) determines the number of significant figures reported. You cannot be more accurate than the least accurate measurement!

- Examples:

	2.0030	21	0.0055
	3.1	4.627	0.03012
	+ 4.051	+ 11.0	+ 0.010
Calculator Answer:	9.1540	36.627	0.04562
Reported Answer:	____________	____________	____________

- **Multiplication and Division**: In these operations, the number of significant figures reported is determined by the one that has the least number of significant figures.

	Calculator Answer	*Reported Answer*
3.2 x 2.060 x 1.44 =	9.49248	_______________
1.53 x 0.002 x 0.018 =	0.00005508	_______________
0.01543 x 0.00035 =	5.4005×10^{-6}	_______________

G. Rounding Rules

1) Rounding should be done at the end of a series of operations!

2) "Odd Up, Even Stays"

Ex. #5: Round each of the following to the tenths position.

0.350 = _____________

0.250 = _____________

0.251 = _____________

Ex. #6: (Progress Check) Give the number of significant figures in each of the following:

a)	453.2	______	f)	6.0002	______	k)	0.1	______
b)	0.00095	______	g)	816,000	______	l)	1,000,000	______
c)	700	______	h)	0.50031	______	m)	1.0×10^6	______
d)	700.1	______	i)	0.00400	______	n)	1.000×10^6	______
e)	700.00	______	j)	1.0	______	o)	1.010×10^{10}	______

II. <u>Historical Perspective</u>

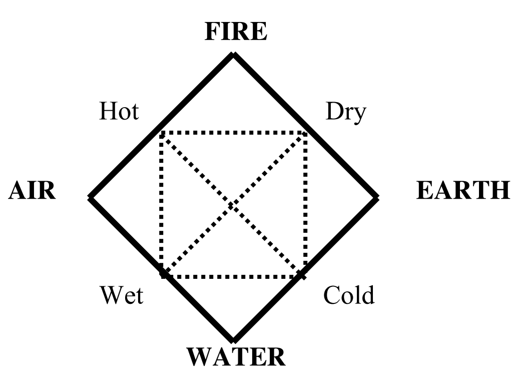
A. Chemistry has two roots:

1) **Craft Traditions** (Metallurgy, brewing, tanning, dyeing, etc. These provided a ***practical*** understanding of how matter behaves.)

   <u>Note</u>: The growth of chemistry is in some respects a reflection of the practical problems overcome in the course of human society's cultural and technical development. However, chemistry is also an expression of humankind's innate curiosity and desire to understand its surroundings without regard to the practical applications of that understanding.

2) **Philosophy** (Philosophers who concerned themselves with questions regarding the basic nature of matter.)

   - Early Greek philosophers (Plato, Aristotle) proposed that all of nature was composed of four elements:



   - Each element has two associated properties
   - Reaction Example: A log can burn to produce smoke (air), ashes (earth), and moisture (water)
   - This concept, and its associated logic persisted through the Middle Ages.

   - Operating from this framework, the **alchemists** found (isolated) many new elements and compounds in their attempts to produce gold.

B. Development

1) **Phlogiston Theory** = Flammable materials contained a substance called "phlogiston". Substances burn only as long as they have (contain) phlogiston. After the phlogiston has escaped, the substance is no longer flammable.

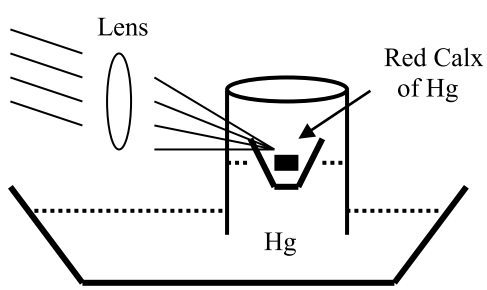
2) **Experiments of Joseph Priestly** (Mid 1700's)
(*or* "The beginning of the End of Phlogiston Theory")

- When wood is burned, it loses mass (loss of phlogiston). But, it was observed that when metals "burn" they get heavier and lose their metallic properties:

iron → rust *(called a "**calx**")*

*(We now think of this reaction as: **metal** → **metal oxide**.)*

- His experiment: Priestly converted the HgO (red calx of mercury) to $Hg_{(l)}$ and discovered that the gas produced during the reaction (He called it "dephlogisticated air", but we know it is $O_{2\,(g)}$), would make things burn much more vigorously.



- Priestly assumed that he had taken phlogiston out of the air and put it back into the mercury. Things would burn more rapidly because the "dephlogisticated gas" had a greater capacity to take phlogiston out of the burning substance.

