Chemistry- study of the composition of matter and the changes that matter undergoes

Types of Chemistry: physical, organic, inorganic, biochemistry, analytical

Physical vs. Chemical Changes

Physical change- change that will alter a substance without changing its composition

Examples: tearing, crushing, melting, evaporating, dissolving

Chemical Change- new substances with different chemical compositions are produced

Examples: burning, rusting, cooking

Signs of chemical change: color change, change in mass, temp. change, gas released, solid forms (precipitate)

Forms of Matter

Matter

Pure Substances- definite compositions

Element- simplest form of matter- represented by symbols

Compound- composed of 2 or more elements chemically combined- represented by a formula

H_2O — CO_2

Mixtures- 2 or more pure substances <u>physically</u> combined

TYPES

Homogeneous- same composition throughout

$\Rightarrow$ solution

Heterogenous- different composition throughout

Chemical Reaction- occurs when one or more substances changes into one or more other substances

Ex.

56g 32g 88g

$Fe + S \rightarrow FeS$

Reactant(s) Product(s)

Law of Conservation of Mass- In any physical or chemical reaction, mass is conserved (matter is not created nor destroyed)

Measurement

Quantitative measurement- gives results in definite form, usually as numbers
Ex. boiling point of a substance

Qualitative measurement- gives results in descriptive, non-numeric form
Ex. determine the identity of a substance

Accuracy- How close the measurement comes to the actual dimension or true value- dependent on quality of measuring device

Precision- How reproducible the measurement is- Dependent on the scale of the device used- Number of decimal places the measurement can be taken to with certainty

Significant Figures

Meaningful digits in a measured or calculated quantity

Rules for Significant figures:

1. **Any nonzero digit (numbers 1-9) is significant**

   Ex. 24.7 3 **0.743** 3 **714** 3

2. **Zeroes between significant digits are always significant**

   Ex. 7003 4 **40.79** 4 **1.503** 4

4 3. **Zeroes to the left of the first nonzero digit (in decimals) are not significant**

   Ex. 0.0071 -2 **0.42** -2 **0.000099** -2

   7.1×10^{-3} 9.9×10^{-5}

~~4~~ 5. Zeroes at the end of a number and to the right of a decimal point are always significant

Ex. 43.00 1.010 9.000

4 4 4

0.0034000

~~5~~ 3. For numbers that do not have decimal points, trailing zeroes are not significant in most cases. If such zeroes are known measured values, they should be written in scientific notation

Ex. 300 7,000 27,210

-1 -4

1

3.00×10^2 -3

7.000×10^3

300. -3

7000.

0.020100

Significant figures in Calculations

Adding and Subtracting:
Answers cannot have more digits to the right of the decimal point than either of the original numbers

Ex. 12.52 meters
349.0 meters
\+ 8.24 meters

369.76
369.8 meters

74.626 meters
-28.34 meters

46.286
46.29

Multiplying and Dividing:
Answer must contain not more significant figures than the measurement with the *least number of total significant figures*

Ex.

7.55 meters X 0.34 meters = 2.5772

3 2 2.6

2.10 meters X 0.70 meters = 1.47

1.5

2.4526 meters^2 / 8.4 meters = 0.29

0.365 meters^2 / 0.0200 meters = 18.3

Significant Figures in conversion problems

When converting, the answer to the conversion problem should have the same number of sig. figs. as the measurement that is being converted. Conversion factors do not count because they are not actually measurements that were taken.

