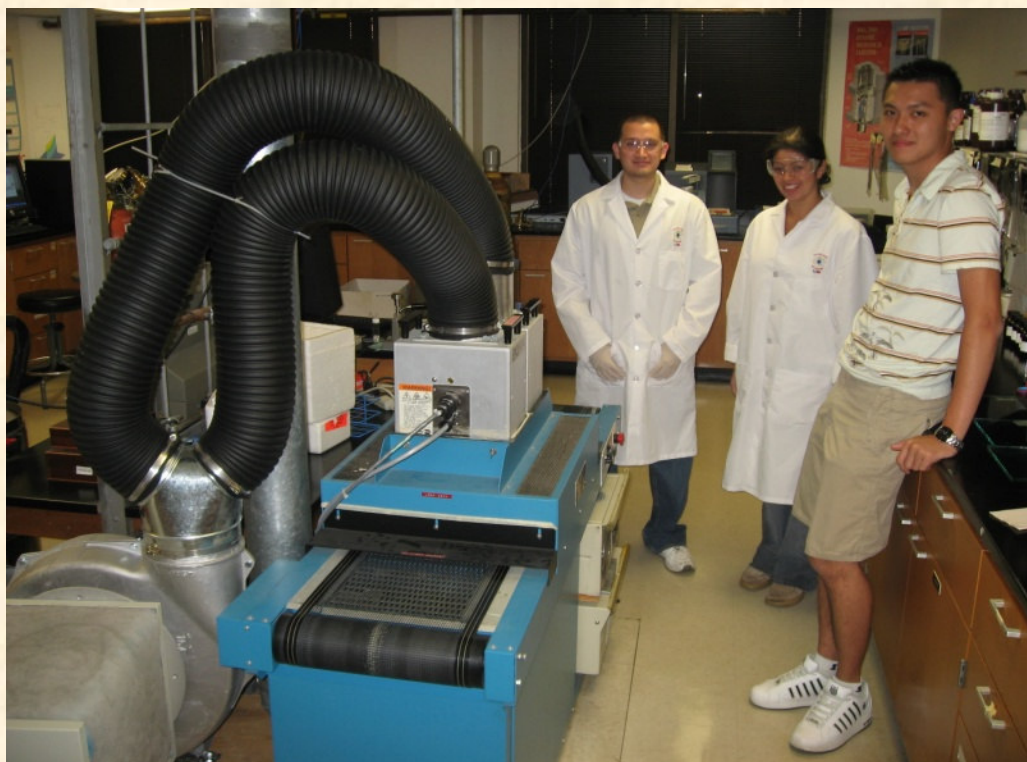


CHEMISTRY 1307

General Chemistry I



CHEMISTRY 1307

General Chemistry I

Instructor: Dr. Byron K. Christmas

http://www.uhd.edu/academic/colleges/sciences/naturalscience/BKC_Homepage.htm

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Office Hours: See Syllabus – *On Course Website!*

Text: *Chemistry: A Molecular Approach* – Nivaldo J. Tro

<http://www.privatehand.com/flash/elements.html>

CHEMISTRY 1307

Pre-requisites: CHEM 1305 or High School Chemistry

**Co-requisite: Credit or Enrollment in MATH 1301 and
CHEM 1107**

**On-Line Mastering General Chemistry Homework/Quizzes:
*More Later!***

Examinations: 4 Examinations & 1 Final Examination

CHEMISTRY 1307

EVALUATION OF PERFORMANCE:

4 Examinations (<i>Final Exam Replaces Lowest</i>)	44%
5 <u>MGC</u> Quizzes (<i>Lowest Score Drops!</i>)	24%
11 <u>MGC</u> Homework Assignments	11%
1 Final Examination (<i>Standardized; Comprehensive!</i>)	21%

Why Study Chemistry?

1. To better understand the world: what it is made of and how it works.
2. Because it is the most practical and relevant of the sciences - chemistry is the study of ***EVERYTHING!***
3. It is the ***“Central Science”*** - All other sciences intersect at and depend on chemistry.
4. It is essential to the national ***and local*** economies.
(Houston is at the center of the world’s largest petro-chemical complex)

Why Study Chemistry?

5. It is required for virtually every major involving science, mathematics, or engineering.
6. An awareness of the principles of chemistry is essential to being an informed and responsible ~~citizen~~ **Consumer** in a highly technical society.
7. It is incredibly fascinating and a lot of fun!
8. *And MANY MORE!*

The Language of Chemistry

- ☞ **Alphabet - Chemical Symbols of the Elements** (*Memorize the first 112 names and symbols*)

H Mg Sc Zr Ta Unh Nd Np Uuu Re

- ☞ **Words - Chemical Formulas of Compounds**

H₂O HNO₃ NaCl C₁₂H₂₂O₁₁ NH₄ClO₄

- ☞ **Sentences - Chemical Equations**

2 Na (s) + Cl₂ (g) -----> 2 NaCl (s) + *energy*

- ☞ **Paragraphs - Reaction Mechanisms**

(Dealt with in Chapter 13 in CHEM 1308)

- ☞ **Using the Language to Express Ideas -**

Definitions, concepts, mathematical skills, etc.

*Master the Language and
You Master the Subject!*

The Periodic Table of the Elements

The Scientific Method



- 1. Collect Facts or Data (*Observe!!*)**
- 2. Search for Generalizations or Laws to Summarize the Facts.**
- 3. Freely Use Your Imagination to Construct Theories or Models of Nature that Will Account for the Laws.**
- 4. Test Theories/Hypotheses for Accuracy.**
- 5. Modify Theories/Hypotheses as Necessary Based on Your Test Results.**

Chapter 1
*Matter, Measurement, &
Problem Solving*

CHEMISTRY: *(Memorize this Definition!)*

Chemistry is the study of the **properties, composition, and structure of *matter***, the **physical and chemical changes** it undergoes, and the ***energy*** liberated or absorbed during those changes.

Got your notebooks? Write this DOWN!

MATTER:

Matter is anything that occupies space and has mass.

Examples:	chairs	gasoline	clothes	batteries
people	the earth	paint	paper	oxygen
water	salt	aluminum	air	rocks

Let's Think About Matter!

What Are the States of Matter?

A. Plasma

B. Gases

C. Liquids

D. Solids



Page 8 in Textbook

***The Law of Conservation of Matter* - Matter is neither created nor destroyed during a chemical reaction or a physical change.**

ENERGY:

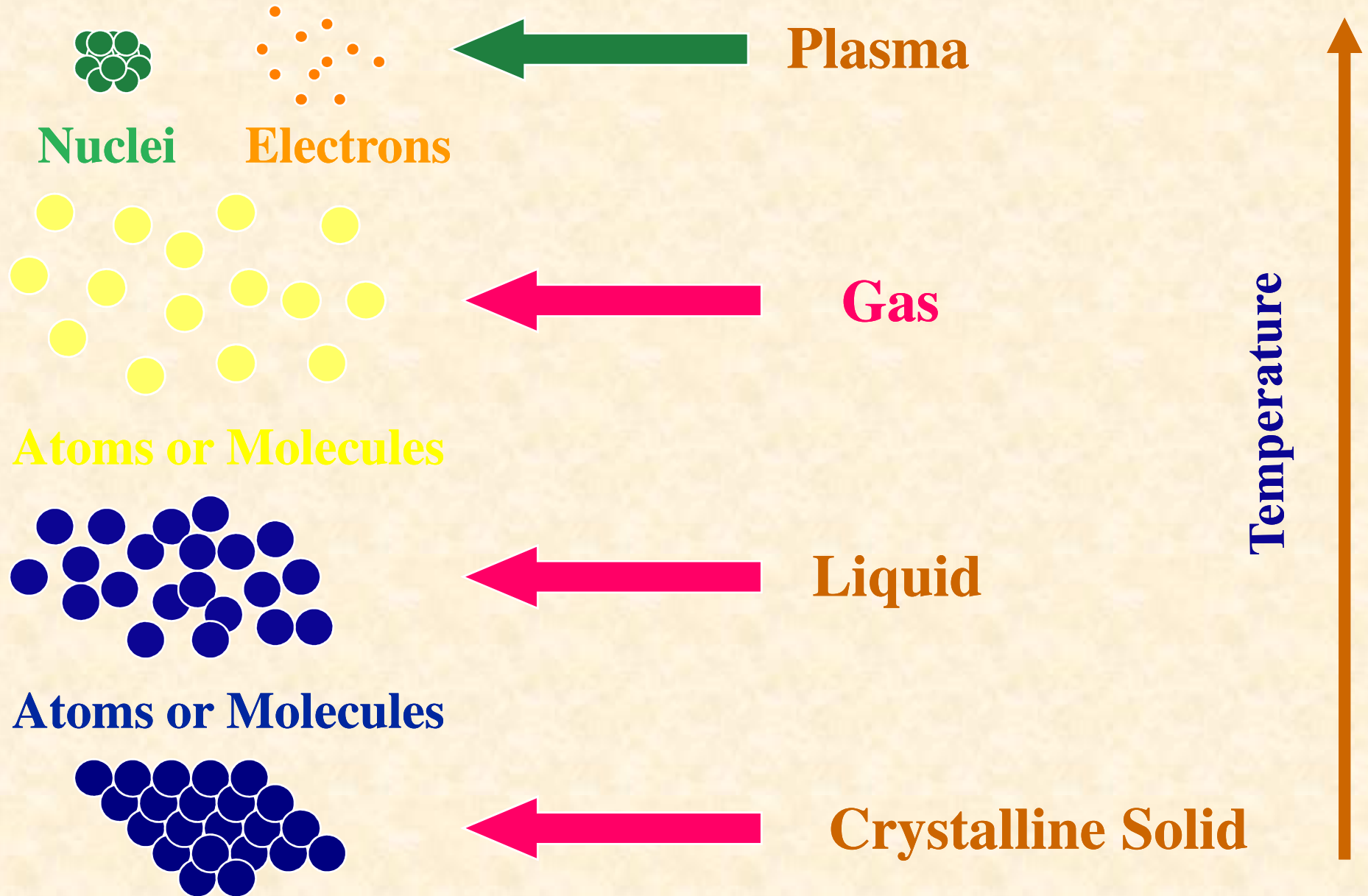
Energy is the ability to do work or generate *thermal energy*.