3) Experiments of **Lavoisier** (Mid 1700's)

- Considered the father of modern chemistry.
- Made careful **quantitative measurements**.
- Discovered that burning produced <u>heavier</u> products, suggesting that something had been taken <u>out</u> of the air.
- Also, the mass lost by the air, equaled the mass gained by sample.
- Also, the products of combustion were the same whether carried out in air <u>or</u> "dephlogisticated air".
- These results contradicted Priestly's conclusions and marked the end of "Phlogiston Theory, and the beginning of "modern" science.

4) We now know that **Scheele**, in Sweden, actually performed the experiments of both Priestly and Lavoisier about two years earlier, and arrived at the same conclusions.

Note: *We are now moving into Chapter 2.*

III. The Basics of Atomic Theory

1) Early Greek and Egyptian Philosophers and "Scientists"

- Is there a "basic building block" of matter?
- Continuous or non-continuous theories of matter.

2) Dalton's Atomic Theory

Simplified Postulates:

(1) Matter is composed of tiny indivisible particles (atoms).
(2) Atoms cannot be changed during chemical reactions.
(3) Compounds are formed from combinations of atoms

3) The Basic Laws *(Historically, these laws contributed to Dalton's Theory)*

Law of Conservation of Mass = atoms are neither created or destroyed during chemical reactions.

Law of Constant (Definite) Composition = a compound always contains the same elements in the same proportions by mass.

Law of Simple Multiple Proportions = When elements combine to form more than one compound, the masses of the one element combine with a fixed mass of a different element are always in a small, whole-number ratio.

Example of how the Law of Multiple Proportions works:

Formula	Fixed Mass of Nitrogen	Resulting Mass of Oxygen	Ratio of Oxygen-to-Oxygen
NO	14		
NO_2	14		
NO_3	14		
N_2O	14		
N_2O_5	14		

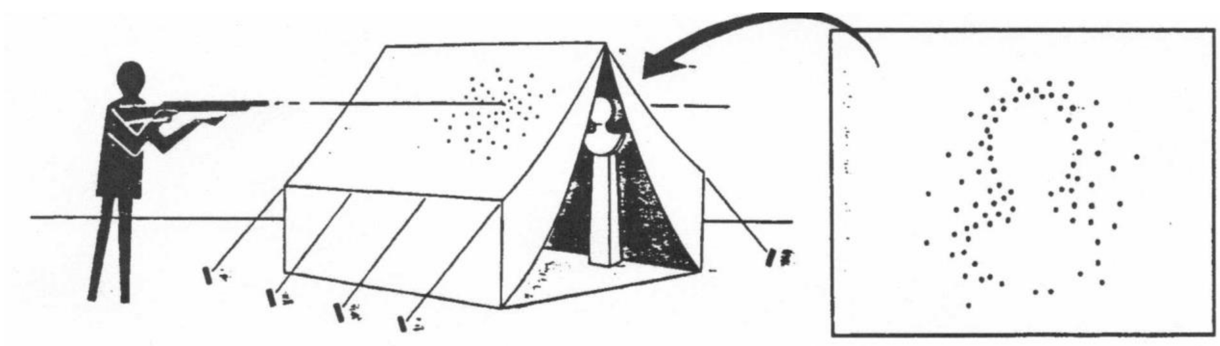
4) **Components of the Atom** (Once the theory of atoms was accepted, the search for subatomic particles had begun.)

   a) **Thomson** confirms the existence of electrons by showing a beam of electrons in a cathode tube could be bent with electric or magnetic fields. Develops model of atom as a jumbled mixture of positive and negative particles.

   b) **Millikan**'s "Oil-Drop" Experiment provided experimental support for Thomson's work.