Convert 23.06 mg to g

Find the sum of the following in liters:

250.0 mL, 1.41. L, 1400 cm^3

$$250.0\ \text{mL} \times \frac{1\ \text{L}}{1000\ \text{mL}} = 0.2500\ \text{L}$$

$$1.41\ \text{L}$$

$$1400\ \cancel{\text{cm}^3}^{\ \text{mL}} \times \frac{1\ \text{L}}{1000\ \text{mL}} = +\ 1.4\ \text{L}$$

$$3.0600\ \text{L}$$

$$3.1\ \text{L}$$

CONVERSIONAND MEASUREMENT

DISTANCE

km - kilo	0.001		English Conversions
X			1 meter = 39.4 inches
X			1 km = 0.621 miles
m	1		1 cm = 0.394 inches
X			
cm - centi	100		
mm -milli	1000		
X		ALL	
X		ARE	
µm - micro	1 X 10^6	EQUAL	
X			
X			
nm - nano	1 X 10^9		
Å - Angstrom	1 X 10^{10}		

VOLUME

1 L = 1000 mL
1 L = 1 X 10^6 µL
1 mL = 1 cm^3 (cc) cubic centimeter

English conversion:
1 L = 1.06 qts.

MASS
1 kg = 1000 g
1000mg = 1 g

English converions:
1 kg = 2.2 lbs.
1 lb. = 454 g

TEMPERATURE
Kelvin (absolute) scale
^{o}Celsius + 273 = ^{o}K
0^{o}K = absolute zero - -273^{o}C

Convert to cm:

A. **36.5 $\cancel{km}$** $\times \frac{100\ cm}{0.001\ \cancel{km}} = 3,650,000cm$

B. **4.5 X 10^8 $\cancel{\mu m}$** $\times \frac{100\ cm}{1\times10^6\ \cancel{\mu m}} = 45,000cm$

C. **67 $\cancel{feet}$** $\times \frac{12\ \cancel{in}}{1\ \cancel{ft.}} \times \frac{1\ cm}{.394\ \cancel{in.}} = 2040.6cm$

Convert to mL:

A. 5.89 $\cancel{L} \times \dfrac{1000\,mL}{1\,\cancel{L}} = 5{,}890\,mL$

B. 78.5 $\cancel{cm^3} \times \dfrac{1\,mL}{1\,\cancel{cm^3}} = 78.5\,mL$

C. 65.4 ~~quarts~~ $\times \dfrac{1\,\cancel{L}}{1.06\,\cancel{qts.}} \times \dfrac{1000\,mL}{1\,\cancel{L}} = 61{,}698.1\,mL$

Convert to grams

A. 8.79 kg $\times \frac{1000g}{1kg} = 8790g$

B. 47.4 lbs. $\times \frac{454g}{1 lb} = 21{,}519.6g$

Density Problems

Calculate volume using significant figure rule for multiplication/division

$V = l \times w \times h$

Find the volume of a container with dimensions that measure 1.36 m x 2.45 m x 4.6 m. Find the volume in cm^3

$V = 15.3272\ m^3$

$15\ m^3$

$$15\ m^3 \times \frac{\overset{1,000,000}{(100\ cm)^3}}{(1\ m)^3} = 15,000,000\ cm^3$$

$1.5 \times 10^7\ cm^3$

Density- the ratio of the mass of an object to its volume

d = m/v where d=density, m=mass, v=volume

Ex. m= 9.60g **v_i=** 5.30mL **v_f=**6.15mL

$V = V_f - V_i = 0.85\,mL$

$d = m/V = \frac{9.60g}{0.85mL}$

$= 11.391\,mL$

Percent error- the ratio of an error to an accepted value

$$\text{Percent error} = \frac{\left|\overset{\text{theoretical}}{\text{accepted value}} - \overset{\text{actual}}{\text{experimental value}}\right|}{\text{accepted value}} \times 100$$

where the accepted value is the value that <u>should</u> be obtained in the experiment and the experimental value is the value that was <u>actually</u> obtained in the experiment

accepted = 11.34 g/mL

exper. = 11.3 g/mL

$$\%\text{error} = \frac{|11.34 - 11.3|}{11.34} \times 100 = 0.353\%$$

Calculate the volume of a metal that has a density of 10.1 g/cm^3 and a mass of 28 g

d m

$V = m/d = \frac{28\,\cancel{g}}{10.1\,\cancel{g}/cm^3} = 2.8\,cm^3$ $\quad \frac{1}{1/cm^3} = cm^3$