***The Law of Conservation of Energy*- Energy is neither created nor destroyed during a chemical reaction or a physical change. It can only be changed from one form into another.**

***The Law of Conservation of Matter-Energy* - The combined amount of matter and energy in the universe is constant.**

$$E = mc^2$$

Atomic and Molecular Concepts

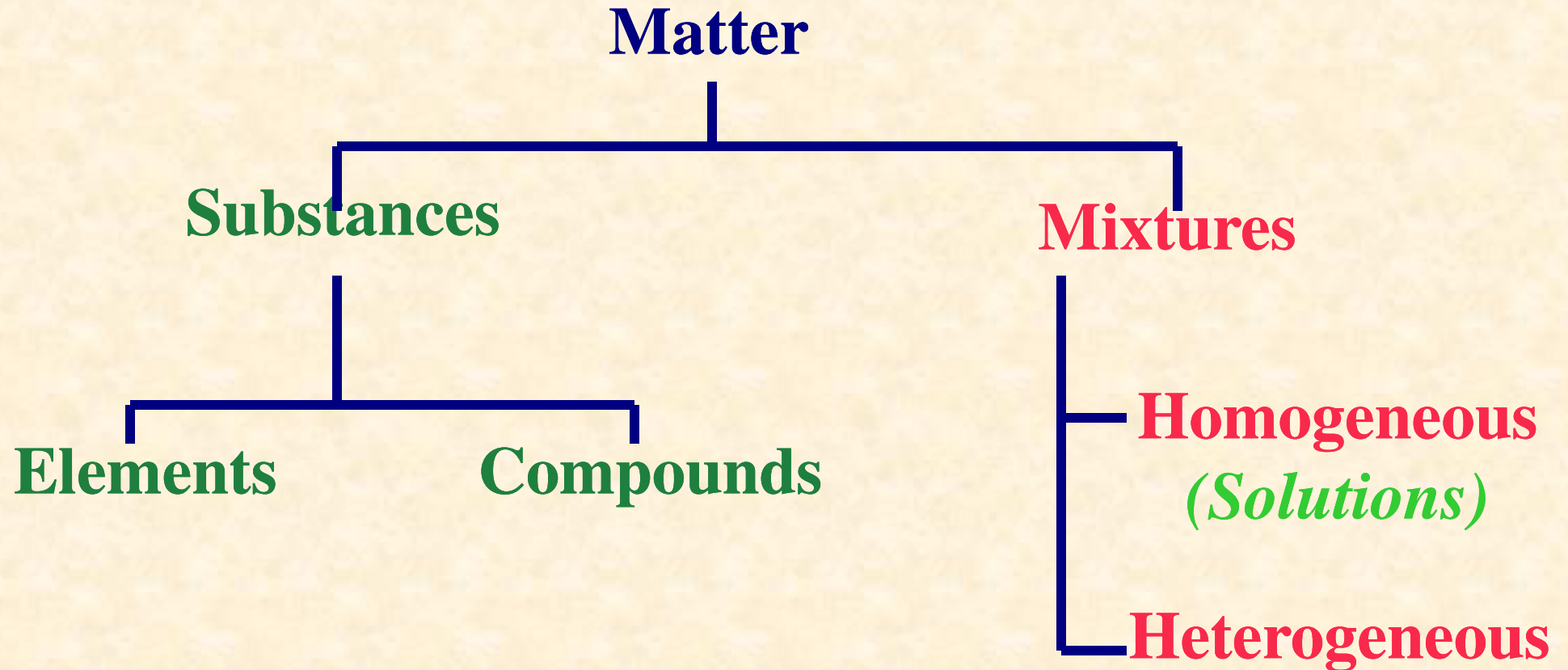


Atomic and Molecular Concepts

2. Classification of Matter

- **Substance** - A distinct type of matter. All samples of a substance have the same properties. **Elements** and **compounds** are **substances**.
- **Mixture** - A sample of matter consisting of two or more **substances** which are **NOT** chemically combined.

Classification of Matter



Memorize

Page 9 in Textbook

Classification of Matter (Substances)

Substances -

- **Element** - A substance that cannot be broken down (decomposed) into simpler substances by chemical reactions.
- **Compound** - A substance composed of two or more elements chemically combined in fixed ratios by mass.

Water - H_2O

Carbon dioxide - CO_2

Sodium Chloride - NaCl

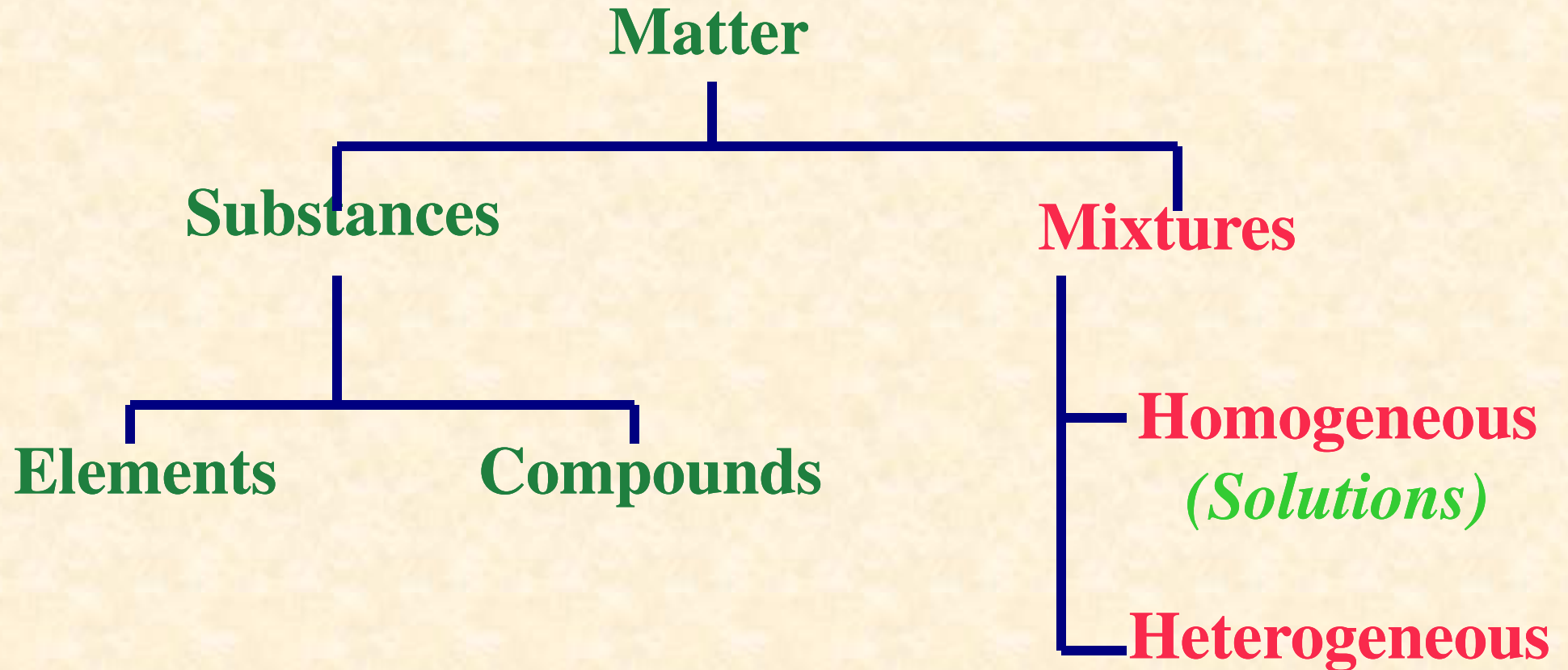
Iron(II) sulfide - FeS

Classification of Matter (Mixtures)

Mixtures -

- ➡ **Homogeneous** - A mixture having only one phase; it is uniform (the same) throughout and has the same properties throughout. These are called *Solutions*.
- ➡ **Heterogeneous** - A mixture with more than one phase. It is non-uniform and does **NOT** have the same properties throughout.

Classification of Matter



Memorize

Matter and Change

Physical Change - A change in which each **substance** involved in the change retains its original identity and no new **elements or compounds** are formed.



Triclosan???

Triclosan

<http://en.wikipedia.org/wiki/Triclosan>

<http://antoine.frostburg.edu/chem/senese/101/consumer/faq/triclosan.shtml>

Matter and Change

Page 10 in Textbook *Physical Changes*

Matter and Change

Physical Changes

Page 11 in Textbook

Matter and Change

Physical Changes

Page 13 in Textbook

Conceptual Connection 1.3

*Which one below represents
 H_2O after vaporization?*

H_2O (l) molecules in pan



Matter and Change

☞ **Chemical Change** - A change in which one or more **elements or compounds (substances)** are formed.



“Reacting”



Matter and Change

*Chemical
Change!*



Matter and Change

Page 12 in Textbook



Properties of Matter

☞ **Physical Properties - Properties that do NOT involve substances changing into other substances.**

Melting Point

Boiling Point

Temperature

Density

Mass

Volume

☞ **Chemical Properties - Properties that involve substances changing into other substances.**

Chemical Reactivity

Reduction Potential

Flammability

Oxidation Potential

Measurement

☞ Chemistry is an *Observational* science.

☞ Chemistry is a *Quantitative* science.

☞ **Measurement** - A quantitative observation.

4.57 mL

103.3 °F



Measurement

All measurements have three parts:

1. A value

2. Units

3. An Uncertainty

26.9762 g

Examples:

33.2 mL

72.36 mm

426 kg

31 people

Measurement

Systems of Units - Standards of Measurement

1. The Need for Standards

2. The English System *(What a pain!!!)*

12 in/ft 3 ft/yd 5280 ft/mi

16 fl.oz/pt 2 pts/qt 4 qt/gal

16 oz/lb 2000 lb/ton

3. The Metric System - A decimal system

meter (m) - Length

liter (L) - Volume

gram (g) - Mass

Measurement

Metric Examples:

 $1 \text{ m} = 1000 \text{ mm}$ $1 \text{ mL} = 0.001 \text{ L}$

 $1 \text{ kg} = 1000 \text{ g} = 1\,000\,000 \text{ mg}$

 $10 \text{ cm} = 0.01 \text{ m} = 0.000\,01 \text{ km}$

 $23 \text{ kL} = 23\,000\,000\,000 \mu\text{L}$

4. The SI System - Système International d'Unitès

A. A complete system of units adequate for the entire realm of physical science.

SI System of Measurement

Prefixes for the Metric and SI Systems

Page 16 in Textbook!



SI System of Measurement

Prefixes for the Metric and SI Systems

Page 19 in Textbook!



SI System of Measurement

Common Conversions for the Metric and SI Systems

Page 19 in Textbook!