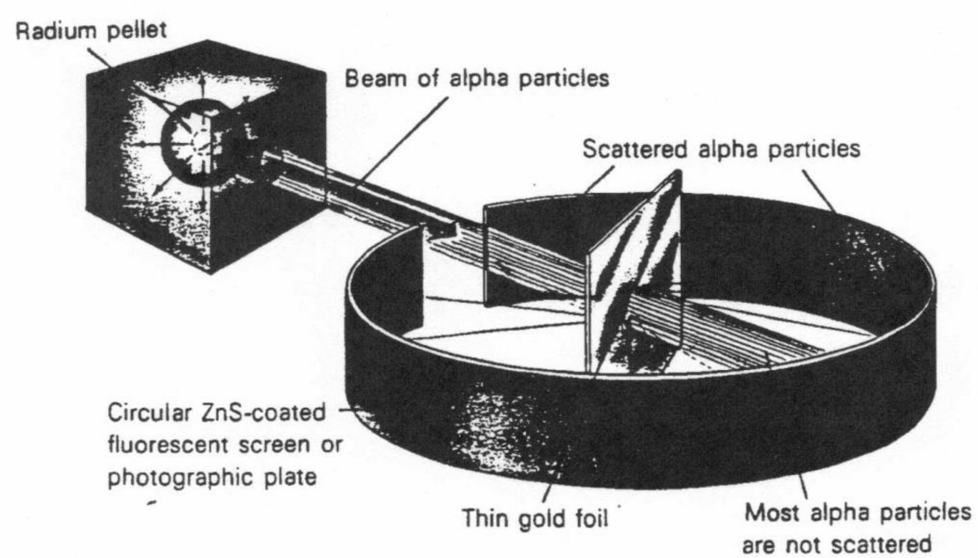
   c) **Rutherford's Experiment**:

   Analogy to Rutherford's scattering experiment: (a) firing rifle bullets at an unseen object in a tent, and (b) the bullet holes seen from the other side of the tent.

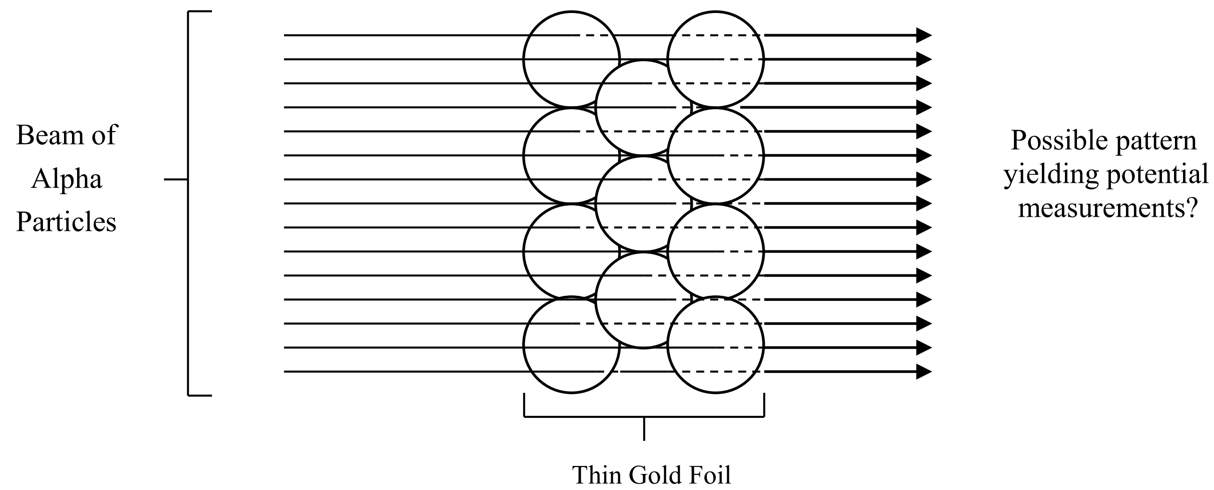


(a) (b)

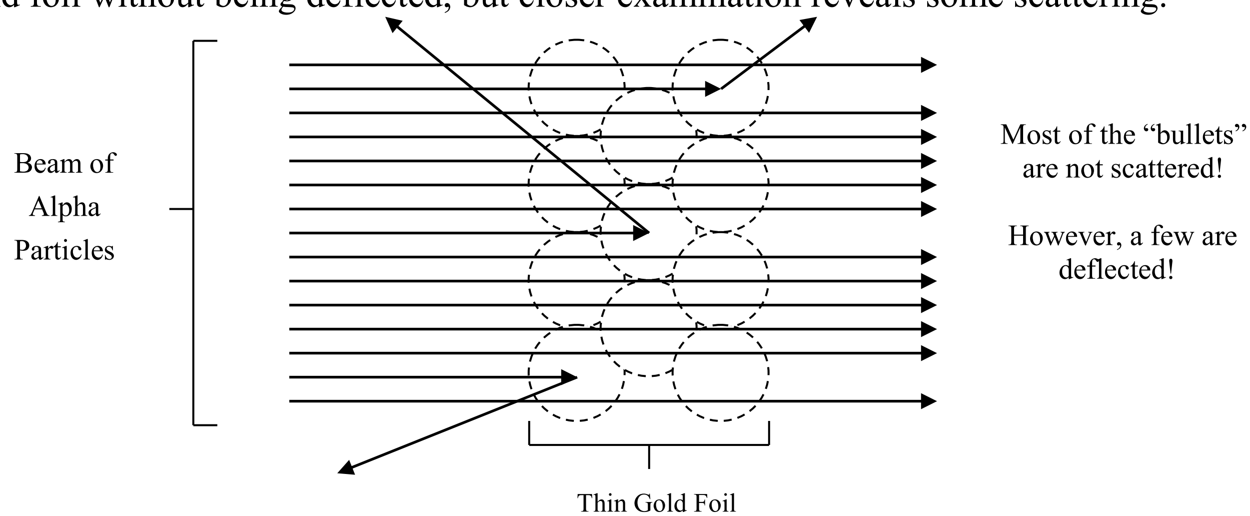
Below is a modern version of the Rutherford's apparatus for observing the scattering of alpha particles by gold foil. (<u>Note</u>: The entire apparatus is enclosed in a vacuum chamber)