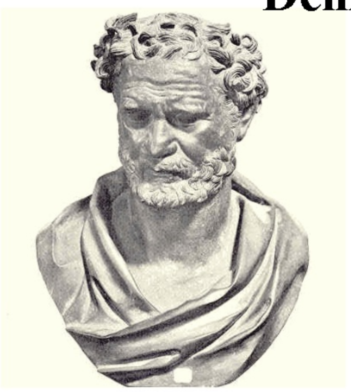
The density of ethanol is 0.798 g/mL. Calculate the mass of 17.4 mL of the liquid.

d V

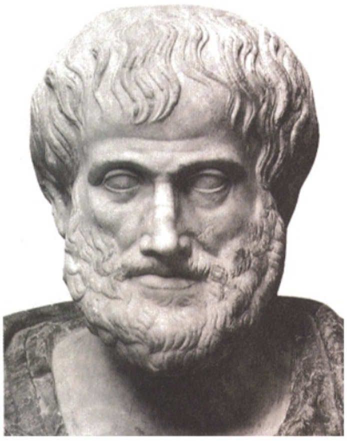
$m = d \times V = (.798\,g/\cancel{mL})(17.4\,\cancel{mL})$

$= 13.9\,g$

Democritus and Aristotle



4^{th} century B.C.-- The first suggestion of atoms was by Democritus (an ancient Greek). He suggested the idea of "indivisible particles", but these were not supported by experimental data

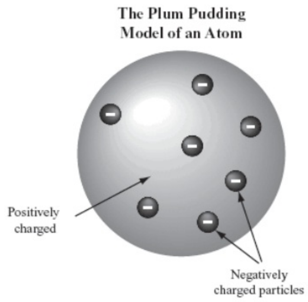


Aristotle believed that all matter could be classified into 4 types: air, earth, fire and water

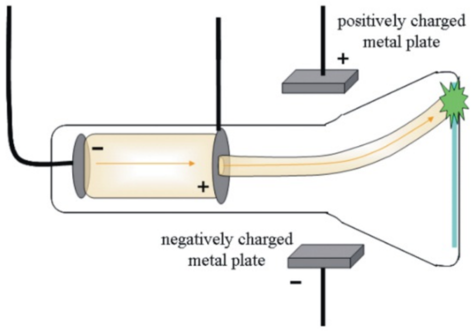
late 1700s—<u>John Dalton</u> (an English chemist and physicist) and other chemists began associating chemical changes to changes in individual atoms

Dalton's theory:

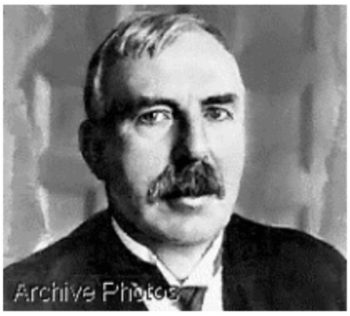
1. Matter consists of definite particles called atoms.
2. Atoms are indestructible.
3. The atoms of one particular element are all identical in mass and other properties.
4. In chemical reactions, the atoms rearrange but they do not themselves break apart.
5. When atoms of different elements combine to form compounds, new and more complex particles form. However, in a given compound the constituent atoms are always present in the same fixed numerical ratio.