SI System of Measurement

5. Rules for Using the SI Systems

1. Use only singular form of units and do **NOT** use a period after the symbol for the unit.

2. Use a dot on the base line for the decimal point.

23.6 m

not

~~23,6 m~~

3. Group digits in threes around the decimal point and do **NOT** use commas.

200 000.000 003 km

SI System of Measurement

4. Do NOT use spaces for four-digit measurements.

1645 mL or 0.2367 μg

5. Do NOT use the degree sign ($^{\circ}$) for temperature recorded for the Kelvin temperature scale.

78.6 K not ~~78.6 $^{\circ}$ K~~

Measurement

Conversion Factors - *A fraction whose numerator and denominator contain the same quantity expressed in **different** units.*

$$1 \text{ mile} = 5280 \text{ ft}$$

$$\frac{1 \text{ mile}}{5280 \text{ ft}} = \frac{5280 \text{ ft}}{1 \text{ mile}} = 1$$

$$1 \text{ cm} = 0.01 \text{ m}$$

$$\frac{1 \text{ cm}}{0.01 \text{ m}} = \frac{0.01 \text{ m}}{1 \text{ cm}} = 1$$

$$1 \text{ in} = 2.54 \text{ cm}$$

$$\frac{2.54 \text{ cm}}{1 \text{ in}} = \frac{1 \text{ in}}{2.54 \text{ cm}} = 1$$

Measurement

6. Uncertainty in Measurements -

Exact Measurements: *Measured values determined by counting or when a value is defined.*

Examples: 31 people 27 rocks
 2.54 cm = 1 in $10^6 \mu\text{L} = 1\text{L}$

The uncertainty in these measurements = 0

Non-exact Measurements: *All other measurements. The last digit recorded is uncertain; it is estimated!!*

Examples: 27.5 g 32.7 mm 12 467 km
 1.156 x 10^2 mL

Accuracy, Precision, & Sensitivity

Accuracy - The degree to which a measured value agrees with the true or “accepted” value.

Precision - The reproducibility of a measured value.

Sensitivity - The “fineness” of a measured value; the number of significant figures it has.

23.5673 g is a more sensitive measurement than 23.57 g.

Accuracy vs. Precision

Page 27 in Textbook!

Measurement

Significant Figures: *Each digit obtained as a result of a measurement. This includes all of the **certain** digits and the first **uncertain** digit. The number of significant figures in a measurement is an indicator of the **SENSITIVITY** of the measurement.*

1.5 How many significant figures are in the following:

65 mL

2

173.4 g

4

12.2 m

3

1×10^9 ns

1

Measurement

The Problem with Zero:

207.1 mm

0.002 36 mm

260.1 mm

0.123 00 mm

2040.0 mm

3600 mm

Rules for Significant Figures:



All non-zero digits are significant.

25.79 km

27 mL



A zero between other significant figures is significant.

207.9 nm

100.7 mL

Measurement

Initial zeros are NOT significant.

0.001 23 cm³

Final zeros after the decimal point ARE significant.

23.100 ps

Final zeros in a measurement with no decimal point may or may not be significant.

3200 cm (*might have 2, 3, or 4 significant figures!!*)

Exact measurements have an infinite number of significant figures. (*They are CERTAIN!!*)

Measurement


Significant Figures in Calculations: *In a measurement, the last significant figure is assumed to be **uncertain**.*

*The result of a calculation involving measured values can be no more **certain** than the least certain measurement.*

*The number of significant figures in a result depends on the number of significant figures in the measurement **and** on the mathematical operation being performed.*

Measurement


Significant Figures in Calculations:

-  **Addition and Subtraction** - *A sum or a difference of two or more measurements has the same number of decimal places as the measurement with the least number of decimal places.*

$$35.2 \text{ mL} + 0.34 \text{ mL} = \underline{\hspace{2cm}} \text{ mL}$$

$$1.00794 \text{ u} + 1.00794 \text{ u} + 15.99994 \text{ u} \\ = \underline{\hspace{2cm}} \text{ u} \quad \textit{u = atomic mass units}$$

Measurement

 **Multiplication and Division** - *A product or quotient of two or more measurements has the same number of significant figures as the measurement with the least number of significant figures.*

$$\begin{aligned}\text{density} &= (9.5760 \text{ g}) / (12.2 \text{ mL}) \\ &= \underline{\hspace{2cm}} \text{ g/mL}\end{aligned}$$

 **Round-off Rules** - *For digits 0 - 4, do not round up. For digits 5 - 9, round up.*

Measurement

1.6 Round-off the following to two decimal places:

$$23.04439 \text{ g} = 23.04 \text{ g} \quad 65.891 \text{ mL} = \underline{\hspace{2cm}} \text{ mL}$$

$$45.106 \text{ ms} = \underline{\hspace{2cm}} \text{ ms} \quad 30.1149 \text{ kg} = \underline{\hspace{2cm}} \text{ kg}$$

$$37.995 \text{ ng} = \underline{\hspace{2cm}} \text{ ng}$$

6. **Dimensional Analysis** - *An extremely useful tool to help you solve mathematical problems. It is based on the fact that when doing calculations involving measured quantities, the **units** must be added, subtracted, divided, or multiplied just like the **value** of the measurements.*

Dimensional Analysis

1.7 How many meters are in each of the following?

21 km

1023 570 μm

$$(\cancel{21 \text{ km}}) \left(\frac{1 \times 10^3 \text{ m}}{\cancel{\text{km}}} \right) = 21 \times 10^3 \text{ m} = \mathbf{2.1 \times 10^4 \text{ m}}$$

$$(\cancel{1023 \ 570 \ \mu\text{m}}) \left(\frac{1 \text{ m}}{\cancel{10^6 \ \mu\text{m}}} \right) = \mathbf{1.023 \ 570 \text{ m}}$$

Dimensional Analysis

1.8 How many mL are in 3.0 ft³?

$$1 \text{ ft} = 12 \text{ in} \quad 1 \text{ in} = 2.54 \text{ cm} \quad 1 \text{ cm}^3 = 1 \text{ mL}$$

$$= \underline{\hspace{2cm}} \text{ mL}$$

1.9 How many ns are in 23.8 s?

$$(23.8 \cancel{\text{ s}}) \left(\frac{10^9 \text{ ns}}{1 \cancel{\text{ s}}} \right) = 23.8 \times 10^9 \text{ ns} = 2.38 \times 10^{10} \text{ ns}$$

Properties of Matter

- **Extensive Properties - Properties that depend on the amount of matter present in a sample.**

Mass

Volume

Heat Capacity

- **Intensive Properties - Properties that do NOT depend on the amount of matter present in a sample.**

Color

Temperature

Density

Melting Point

Specific Heat

Boiling Point

Mass and Weight

Mass: the measure of the quantity or amount of matter in an object. The mass of an object does not change as its position changes.

Mass is measured using a ***BALANCE***.

Weight: A measure of the gravitational attraction of the earth for an object. The weight of an object changes with its distance from the center of the earth.

Weight is measured using ***SCALES***.

Sample Calculations Involving Masses

1.1 How many μg are in 2.56 kg?

= _____ μg

1.2 How many g are in 2.578×10^{12} ng?

= _____ g

Sample Calculations Involving Volumes

1.3 How many mL are in 3.456 L?

= _____ mL

1.4 How many μL are in 23.7 cm^3 ?

= _____ μL

Density

Density - *The mass of a unit volume of a material.*

$$\text{density} = \text{mass}/\text{volume}$$

1.5 What is the density of a cubic block of wood that is 2.4 cm on each side and has a mass of 9.57 g?

$$\text{volume} = [2.4 \text{ cm} \times 2.4 \text{ cm} \times 2.4 \text{ cm}]$$

$$\text{density} = (9.57 \text{ g})/(\text{_____ cm}^3)$$

$$= \text{_____ g/cm}^3 = \boxed{\text{_____ g/mL}}$$

Note that 1 cm³ = 1 mL

Page 20 in Textbook!

Temperature and Thermal Energy

Temperature: A measure of the “hotness” and “coldness” of an object; a measure of the average kinetic energy of the atoms and molecules of the object. The higher the temperature, the more kinetic energy the atoms and/or molecules have. This is an **INTENSIVE** property.

<http://www.uh.edu/engines/epi1764.htm>

Thermal Energy: Often called “heat”, it is the form of energy toward which all other forms tend to go. This is an **EXTENSIVE** property.

Sample Calculations Involving Temperatures

1.6 Convert 73.6°F to Celsius and Kelvin temperatures.

$$^{\circ}\text{C} = (5/9)(^{\circ}\text{F} - 32) \quad \text{K} = ^{\circ}\text{C} + 273.15$$

Memorize

$$\begin{aligned}^{\circ}\text{C} &= (5/9)(73.6^{\circ}\text{F} - 32) = (5/9)(41.6) \\ &= \mathbf{23.1^{\circ}\text{C}}\end{aligned}$$

$$\text{K} = 23.1^{\circ}\text{C} + 273.15 = \mathbf{296.3 \text{ K}}$$

Temperature Scales

*Page 17 in
Textbook*

Chapter 2 - Pages 44-81

Chapter 2

Atoms and Elements

Thompson
1897

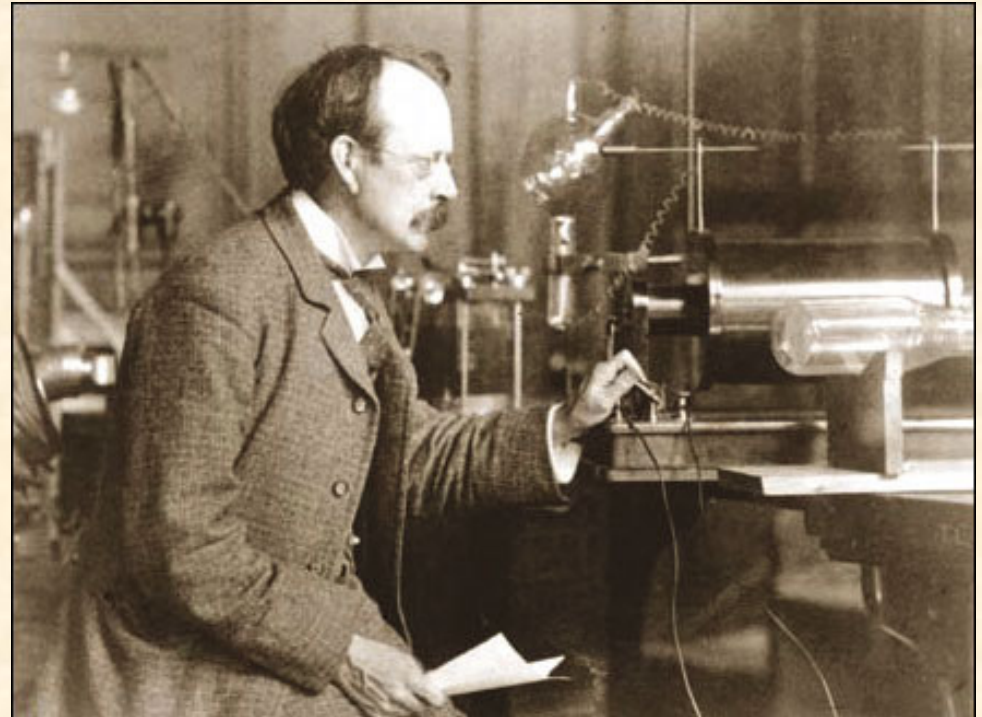
Millikan
1905

Rutherford
1911

John Dalton



J. J. Thompson



Not to be confused with this JJ:

J. J. Thompson's Cathode Ray Experiments

ATOMIC STRUCTURE

With the magnetic field present, the cathode ray is deflected out of the magnetic field. The stronger the magnetic field, the greater the amount of deflection.

$$e/m = E/H^2r$$

e = the charge on the electron

m = the mass of the electron

E = the electric field strength

H = the magnetic field strength

r = the radius of curvature of the electron beam

Thompson, thus, measured the charge/mass ratio of the electron - 1.759×10^8 C/g

ATOMIC STRUCTURE

Summary of Thompson's Findings:

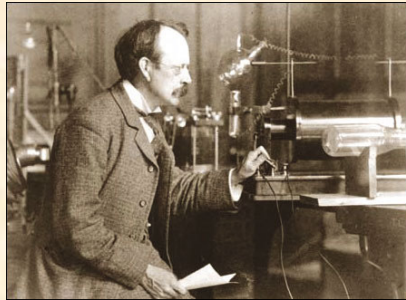
- ❑ Cathode rays had the same properties no matter what metal was being used.
- ❑ Cathode rays appeared to be a constituent of all matter and, thus, appeared to be a “sub-atomic” particle.
- ❑ Cathode rays had a negative charge.
- ❑ Cathode rays have a charge-to-mass ratio of 1.7588×10^8 C/g.

Ernest Rutherford

Metal Foil Experiment

ATOMIC

Ernest Rutherford - Developed the “**nuclear**” model of the atom.



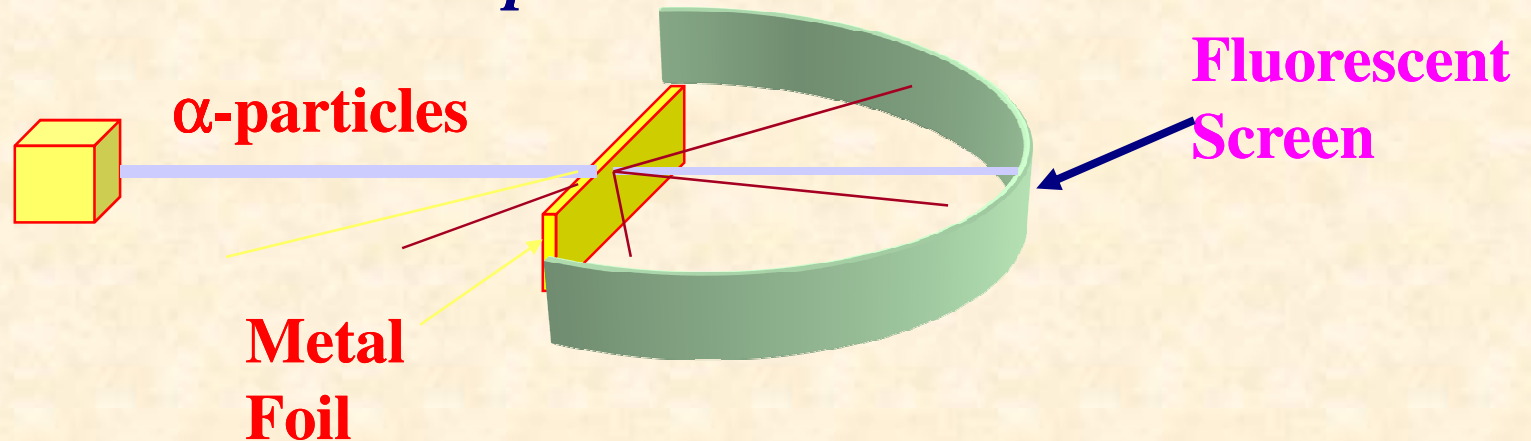
The Plum Pudding Model of the atom:

A smeared out “pudding” of positive charge with negative electron “plums” imbedded in it.

Electrons

The Metal Foil Experiments:

Radioactive Material in Pb box.



ATOMIC STRUCTURE

If the plum pudding model is correct, then all of the massive α -particles should pass right through without being deflected.

In fact, most of the α - particles DID pass right through. However, a few of them were deflected at high angles, disproving the “plum pudding” model.

Rutherford concluded from this that the atom consisted of a very dense nucleus containing all of the positive charge and most of the mass surrounded electrons that orbited around the nucleus much as the planets orbit around the sun.

ATOMIC STRUCTURE

Assignment:

Assume the diameter of the nucleus of a hydrogen atom is 1×10^{-13} cm and the diameter of the atom is 1×10^{-8} cm.

1. Calculate the volume of the nucleus and the volume of the atom in cm^3 .
2. Calculate the volume of empty space in the atom.
3. Calculate the ratio of the volume of the nucleus to volume of the whole atom.
4. Calculate the density of the nucleus if the proton's mass is 1.6726×10^{-24} g

ATOMIC STRUCTURE

Solution:

1. $V_{\text{nuc}} = (4/3)\pi(10^{-13} \text{ cm})^3 = (4/3)\pi(10^{-39} \text{ cm}^3) = \underline{4.2} \times 10^{-39} \text{ cm}^3$

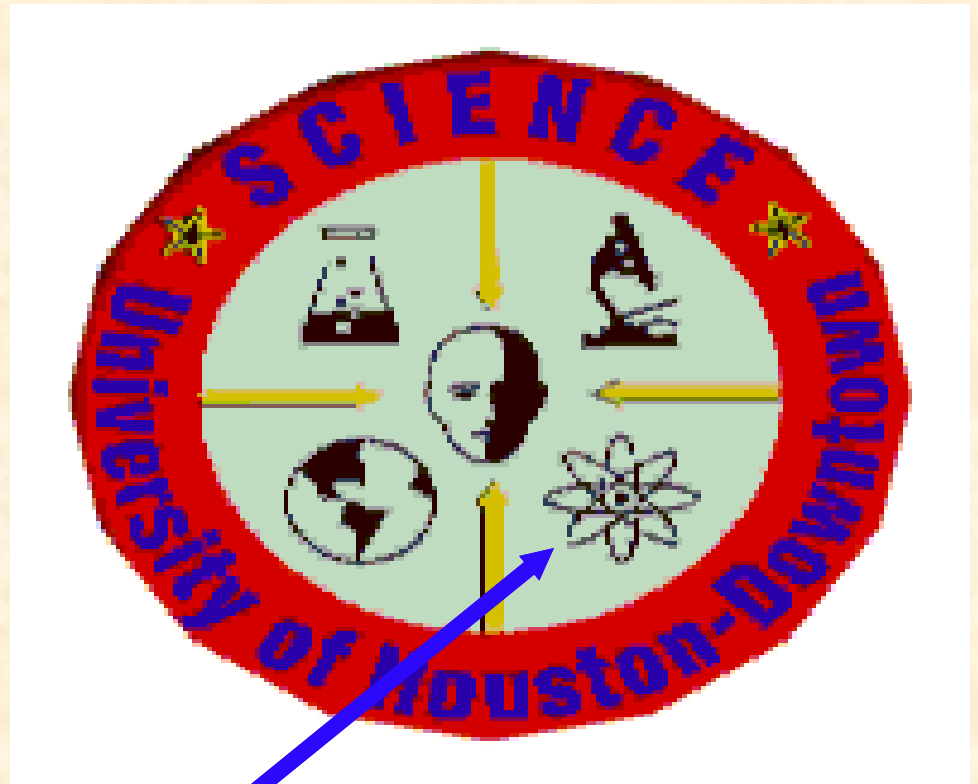
$$V_{\text{atom}} = (4/3)\pi(10^{-8} \text{ cm})^3 = (4/3)\pi(10^{-24} \text{ cm}^3) = \underline{4.2} \times 10^{-24} \text{ cm}^3$$

2. $\text{Vol}_{\text{atom}} - \text{Vol}_{\text{nuc}} = \underline{4.2} \times 10^{-24} \text{ cm}^3 !$

3. $V_{\text{nuc}}/V_{\text{atom}} = (\underline{4.2} \times 10^{-39} \text{ cm}^3)/(\underline{4.2} \times 10^{-24} \text{ cm}^3) = \underline{1} \times 10^{-15}$

4. $D = \text{mass/volume} = (1.6726 \times 10^{-24} \text{ g})/(\underline{4.2} \times 10^{-39} \text{ cm}^3) =$
 $\underline{4} \times 10^{14} \text{ g/cm}^3$

ATOMIC STRUCTURE



The Periodic Table

*Learn how to read
the Periodic Table*

The Periodic Table

The Periodic Law: The properties of the elements are a periodic function of their atomic number (Z).

The Structure of the Table:

1. Groups or Families - The vertical columns in the periodic table containing elements of similar properties.
2. Periods or Rows - The horizontal rows containing elements with continuously increasing atomic numbers. From left to right, they become more non-metallic.

The Periodic Table

3. Metals - Elements to the *left* and *down* in the periodic table. The majority of elements are metals.

Na Ca Mn Zn Bi Os

4. Metalloids (Semi-metals) -

Elements that border the line between metals and non-metals.