<u>What He Expected</u>: The bombarding alpha particle "bullets" might produce a pattern that would possibly disclose the size, arrangement, position, etc. of the gold atoms.



<u>What He Discovered</u>: The alpha particle "bullets" appear to pass through Rutherford's gold foil without being deflected, but closer examination reveals some scattering.



d) <u>Conclusions from Rutherford's Experiment</u>:

(1)

(2)

(3)

e) **Chadwick**'s experiments identified and confirmed the presence of neutrons in the nucleus.

f) The work of these four (and a few others) brought us to our modern understanding of the atom

5) Summary – Models of the Atom

Dalton =

Thomson =

Rutherford =

<u>Note</u>: *We are now moving into Chapter 25.*

IV. Components of the Nucleus – Nuclear Symbols and Reactions

("We can better understand how things are put together by watching how they fall apart.")

- <u>Reminders</u>:

atomic number = number of protons (Find on Periodic Table)

mass number = sum of protons and neutrons (Periodic Table shows the "weighted average" of the mass numbers of the isotopes of a given element.)

isotope = same atomic number, but different mass numbers

radioactivity = process of self-decay (an unstable nucleus changes)

<u>Note</u>: In general, if an isotope is within the "**Belt of Stability**", it will be stable. Outside that belt, it will probably have too many or too few neutrons to be stable. *(See Page 23)*

A. **Nuclear Symbols**: $^{A}_{Z}X$

A = mass number
Z = atomic number
X = symbol of the element

(See Examples on Next Page)

Examples:

carbon – 12 = $^{12}_{6}C$ = 6 protons + 6 neutrons

carbon – 13 = $^{13}_{6}C$ = 6 protons + 7 neutrons

carbon – 14 = $^{14}_{6}C$ = 6 protons + 8 neutrons

Quick Review: Complete the following table for the given neutral atoms:

Element	Mass Number	Protons	Neutrons	Electrons	Nuclear Symbol
Hydrogen	1	______	______	______	
Helium	4	______	______	______	
Scandium	45	______	______	______	
Rhodium	103	______	______	______	

- Note:
  a) What makes a nucleus "unstable"?

  b) What happens when a nucleus is unstable?

  c) What is the "Belt of Stability"?

B. Three Common Forms of Radiation

(1) **Alpha Decay =** The unstable nucleus attempts to become more stable by breaking of a piece of itself. In alpha decay, the "chunk" of matter that breaks off is always an "alpha particle" (or helium nucleus)

alpha particle = helium nucleus = $^{4}_{2}He$

Example: $^{238}_{92}U \rightarrow {}^{234}_{90}Th + {}^{4}_{2}He$

Practice: $^{238}_{94}\text{Pu} \rightarrow$

$^{162}_{66}\text{Dy} \rightarrow$

$^{263}_{106}\text{Sg} \rightarrow$

(2) **Beta Decay =** This time the unstable nucleus releases a high-energy electron (beta particle)

$$\text{beta particle} = {}^{0}_{-1}e \text{ or } {}^{0}_{-1}\beta$$

Example: $^{234}_{90}\text{Th} \rightarrow {}^{234}_{91}\text{Pa} + {}^{0}_{-1}e$

Note Totals: The atomic number has gone up and the mass number stays the same! (This means a neutron has transformed into a proton.) This conversion happens every time there is a beta decay.