- **1897—J.J. Thomson discovered the existence of the electron using evidence from cathode ray tubes. Prior to this, hydrogen was thought to be lightest form of matter (Thomson's evidence showed the electron to be almost 2000 times lighter)**

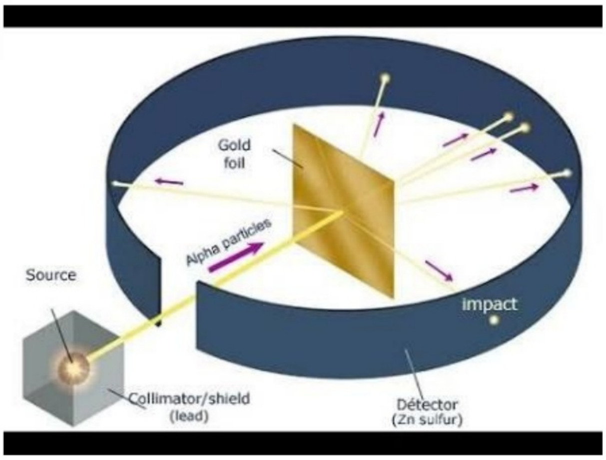


1923- Robert Millikan showed the exact charge of an electron with his oil drop experiment



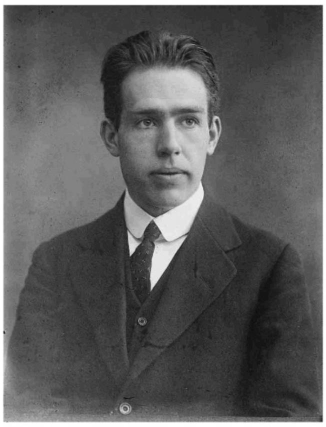
- **1911 – Ernest Rutherford provides the evidence for the positively charged nucleus and, therefore, the proton. Prior to Rutherford, it was thought that protons were distributed throughout the atom.**

http://chemmovies.unl.edu/ChemAnime/RUTHERFD/RUTHERFD.html



1886- Eugen Goldstein- found evidence for positively charged particles (protons)

1932- Sir James Chadwick, an English physicist, confirmed the existence of the neutron, a particle with no charge and a mass nearly equal to the proton.



1913- Niels Bohr

- Electrons in a particular path have fixed energy which prevents them from falling into the nucleus
- Energy levels are quantized (fixed)
- Energy levels are like the rungs of a ladder

Quantum Mechanical Model- 1926 (Schrodinger, Heisenberg, deBroglie)

- The quantum-mechanical approach acknowledges the wavelike character of electrons and provides the framework for viewing the electrons as fuzzy clouds of negative charge.
- Electrons still have assigned states of motion, but these states of motion do not correspond to fixed orbits. Rather than orbits, the locations of electrons are viewed as cloud-like arrangements

http://www.edumedia-sciences.com/en/a66-electron-distribution
http://www.physics.ucla.edu/~dauger/orbitals/

Summary of Subatomic particles

Properties of Subatomic Particles

Particle	Symbol	Relative Electrical Charge	Approximate Relative Mass (amu)	Actual Mass (kg)	Location
Electron	e^-	-1	1/1840	9.11×10^{-28}	elec. clouds
Proton	p^+	+1	1	1.67×10^{-24}	nucleus
Neutron	n^0	0	1	1.67×10^{-24}	nucleus

Each element is unique according to the number of protons, neutrons, and electrons in an atom

<u>ATOMIC NUMBER</u>- the number of protons in the nucleus of an atom

atom# = protons = electrons (if neutral)

For <u>neutral atoms</u>, the number of electrons in the atom would be the same as the atomic number, as well

<u>MASS NUMBER</u>- the total number of neutrons and protons present in the nucleus of an atom mass # - 9

atom. # - 4 Be

	Atomic number	Mass number	p^+	n^0	e^-	Symbol
Beryllium	4	9	4	5	4	^{9}Be
Neon	10	20	10	10	10	^{20}Ne
Sodium	11	23	11	12	11	^{23}Na

ISOTOPES- Atoms having the same atomic number but different mass numbers (due to different numbers of neutrons)

Two types of notation

^{12}C ^{14}C carbon-12 carbon-14

Example

	Atomic number	Mass number	p^+	n^0	e^-	Symbol
carbon-12	6	12	6	6	6	^{12}C
carbon-14	6	14	6	8	6	^{14}C