Si Ge As Sb

5. Non-metals - Elements to the *right* and *up* in the periodic table.

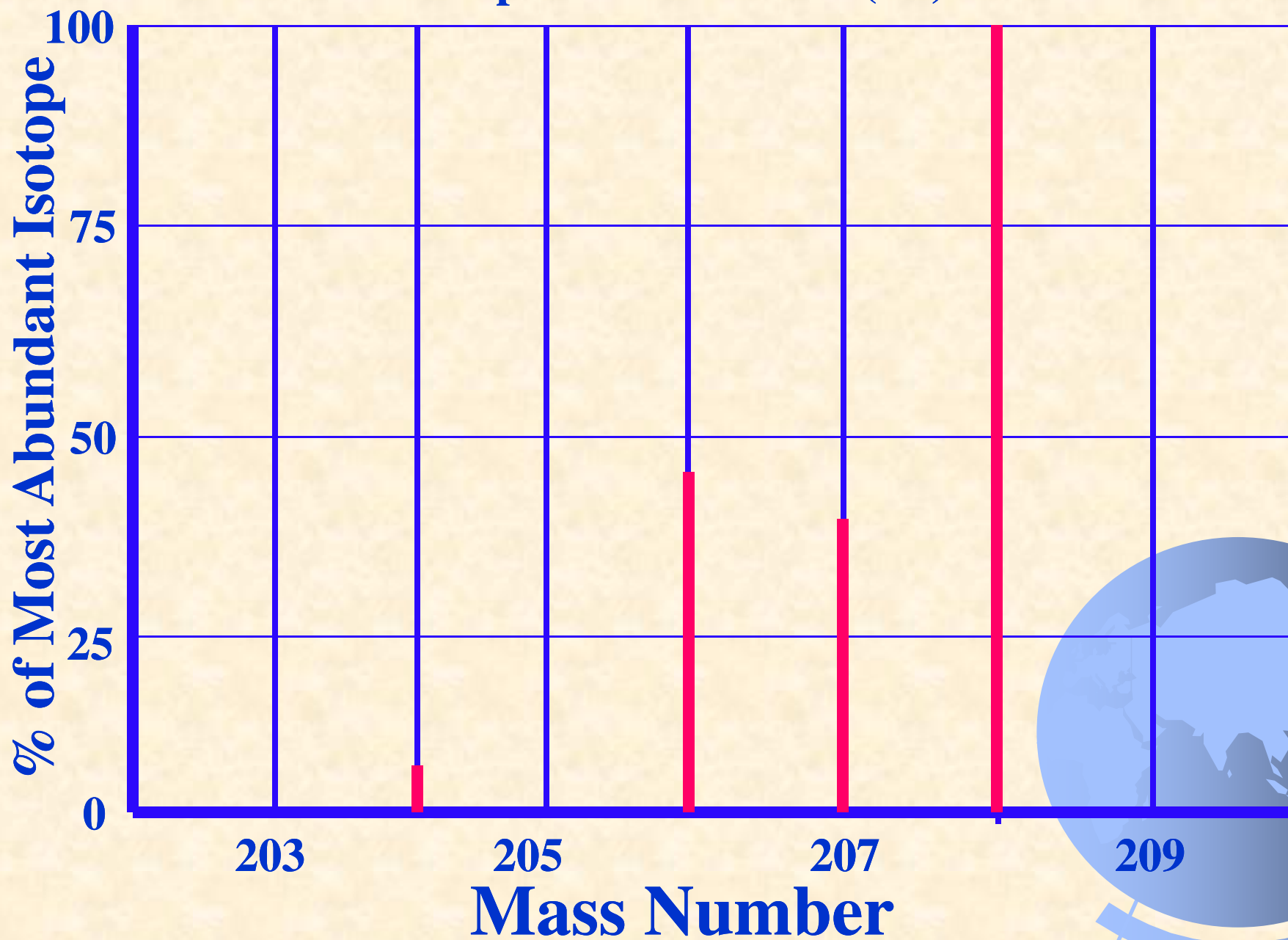
H C O Cl Se Br P Ne

How do we measure the atomic masses and the fractional abundances of isotopes?

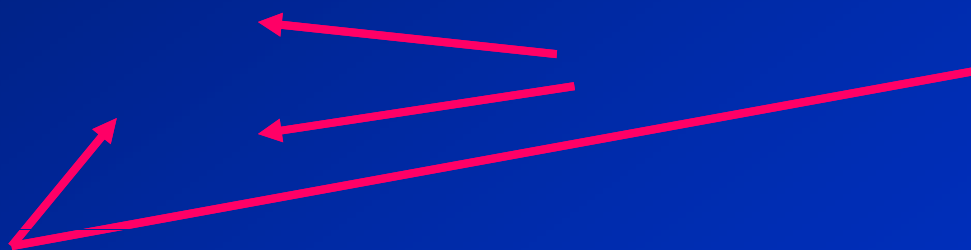
Mass Spectrometer

Isotopes

Mass Spectrum for Lead (Pb)



Molecular Concepts:



Atomic Masses:

Definition - The relative mass of an atom usually expressed in relative **atomic mass units (amu or u)**.

Atomic Mass Unit - A unit of mass that is equivalent to 1/12 of the mass of a single carbon-12 (^{12}C) atom.

A single carbon-12 atom has a mass of **EXACTLY** 12 u.

A single oxygen-16 atom has a mass that is 1.332 9096 times that of a carbon-12 atom.

$$(12 \text{ u})(1.332 \ 9096) = 15.994 \ 915 \text{ u}$$

The Atomic Masses of Elements:

The atomic masses shown on the Periodic Table are weighted average masses of naturally occurring samples of the elements. These samples consist of mixtures of **isotopes** with different atomic masses.

Isotope	Mass (u)	Fractional Abundance	Atomic Mass (u)
^{16}O	15.994 915	$\times 0.997\ 587$	$= 15.956\ 31$
^{17}O	16.999 133	$\times 0.000\ 374$	$= 0.006\ 358$
^{18}O	17.999 16	$\times 0.002\ 039$	$= 0.036\ 700$

Weighted Average = 15.9994 u

Practice Problem: Given the following data from the mass spectrometer, calculate the average atomic mass for the element, silicon (Si).

^{28}Si - 92.23 % (Atomic mass - 27.976 93 u)

^{29}Si - 4.67 % (Atomic mass - 28.976 49 u)

^{30}Si - 3.10 % (Atomic mass - 29.973 76 u)

$$\frac{(27.976\ 93\ \cancel{\text{u}})(\cancel{92.23\ \text{u}})}{(100\ \cancel{\text{u}})} + \frac{(28.976\ 49\ \cancel{\text{u}})(\cancel{4.67\ \text{u}})}{(100\ \cancel{\text{u}})}$$

$$+ \frac{(29.973\ 76\ \cancel{\text{u}})(\cancel{3.10\ \text{u}})}{(100\ \cancel{\text{u}})} =$$

28.09 u

Formula Masses: The relative mass of a compound is equal to the sum of the atomic masses of the elements forming the compound.

Calculate the formula masses for NaCl, H₂O, and H₃PO₄ -



$$1 \text{ Na @ } 22.989 \text{ 77 u} = 22.989 \text{ 77 u}$$

$$1 \text{ Cl @ } 35.453 \text{ u} = 35.453 \text{ u}$$

$$1 \text{ NaCl} = 58.443 \text{ u}$$



$$2 \text{ H @ } 1.007 \text{ 94 u} = 2.015 \text{ 88 u}$$

$$1 \text{ O @ } 15.9994 \text{ u} = 15.9994 \text{ u}$$

$$1 \text{ H}_2\text{O} = 18.0153 \text{ u}$$



$$3 \text{ H @ } 1.007 \text{ 94 u} = 3.023 \text{ 82 u}$$

$$1 \text{ P @ } 30.973 \text{ 76 u} = 30.973 \text{ 76 u}$$

$$4 \text{ O @ } 15.9994 \text{ u} = 63.9976 \text{ u}$$

$$1 \text{ H}_3\text{PO}_4 = 97.9952 \text{ u}$$

Pair: A group containing **2** items or objects.

Dozen: A group containing **12** items or objects.

Gross: A group containing **144** items or objects.

Mole: A group containing **6.0221×10^{23}** items or objects.

One Mole = the amount of substance (***element or compound***) that contains the same number of particles as there are atoms in **EXACTLY** 12 g of ^{12}C .

$6.0221 \times 10^{23} = \text{Avogadro's Number}$

Concept!

Page 72

“When in doubt, calculate moles!!”

of Particles

Moles

Avogadro's Number

How many atoms of carbon are in 2.6 moles of C?

$$\frac{(2.6 \text{ mol C}) (6.0221 \times 10^{23} \text{ atom C})}{\text{mol C}} = 1.6 \times 10^{24} \text{ atom C}$$

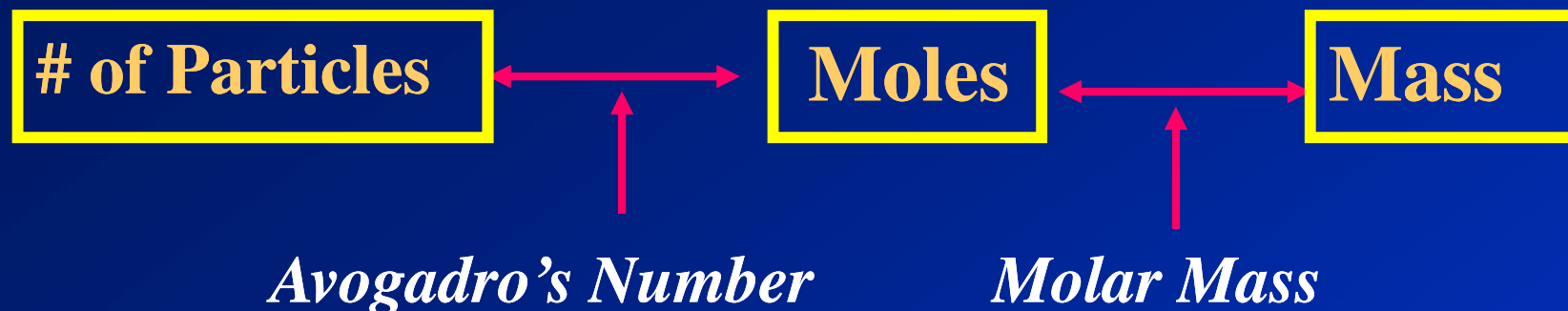
How many moles of water are present when 2.673×10^{25} molecules of water are present?

$$= \frac{(2.673 \times 10^{25} \text{ molecules H}_2\text{O})}{(6.0221 \times 10^{23} \text{ molecules H}_2\text{O})} (1 \text{ mol H}_2\text{O}) = 44.39 \text{ mol H}_2\text{O}$$

Molar Mass: The mass in grams of one mole of any substance. The molar mass is numerically equal to the formula mass.