$$^{1}_{0}\text{n} \rightarrow {}^{1}_{1}\text{H} + {}^{0}_{-1}e$$

Practice: $^{43}_{19}\text{K} \rightarrow$

$^{147}_{61}\text{Pm} \rightarrow$

$^{37}_{18}\text{Ar} \rightarrow$

(3) **Gamma Decay =** This time the unstable nucleus simply releases energy. The nucleus does not change its identity, but simply goes from a "higher" energy-state to a "lower" energy-state. The released energy is the "gamma ray"

$$\text{gamma ray} = {}^{0}_{0}\gamma$$

Example: $^{254}_{102}\text{No}^{*} \rightarrow {}^{254}_{102}\text{No} + {}^{0}_{0}\gamma$

[* = indicates that the nucleus is in a "high-energy" state]

C. Some Other Examples of Nuclear Reactions

- Rutherford Transmutation/Bombardment:

  Example: $^{14}_{7}N + ^{4}_{2}He \rightarrow ^{17}_{8}O + ^{1}_{1}H$

- Neutron Bombardment: *Example:* $^{14}_{7}N + ^{1}_{0}n \rightarrow ^{14}_{6}C + ^{1}_{1}H$

- Electron Capture: *Example:* $^{7}_{4}Be + ^{0}_{-1}e \rightarrow ^{7}_{3}Li$

- Positron Emission: *Example:* $^{207}_{84}Po \rightarrow ^{0}_{+1}\beta + ^{207}_{83}Bi$

Practice: Complete each of the following nuclear equations, and identify the type of nuclear reaction:

Equation	Type
1) $^{116}_{49}In \rightarrow ^{0}_{-1}\beta +$ ______	__________________
2) $^{198}_{78}Pt \rightarrow ^{4}_{2}He +$ ______	__________________
3) $^{27}_{13}Al + ^{4}_{2}He \rightarrow$ ______ $+ ^{1}_{0}n$	__________________
4) ______ $+ ^{1}_{0}n \rightarrow ^{144}_{58}Ce + ^{90}_{38}Sr + 2\,^{1}_{0}n + 4\,^{0}_{-1}e$	__________________
5) $^{238}_{92}U +$ ______ $\rightarrow ^{246}_{98}Cf + 4\,^{1}_{0}n$	__________________

D. Summary of Particles

$^{1}_{1}H$ = proton

$^{4}_{2}He$ = alpha particle

$^{1}_{0}n$ = neutron

$^{0}_{-1}e$ = electron

$^{0}_{-1}e = ^{0}_{-1}\beta$ = beta particle

$^{0}_{+1}e = ^{0}_{+1}\beta$ = positron

$^{0}_{0}\gamma$ = gamma ray

$^{A}_{Z}X$ = generic form for elements

Note: You should probably memorize these (?)

V. Atomic and Formula Masses *(Back to Chapter 2)*

Reminder: Mass numbers cannot be found on the Periodic Table. The Periodic Table gives the relative (weighted) average of the naturally occurring isotopes of a given element, which is called "atomic weight" or "atomic mass".

A. **Formula for determining atomic mass:** ("AM" = atomic mass, "I" = isotope)

$$\text{AM} = \text{AM}_1\left(\frac{\% I_1}{100}\right) + \text{AM}_2\left(\frac{\% I_2}{100}\right) + \text{AM}_3\left(\frac{\% I_3}{100}\right) + \text{etc}$$

Note: This formula is not on your Blue Reference Sheet

Examples:

Ex #7:

Isotope	Atomic Mass	Percent
Ne – 20	20.00	90.92
Ne – 21	21.00	0.26
Ne – 22	22.00	8.82

What is the atomic mass of neon, based on this data?

Ex #8: Antimony occurs in nature as a mixture of two isotopes ($^{123}_{51}Sb$ and $^{121}_{51}Sb$). The Sb –121 isotope has a mass of 120.9 amu and an abundance of 57.25%

a) What is the abundance of the $^{123}_{51}Sb$ isotope?

b) What is its atomic mass?

Ex #9: Data Given:

Isotope	Atomic Mass	Percent
${}^{28}_{14}Si$	27.98	92.2
${}^{29}_{14}Si$	28.98	4.7
${}^{30}_{14}Si$	29.97	3.1

Calculate the atomic mass of silicon from this data.

Note: Usually, atomic mass increases with atomic number. However, because of isotopic abundance, it is possible to have an **anomaly**.