Substance	Formula Mass	Molar Mass
H₂	2.015 88 u	2.015 88 g
Fe	55.847 u	55. 847 g
H₃PO₄	97.9952 u	97.9952 g
SO₃	80.06 u	80.06 g
NaCl	58.443 u	58.443 g

“When in doubt, calculate moles!!”



How many moles of carbon are in 26.78 g of C?

$$\frac{(26.78 \text{ g C}) \cancel{(1 \text{ mol C})}}{\cancel{(12.011 \text{ g C})}} = \mathbf{2.230 \text{ mol C}}$$

How many grams of water are in 32.6 moles of water?

$$\frac{(32.6 \text{ mol H}_2\text{O}) \cancel{(18.015 \text{ g H}_2\text{O})}}{\cancel{(1 \text{ mol H}_2\text{O})}} = \mathbf{587 \text{ g H}_2\text{O}}$$

How many atoms of hydrogen are present in 36.5 g of water?

“When in doubt, calculate moles!!”

$$\frac{(36.5 \text{ g H}_2\text{O})(1 \text{ mol H}_2\text{O})}{(18.015 \text{ g H}_2\text{O})} \frac{(2 \text{ mol H})}{(\text{mol H}_2\text{O})} \frac{(6.0221 \times 10^{23} \text{ atom H})}{(\text{mol H})}$$

= **2.44 x 10²⁴ atom H**

How many grams of Au atoms are present in 1.45 x 10²⁰ Au atoms?

$$\frac{(1.45 \times 10^{20} \text{ atom Au})(1 \text{ mol Au})}{(6.0221 \times 10^{23} \text{ atom Au})} \frac{(196.967 \text{ g Au})}{(\text{mol Au})}$$

= **0.0474 g Au**

How many moles of uranium are there in 1.23 kg U?

Molar Mass of U = 238.019 g/mol

$$\frac{(1.23 \text{ kg U})}{(1 \text{ kg U})} \frac{(10^3 \text{ g U})}{(238.019 \text{ g U})} (1 \text{ mol U}) = 5.17 \text{ mol U}$$

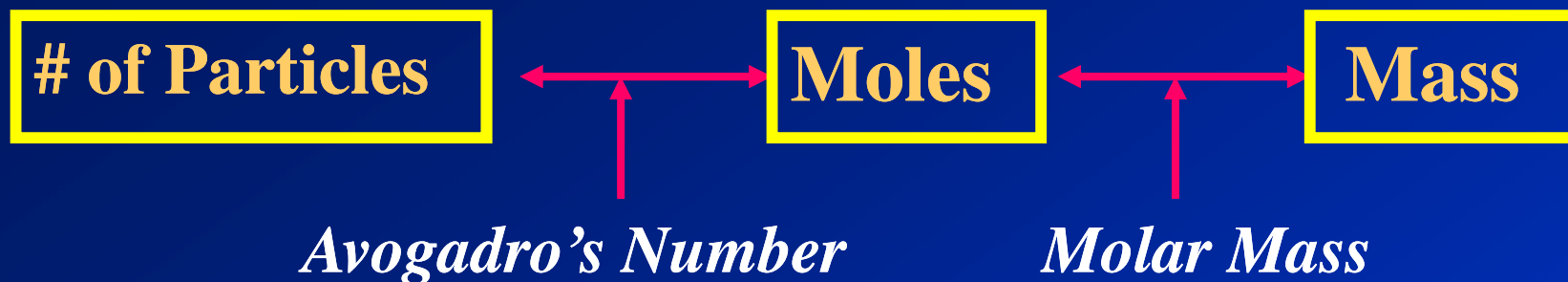
How many atoms of sulfur are there in 7.62 g S?

$$(7.62 \text{ g S}) \frac{(1 \text{ mol S})}{(32.066 \text{ g S})} (6.0221 \times 10^{23} \text{ atom S/mol S}) = 1.43 \times 10^{23} \text{ atom S}$$

What is the mass of a single Pt atom?

$$(1 \text{ Pt atom}) \frac{(1 \text{ mol Pt})}{(6.0221 \times 10^{23} \text{ atom Pt})} \frac{(195.08 \text{ g Pt})}{(1 \text{ mol Pt})} = 3.2394 \times 10^{-22} \text{ g Pt}$$

“When in doubt, calculate moles!!”



This allows you to do all types of calculations involving the conversion of particles to mass, mass to particles, mass to moles, moles to mass, particles to moles, or moles to particles.

If the world contains 6.0×10^9 people, how many moles of people are on earth?

$$\frac{(6.0 \times 10^9 \text{ people})(1 \text{ mol people})}{(6.0221 \times 10^{23} \text{ people})} = 1.0 \times 10^{-14} \text{ mol people}$$

If the earth has a radius of 4×10^3 miles what is its surface area in ft^2 ? If each person was allowed one ft^2 of space on earth, how many moles of people would fit on the earth's surface (counting the oceans)? Assume that the earth is a perfect sphere.

$$\text{Area of Sphere} = 4\pi r^2 \qquad 1 \text{ mile} = 5280 \text{ ft}$$

$$4\pi(4000 \text{ mi})(4000 \text{ mi})(5280 \text{ ft})(5280 \text{ ft}) \qquad = 5.6 \times 10^{15} \text{ ft}^2$$

$\frac{\text{mi} \quad \text{mi}}{\text{mi} \quad \text{mi}}$

$$\frac{(5.6 \times 10^{15} \text{ ft}^2)}{\text{ft}^2} \frac{(1 \text{ person})}{\text{ft}^2} \frac{(1 \text{ mol people})}{(6.0221 \times 10^{23} \text{ persons})}$$

$$= 9 \times 10^{-9} \text{ mol people}$$

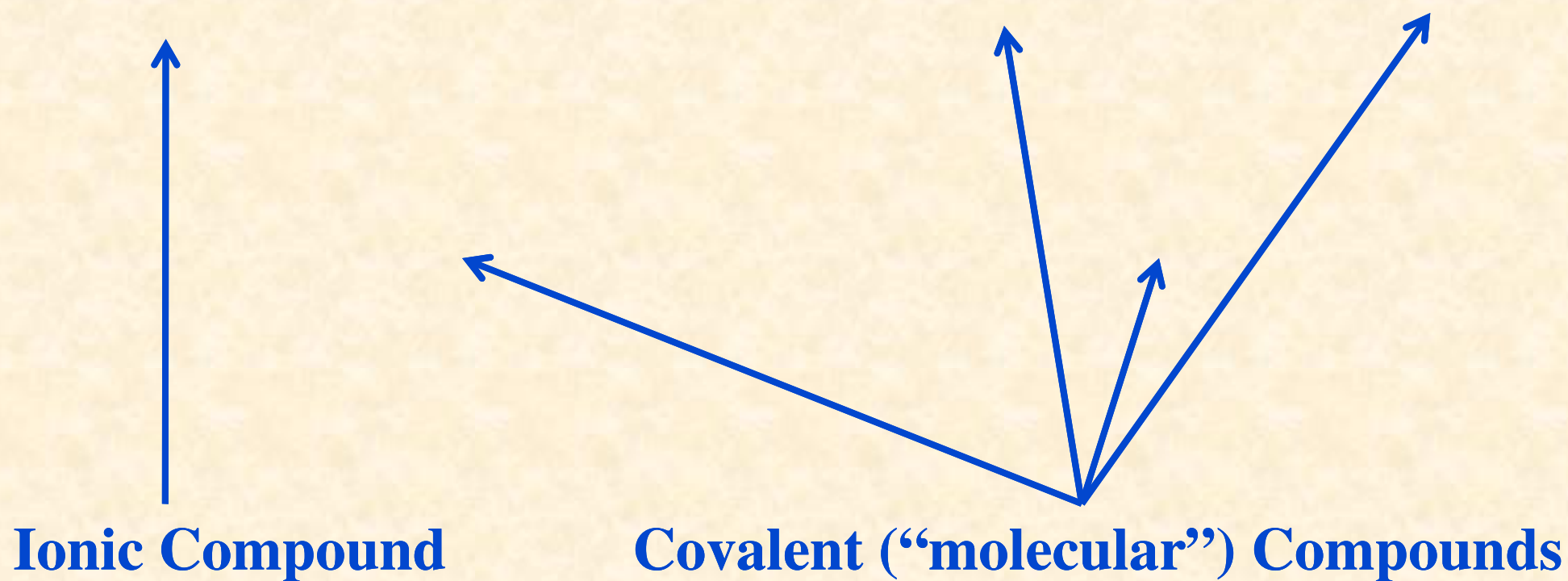
$$= 0.000 \ 000 \ 009 \text{ mol people}$$

Chapter 3

Molecules, Compounds, and Chemical Equations

Chemical Bonds

Covalent Bonds



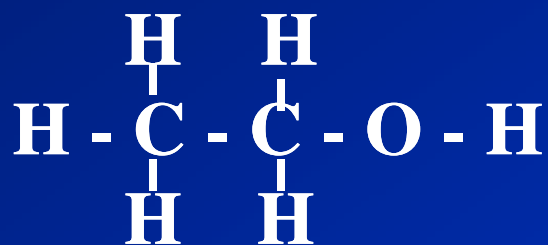
Compound Formulas

Empirical Formula - The formula of a compound that gives the simplest whole number ratio of the atoms in the compound.

Molecular Formula - The formula that gives the actual number of atoms of each element in a molecule of a *molecular compound*.