Element	Atomic Number	Atomic Mass
Ar	18	39.95
K	19	39.10

If the Periodic Table were arranged based on increasing mass, Ar and K would be Listed in reverse order. There are six other similar anomalies on the Table. *(Can you find them?)*

VI. The Mole

Remember: A "mole" is just a grouping! (Like "dozen" or "pair")

- The number of things in a mole grouping is known as "Avogadro's Number".
- The common symbol for Avogadro's Number is "N" or "N_A".

1 dozen = 12

1 mole = $\mathbf{6.022 \times 10^{23}}$ *(Value we will use)*

= $6.02216940 \times 10^{23}$ *(CRC Handbook)*

A. Related Terms

molar mass = mass of one mole of a substance (units = grams/mole)

formula mass = same as molar mass, but without units

Reminder: The power of the "mole" is that we can use the number value from the amu relative scale on the Periodic Table and simply change the units! The mole is just the right size group to make this happen!

One atom of H = 1.008 amu	One mole of H = 1.008 g/mole
One molecule of HCl = 36.46 amu	One mole of HCl = 36.46 g/mole

- NOTE: *Numerically* all of the following are equal:

molar mass	=	**atomic/molecular mass**	=	**formula mass**
(units = g/mole)		(units = amu)		(no units)

B. Using Molar Mass ("Gram → Mole" and "Mole → Gram" Conversions)

- If you have ever had a chemistry course before, you know these are the most common calculations in chemistry.
- We use "molar mass" to get to, or from, mass and moles. This is critical in the lab where we use balances (that measure in "mass") and relate our measurements to reactions (which are described in "moles").

C. Simple (Elemental) Conversions

Ex #10: During a chemical reaction, 1.271 grams of copper were produced. How many moles of copper have been formed?

Ex #11: A certain step in a lab procedure asks you to use 0.250 moles of zinc in a certain reaction. How many grams zinc will you weigh on the balance?

D. Masses of Individual Atoms and Molecules

- We normally use the relative scale (amu or just "u") for individual atoms or molecules, because they are way too small for "grams".

- However, in certain situations, it may be useful or necessary. (Historically, these calculations were used to determine Avogadro's Number!)

Ex #12: What is the mass of a single hydrogen atom in units of grams?

Ex #13: How many atoms are in one gram of nickel?

VII. Periodic Table Review

A. Historical Background

- **Dalton**: Found accurate atomic masses for about 20 elements.

- **Dobereiner**: Proposed "Triads", groups that had similar properties.

- **Newlands**: Early table; had seven families going up to calcium. Also proposed "Law of Octaves" (every eighth element has similar properties).

- **Mendeleev**: Arranged elements by increasing <u>atomic</u> <u>mass</u> *and* by families with similar properties; better than Newlands because it went beyond calcium. Also predicted properties (and left spaces for) some undiscovered elements. However, he didn't know that the Noble gases existed.

- **Mosely**: Arranged the table based on increasing atomic number (based on x-ray data) and included Noble Gases. Also left gap for technetium.

- **Seaborg**: Gives table its "modern arrangement" and predicts the existence of the "super-heavies".

B. Review “**Nomenclature**” of the Periodic Table

- In the modern Periodic Table (and most other periodic tables), **elements with similar chemical properties are arranged in vertical columns**. This periodic repetition of properties based on the atomic number of the elements is known as **chemical periodicity**.
- Vertical columns are called: ___________________ or _____________________
- Horizontal rows are called: ___________________ or _____________________
- The diagonal “staircase” on the right side of the table is a “dividing line” that separates the _______________________ from the ________________________.
- Hydrogen is a unique element. It is a “family of one”. Some of its properties and reactions are similar to Group 1(IA) metals. Other properties and reactions are similar to Group 17 (VIIA) nonmetals. Still other properties are unique to hydrogen itself. Therefore, it should really be placed separately on the Periodic Table.

Location	**Name**
Left of Staircase	
Right of Staircase	
Bordering Staircase	Exceptions: ______________________
Group 1 (IA)	
Group 2 (IIA)	
Groups 1 and 2 Combined	
Group 17 (VIIA)	
Group 18 (VIIIA)	
“Middle Block” [Groups 3 (IIIB) – 12 (IIB)]	
1st Partial Row Below Table	
2nd Partial Row Below Table	
1st and 2nd Partial Rows Combined	
Elements below the Staircase (Except for metalloids)	
Groups 1, 2, and 13-18	
Groups 8–10 (VIIIB)	

C. **Other Notes About the Periodic Table**

- Symbols of Elements:
  First Letter: H, B, C, etc.
  First Two Letters: He, Li, Ne, etc.
  First Letter and Another Letter: Mg, Cl, Zn, etc
  Latin: Na (natrium), Cu (cuprum), Au, (aurum), Pb (plumbum), etc.
  German: W (wolfgram)
  Lab Names of New Elements: Uuo (ununoctium = 118), etc.