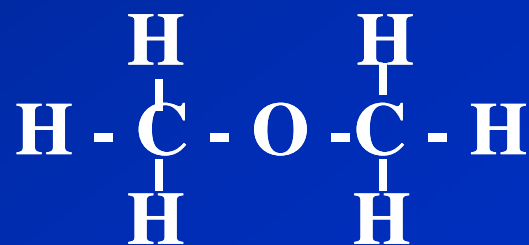
Compound	Empirical Formula	Molecular Formula
Benzene	CH	C ₆ H ₆
Water	H ₂ O	H ₂ O
<i>n</i> -Hexane	C ₃ H ₇	C ₆ H ₁₄
Hydrogen Peroxide	HO	H ₂ O ₂
Sodium Chloride	NaCl	-----

Structural Formula - The formula of a molecular compound that shows the way the atoms in the molecules are joined together.



ethyl alcohol

(ethanol)



dimethyl ether

(CH₃OCH₃)



“Condensed Structural Formulas”

Atomic and Molecular Concepts

“Space-Filling” Models

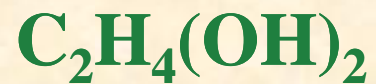
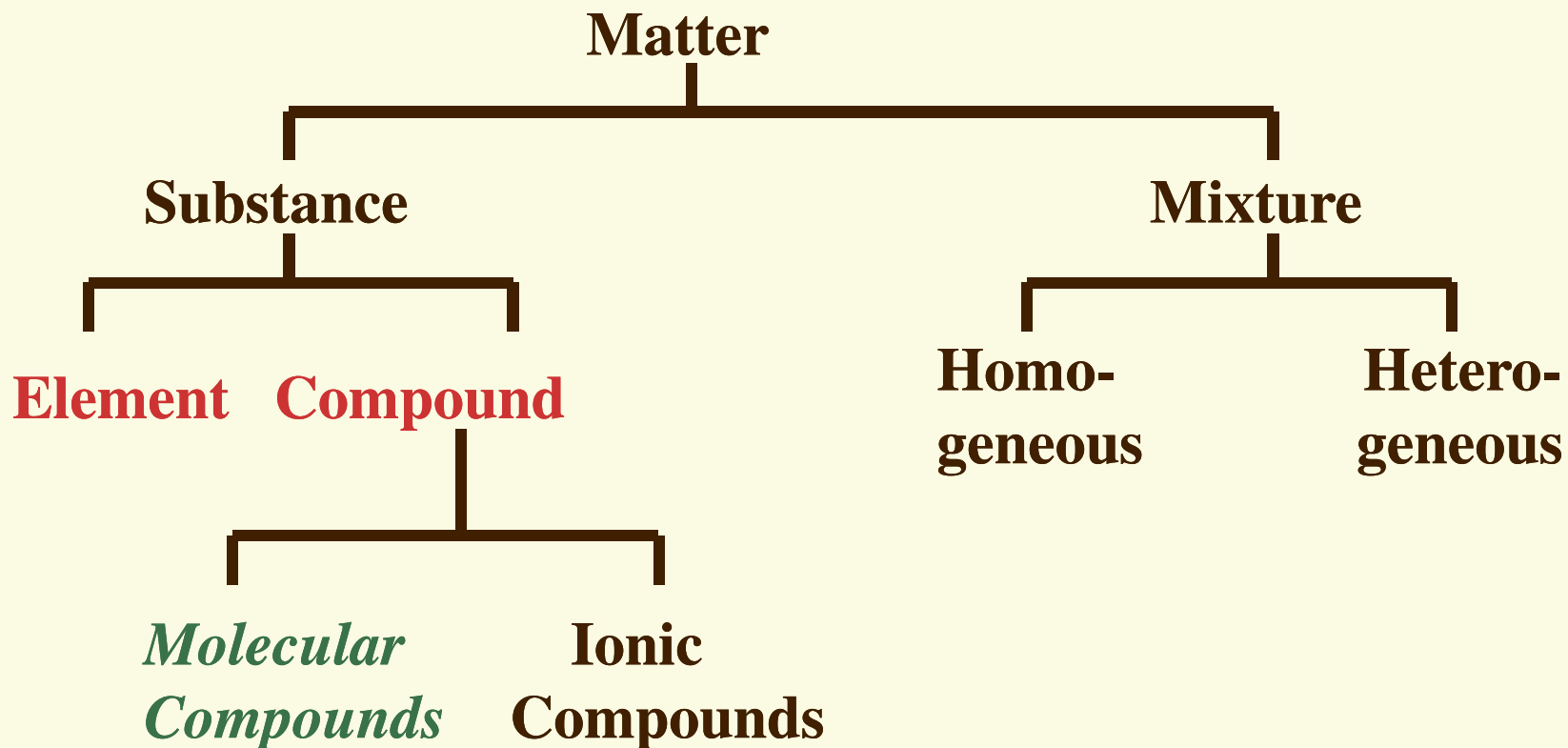


Figure 3.3

Molecular Compounds

Chemistry is a language that you must be able to read, write, and speak to be successful in this course.



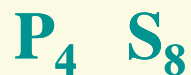
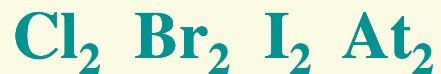
Memorize this chart!!!

Molecular Compounds

Atom - The smallest particle of an element that retains the properties of that element.

Molecule - An assembly of two or more atoms that are chemically bound together. It is the smallest particle of an element or molecular compound that retains the properties of that element or compound.

Elements



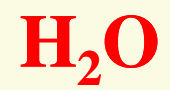
Compounds



Molecular Compounds

Molecular Compound - A compound that is composed of discrete molecules. *Normally*, they contain non-metal atoms chemically combined with one another.





Molecular Compounds

Nomenclature (Naming):

Binary Molecular Compounds -

1. Contain only *TWO* elements.
2. Usually contain only non-metals.
3. The element further to the left in the Periodic Table is usually written first.



hydrogen

chloride



hydrogen

sulfide



phosphorus

pentachloride



sulfur

dioxide



antimony

pentafluoride



dinitrogen

pentoxide

Molecular Compounds

6. Some molecular compounds have common names.



water



ammonia



phosphine

7. Organic compounds have their own naming system.



methane



ethane



propane



n-butane



ethyl alcohol

Molecular Compounds

4. For elements in the same family, the lower one in the Periodic Table is written first.
-

SO₂ - sulfur dioxide

BrCl - bromine chloride

SiC₂ - silicon carbide

5. When two elements form more than one compound, use prefixes.

CO - carbon monoxide

CO₂ - carbon dioxide

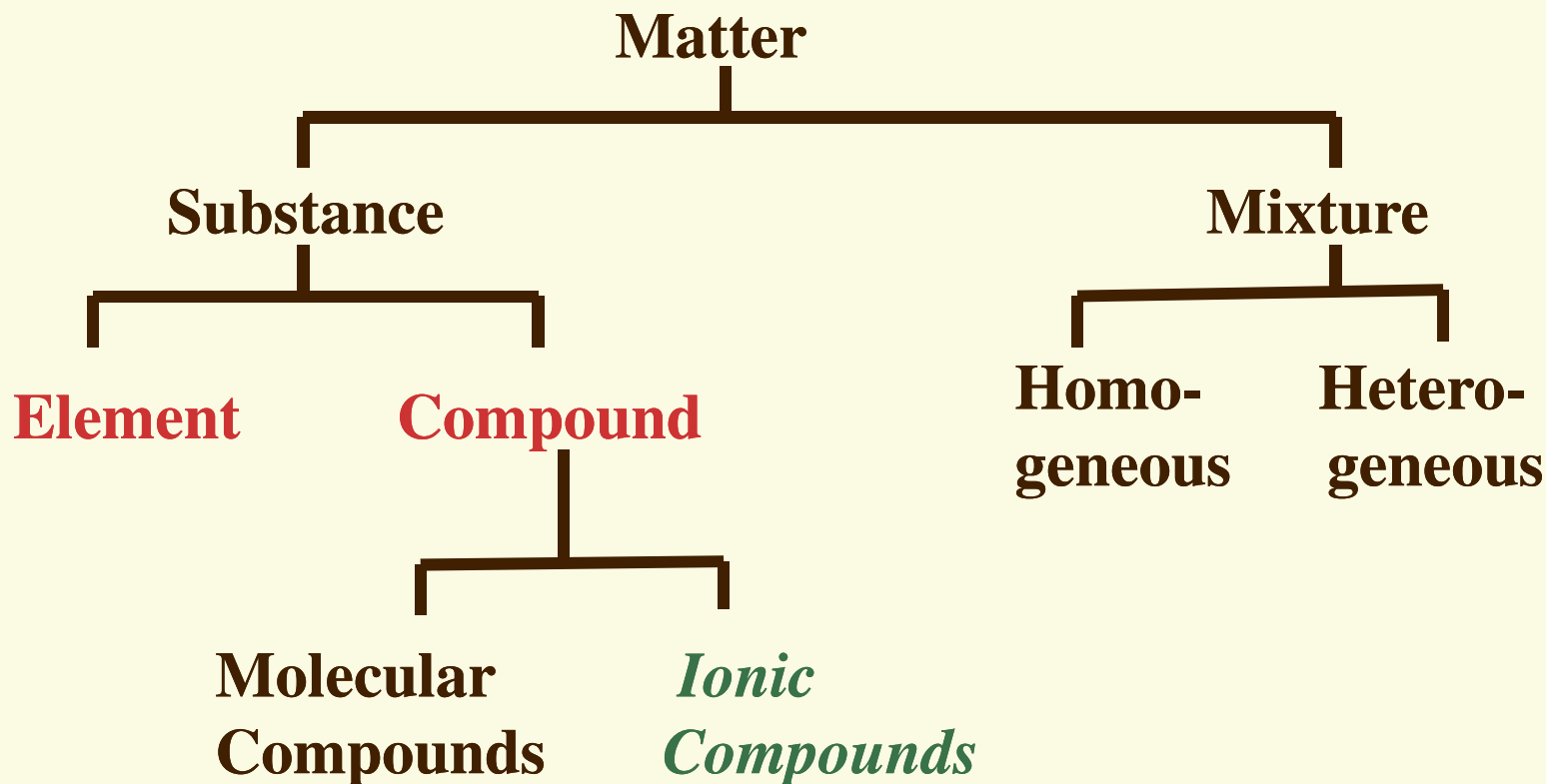
NO₂ - nitrogen dioxide

Molecular Compounds (*Organics*)

Molecular Compounds (*Organics*)

Ionic Compounds

Ion - An atom or group of atoms carrying an excess positive or negative charge.

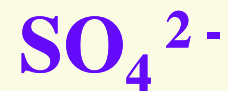


Remember!!!
Memorize this chart!!!

Ionic Compounds

Monatomic Ions - Ions that contain only one atom.

Polyatomic Ions - Ions that contain more than one atom.



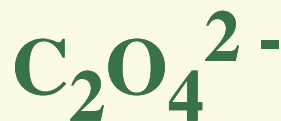
Ionic Compounds

Class Problems:

Classify each of the following as either a molecule or an ion:



Classify each of the following as either anions or cations:



Did you write down the names of these ions??

*If not, **WHY NOT???***

Ionic Compounds

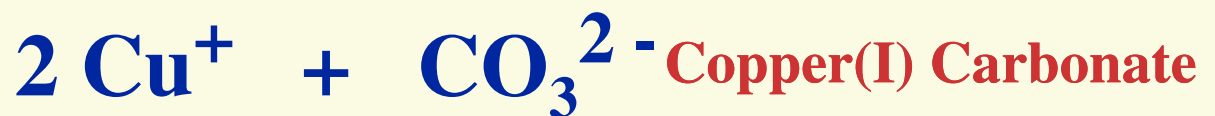
Ionic Compounds – Compounds that have *cations* and *anions* as their smallest particles. The charges on the ions must add to zero, i.e., the compounds are neutral.



Sodium Chloride



Calcium Fluoride



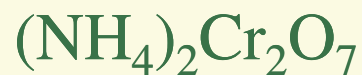
Copper(I) Carbonate

Ionic Compounds

Nomenclature (Naming):

Formulas for Ionic Compounds

Ammonium Dichromate:



Barium Phosphate: $\text{Ba}^{2+} + \text{PO}_4^{3-}$



Magnesium Nitride: $\text{Mg}^{2+} + \text{N}^{3-}$



Ionic Compounds

Ionic Compounds - Name the cation first, and then the anion.

Cations:

1. From Metal Atoms

Na^+ Sodium ion Ca^{2+} Calcium ion

Zn^{2+} Zinc ion Fe^{3+} Iron(III) ion

Fe^{2+} Iron(II) ion

2. From Non-metal Atoms

NH_4^+ Ammonium ion

H_3O^+ Hydronium ion

H^+ Hydrogen ion - Called a “proton”

WHY???

Ionic Compounds

Anions:

1. From *Simple Anions*

I^- - Iodide

O^{2-} - Oxide

S^{2-} - Sulfide

H^- - Hydride

N^{3-} - Nitride

Te^{2-} - Telluride

2. For Polyatomic Anions

NO_3^- - Nitrate

NO_2^- - Nitrite

SO_4^{2-} - Sulfate

SO_3^{2-} - Sulfite

CO_3^{2-} - Carbonate

$C_2H_3O_2^-$ - Acetate

ClO_4^- - Perchlorate

ClO_2^- - chlorite

ClO_3^- - chlorate

ClO^- - Hypochlorite

Ionic Compounds

3. Other Special Cases

OH^- **Hydroxide**

CN^- **Cyanide**

O_2^{2-} **Peroxide**

Class Problems:

Name the following ionic compounds:

$(\text{NH}_4)_2\text{CrO}_4$

$\text{Fe}_2(\text{SO}_4)_3$

**Ammonium
Chromate**

**Iron(III)
Sulfate**

$\text{Fe}_3(\text{PO}_4)_2$
Iron(II) Phosphate

Cu_2O
Copper(I) Oxide

$\text{KC}_2\text{H}_3\text{O}_2$
**Potassium
Acetate**

CaCr_2O_7
**Calcium
Dichromate**

Ions to Memorize!

pp. 94, 95 in Tro Textbook!

Composition Stoichiometry

Percent Composition: The percent by mass of each element in a compound. It can be determined from the compound's formula.

Calculate the percent composition of ethyl alcohol (ethanol), C_2H_6O or CH_3CH_2OH .

1. Calculate the molar mass of C_2H_6O :

$$2 \text{ C @ } 12.011 \text{ u} = 24.022 \text{ u}$$

$$6 \text{ H @ } 1.007 \text{ 94 u} = 6.047 \text{ 64 u}$$

$$1 \text{ O @ } 15.9994 \text{ u} = 15.9994 \text{ u}$$

$$1 \text{ C}_2\text{H}_6\text{O} = 46.069 \text{ u}$$

2. Calculate the percent by mass of each element in the compound:

$$\frac{(24.022 \text{ g C})}{(46.069 \text{ g ethanol})} (100\%) = \boxed{52.144 \% \text{ C}}$$

$$\frac{(6.04764 \text{ g H})}{(46.069 \text{ g ethanol})} (100\%) = \boxed{13.127 \% \text{ H}}$$

$$\frac{(15.9994 \text{ g})}{(46.069 \text{ g ethanol})} (100\%) = \boxed{34.729 \% \text{ O}}$$

$$\text{Total} = 100.000 \%$$

Empirical Formulas from Percent Composition:

What is the empirical formula for a compound with the following percent composition?

C - 64.80 %

H - 13.60 %

O - 21.60 %

Assume a 100.00 g sample!

“When in doubt, calculate moles”

$$(64.80 \text{ g C})(1 \text{ mol C}/12.011 \text{ g C}) = 5.395 \text{ mol C}$$

$$(13.60 \text{ g H})(1 \text{ mol H}/1.00794 \text{ g H}) = 13.49 \text{ mol H}$$

$$(21.60 \text{ g O})(1 \text{ mol O}/15.9994 \text{ g O}) = 1.350 \text{ mol O}$$



Find the smallest whole number ratio of subscripts!



Molecular Formulas from Empirical Formulas:

The molecular mass of $\text{C}_4\text{H}_{10}\text{O}$ was found by experiment to be 148.25 u. What is its molecular formula?

*The **Formula Mass** of the compound is:*

$$\begin{array}{r} 4 \text{ C @ } 12.011 \text{ u} = 48.044 \text{ u} \\ 10 \text{ H @ } 1.007 \text{ u} = 10.0794 \text{ u} \\ 1 \text{ O @ } 15.9994 \text{ u} = 15.9994 \text{ u} \\ \hline 1 \text{ C}_4\text{H}_{10}\text{O} = 74.1228 \text{ u} \end{array}$$

$$n(74.1228 \text{ u}) = 148.25 \text{ u} \quad \therefore n = 2$$



The percent composition of a compound can also be determined by experiment. This type of experiment is known as *quantitative analysis*.

An oxide of mercury (Hg) is decomposed to oxygen gas and mercury. If a 0.8349 g sample of the oxide gives 0.7732 g of Hg, find the percent by mass of Hg and oxygen in the compound.

$$\left[\frac{0.7732 \text{ g Hg}}{0.8349 \text{ g sample}} \right] [100 \%] = 92.61 \% \text{ Hg}$$

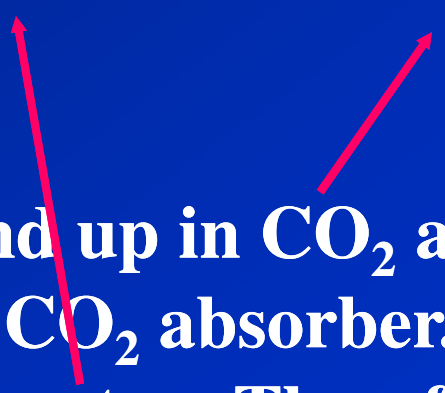
$$100 \% \text{ total} - 92.61 \% \text{ Hg} = 7.39 \% \text{ O}$$

or

$$\left[\frac{(0.8349 \text{ g total} - 0.7732 \text{ g Hg})}{(0.8349 \text{ g total})} \right] [100 \%] = 7.39 \% \text{ O}$$

How Combustion Reactions are Done for Hydrocarbon Analyses:

All of the Carbon atoms end up in CO₂ and are, therefore, captured by the CO₂ absorber. All of the hydrogen atoms end up in water. Therefore, they are captured by the water absorber.



Practice Problem:

A hydrocarbon compound is analyzed by burning it with an excess of oxygen gas. In this analysis, 4.80 g of H₂O were produced and 8.79 g CO₂. What was the empirical formula of the hydrocarbon?

1. First recognize that all of the carbon from the hydrocarbon ends up in the CO₂ and all of the hydrogen ends up in the water.
2. Calculate the moles of hydrogen in water and the moles of carbon in CO₂:

$$\begin{aligned} & \cancel{(4.80 \text{ g H}_2\text{O})} \cancel{(1 \text{ mol H}_2\text{O}/18.0153 \text{ g H}_2\text{O})} \cancel{(2 \text{ mol H/mol H}_2\text{O})} \\ & = \underline{0.5329} \text{ mol H} \end{aligned}$$

$$\cancel{(8.79 \text{ g CO}_2)} \cancel{(1 \text{ mol CO}_2 / 44.010 \text{ g CO}_2)} \cancel{(1 \text{ mol C} / \text{mol CO}_2)} \\ = 0.1997 \text{ mol C}$$

3. Calculate the empirical formula of the hydrocarbon:

$$\frac{\text{C}_{0.1997} \text{H}_{0.5329}}{0.1997 \quad 0.1997} = (\text{CH}_{2.67}) \times 3 = \boxed{\text{C}_3\text{H}_8}$$

The initial result was a fraction. Since a molecule cannot have a partial atom, the formula must be multiplied by a factor that will make the subscripts whole numbers.

A 0.018 69 g sample of an organic compound containing only C, H, and O gave 0.044 38 g of CO₂ and 0.022 72 g of H₂O on combustion. What is the percent composition of the compound?

First, recognize that ALL of the carbon in the compound ends up in CO₂ and ALL of the hydrogen ends up in the water.

$$1 \text{ C @ } 12.011 \text{ u} = 12.011 \text{ u}$$

$$2 \text{ O @ } 15.9994 \text{ u} = 31.9988 \text{ u}$$

$$1 \text{ CO}_2 = 44.010 \text{ u}$$

$$\frac{(0.044 \text{ 38 g CO}_2)(1 \text{ mol CO}_2) (1 \text{ mol C}) (12.011 \text{ g C})}{(44.010 \text{ g CO}_2)(\text{mol CO}_2) (1 \text{ mol C})}$$

$$= \boxed{0.012 \text{ 11}_2 \text{ g C}}$$

$$\frac{(\cancel{0.012112 \text{ g C}}) (\cancel{100\%})}{(\cancel{0.01869 \text{ g sample}})} = \boxed{64.80\% \text{ C}}$$

$$\frac{(\cancel{0.02272 \text{ g H}_2\text{O}})(\cancel{1 \text{ mol H}_2\text{O}}) (\cancel{2 \text{ mol H}}) (\cancel{1.00794 \text{ g H}})}{(\cancel{18.015 \text{ g H}_2\text{O}}) (\cancel{\text{mol H}_2\text{O}}) (\cancel{\text{mol H}})}$$
$$= \boxed{0.0025424 \text{ g H}}$$

$$\frac{(\cancel{0.0025424 \text{ g H}}) (\cancel{100\%})}{(\cancel{0.01869 \text{ g sample}})} = \boxed{13.60\% \text{ H}}$$

$$100.00\% \text{ total} - (64.80\% \text{ C} + 13.60\% \text{ H}) =$$

$$\boxed{21.60\% \text{ O}}$$