- Phase Color Code: (Many Periodic Tables have a color code for solid, liquid, or gas phases at room temperature. Typically, the synthetic elements are designated with a different color.)

- Origins of Element Names: *See* **webelements.com**
  (Have you ever heard of "Copernicium"?)

VIII. "Enrichment"

A. Phases of Matter

B. More About Quarks

- Quarks come in six "flavors":

Up	Top (Truth)	Strangeness
Down	Bottom (Beauty)	Charm

- Protons and neutrons are NOT indivisible particles. We now know that they are "**quark clusters**". A proton is an UUD combination of quarks and a neutron is a UDD combination

(Continued on next page.)

- In protons and neutrons, "Up" quarks have a charge of $+\frac{2}{3}$, and "Down" quarks have a charge of $-\frac{1}{3}$.
- Sketch and net charge calculation:

IX. *(Addendum)*: **Balancing Equations** *(An application of the Law of Conservation of Mass)*

Law of Conservation of Atoms (Law of Conservation of Mass)

Since atoms cannot be created or destroyed in a chemical reaction, the total number of each kind of atom must be the same before and after the reaction.

Note: Since atoms are what give mass to matter, this law is just a re-statement of the Law of Conservation of Mass

Atoms can be re-arranged, but totals before and after the reaction are the same. (It is this fact that allows us to "balance" equations.)

$\boxed{A}$ $\boxed{C}\boxed{A}\boxed{R}$ $\boxed{T}\boxed{A}\boxed{K}\boxed{E}\boxed{N}$ = $\boxed{N}\boxed{E}\boxed{A}\boxed{R}$ $\boxed{A}$ $\boxed{T}\boxed{A}\boxed{C}\boxed{K}$

A. Important Reminders

Balancing equations is a "Trial-and-Error" process.

We have already done some balancing when we did dissolving and precipitation equations.

Balancing means adjusting counts (or totals), so we use **coefficients**. (Subscripts are used when we write formulas, but NOT when we balance equations!)

A balanced chemical equation:

- gives formulas of reactants and products
- gives relative proportions of species in the reaction
- does NOT give actual gram amounts or a gram-ratio!

Practice: Balance the following equation

$$Fe + Al_2O_3 \rightarrow Al + Fe_2O_3$$

Remember: When balancing equations, we cannot change formulas!!! We can only use coefficients!

Additional Practice:

Examples: Balance each of the following equations.

a) $SiO_2 + C \rightarrow SiC + CO$

b) $KCN + FeCl_2 \rightarrow K_4Fe(CN)_6 + KCl$

c) $NO_2 + H_2O \rightarrow HNO_3 + NO$

d) $Mg + O_2 \rightarrow MgO$

e) $N_2 + O_2 \rightarrow NO$

f) $Na + Cl_2 \rightarrow NaCl$

g) $Fe + O_2 \rightarrow Fe_2O_3$

h) $CuS + O_2 \rightarrow CuO + SO_2$

i) $N_2H_4 + O_2 \rightarrow N_2 + H_2O$

Fractions are OK!

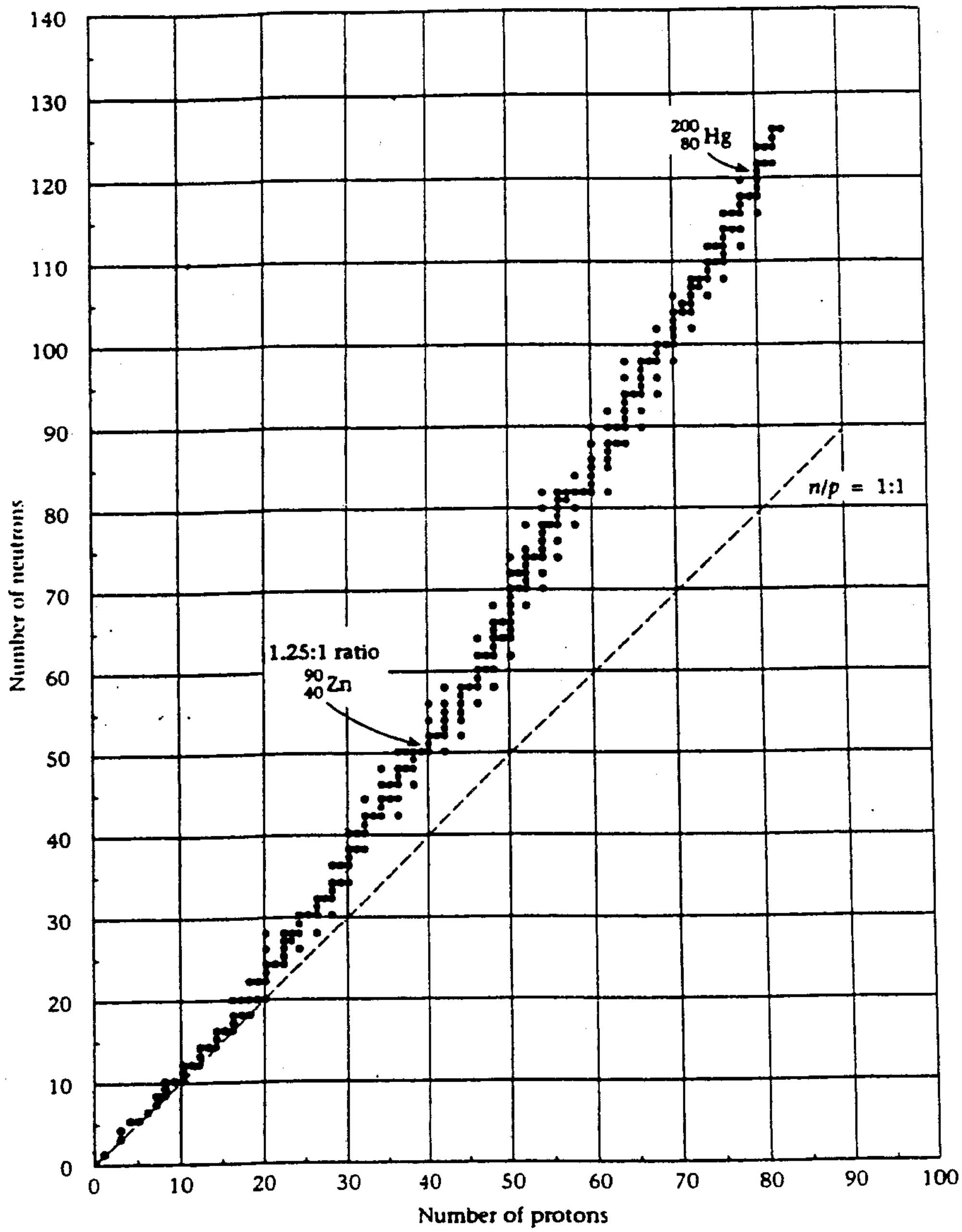
- Sometimes fractions can be used to make the balancing process faster and more efficient. For example, (d) above could have been balanced as:

$$Mg + \tfrac{1}{2}\,O_2 \rightarrow MgO$$

- Although this may seem odd if you think of individual atoms or molecules, coefficients can also represent moles. So even if you can't have ½ molecule, you could have ½ of a mole (like having ½ dozen).

X. Appendices

Belt of Stability:



Note: Stable isotopes have neutron-to-proton ratios that fall within a narrow range, referred to as a "**belt of stability**". For light isotopes of small atomic number, the stable ratio is 1.0. With heavier isotopes, it increases to about 1.5. There are no known stable isotopes for elements of atomic number greater than 83 (Bi).

Summary: Equations and Constants

***Note*: The following will *not* be provided on Exams in this course. It is recommended that these items be memorized. (Compare to your "Blue Reference Sheet")**

Density: $$D = \frac{mass}{volume}$$

Avogadro's Number: $$\frac{6.022 \ x \ 10^{23}}{1 \ \text{mole}}$$

Isotopic Abundance: $$\text{AM} = \text{AM}_1 \left(\frac{\% I_1}{100} \right) + \text{AM}_2 \left(\frac{\% I_2}{100} \right) + \text{AM}_3 \left(\frac{\% I_3}{100} \right) + \text{etc}$$

Error Analysis:

$$\textbf{Error} = \textit{Measured Value} - \textit{Actual Value}$$

$$\textbf{\% Error} = \textbf{\% Difference} = \left| \frac{Measured - Actual}{Actual} \right| \ x \ 100$$

$$\textbf{Magnitude of Range} = \textit{Highest Value} - \textit{Lowest Value}$$

$$\textbf{\% Range} = \frac{Highest\,Value - Lowest\,Value}{Average\,Value} \ x \ 100$$