CHEM 115
Structure of Atoms

Lecture 2
Prof. Sevian

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Agenda

- A few announcements
  - First discussion section is tomorrow (Friday) – Attendance will be taken and Assignment #1 will be discussed
  - If you are waiting to get into this class and if you do not need to take Chem 117 (because you already passed it), please email me to tell me so, and also give me your UMS number and which discussion section you'd like to be enrolled in
- Calculation rules, through density calculations examples
  - Dimensional analysis
  - Significant figures
  - Scientific notation
- Structure of atoms & evidence
- Isotopes
- Facilitated study group – poll to find out optimal times

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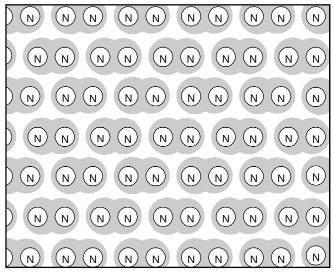
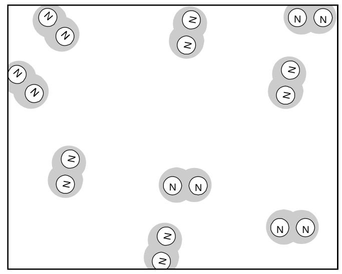
Density = A measure of concentration



N_2 (*s*)



N_2 (*g*)

Which is more dense, the solid or the gas?

Density measures how many particles (how much "stuff") are in a given volume (space). To compare two samples, you can either:

- **Compare equal volumes and then see which one has more particles (stuff) in that same volume (space)**
- **Compare equal amounts of particles (stuff) and see which one takes up more volume (space)**

Density consequences

1.00 cm^3 of liquid water has a mass of 1.00 g
1.00 cm^3 of ice has a mass of 0.92 g

What are the densities of each?
Liquid water density = 1.00 g / 1.00 cm^3 = 1.00 g/cm^3
Ice density = 0.92 g / 1.00 cm^3 = 0.92 g/cm^3

Which is less dense? Does this make sense?
Ice is less dense

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Key points about density

- Density is:
  - A measure of concentration (crowdedness) because it tells you how much stuff there is in a certain amount of space
  - Amount of matter per unit of volume
    - Amount of matter can be measured in grams (mass)
    - Volume can be measured in cubic distance (e.g., cm^3 or m^3) or space occupied (e.g., milliliters or liters). These are actually the same thing because 1 mL = 1 cm^3 by definition.
  - D = (mass) / (volume), most often as g/mL or g/L

Some Important Chemistry Skills

- Using dimensional analysis to solve problems

  Why is it useful?
  - Ensures correct units
  - Can find missing information
- Keeping track of significant digits in calculations

  Why is this important?
  - Measurements taken by instruments limit accuracy of information

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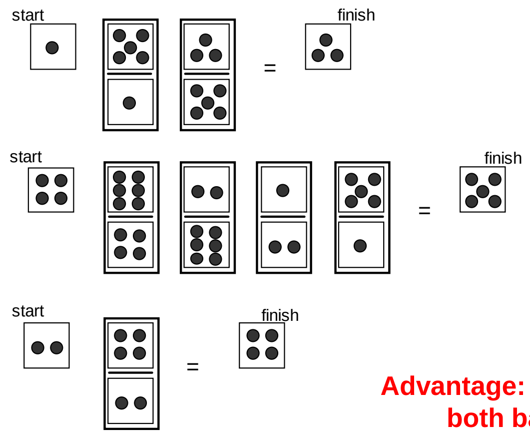
Dimensional Analysis

There is only one rule



Advantage: You can work both backwards and forwards to solve a problem

Dominoes image from Sevian et al, ***Active Chemistry*** (2006)

Example: Density Problem

Mercury has a density of 13.534 g/mL. What is the mass of 24 mL of mercury?

Units are grams

Start | End

$$\boxed{24\ \text{mL}} \times \frac{\boxed{13.534\ \text{g}}}{\boxed{1\ \text{mL}}} = \boxed{320\ \text{g}}$$

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Dimensional analysis

What to look for:

1. What data are given?
2. What quantity do you need?
3. What conversions are available to take you from start to end?

What you need to be able to do:

- Recognize dimensions by both names and units
  - Volume or (mL, L, cm^3; and mL is the same as cm^3)
  - Mass or (g, kg)
  - Density or (units of mass/units of volume)
- Determine starting and ending information
- Work toward the middle from both ends
- Do the calculations properly

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Another density example

A 2.0 cm x 2.0 cm x 3.0 cm rectangular piece of metal has a mass of 32.4 g. What is the metal's density?

Significant Digits

Why are they important?

- Tell you to what extent you can "trust" the data
- Tell you how reliable is the least reliable instrument that was used in determining the data

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What goes into any calculation?

Three kinds of information

1. Measured values
2. Exact values
3. Derived (calculated) values

The least reliable one determines the reliability of the final calculated result.

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Measured Values

Measurement	Units most often used	Instrument often used
Length (Wavelength)	m, cm nm	Ruler (spectrometer)
Mass	g, kg	Balance
Volume	L, mL, m^3, cm^3	Ruler (to measure dimensions), graduated cylinder (liquids)
Time	s	Clock
Temperature	ºC, ºF, K	Thermometer
pH	No units	pH gauge
Pressure	kPa, atm, mmHg	manometer

How many significant figures matters for these numbers

Exact Values

- Integer-based values
  - Fractions: ½, ¾
  - Counting numbers: 2 electrons
  - Metric conversions: 1 meter = 100 cm
  - Two important exact conversions you need to remember
    - Distance: 1 inch = 2.54 cm
    - Energy: 1 calorie = 4.184 Joules
- Constants of nature are usually treated as exact values
  - Speed of light in vacuum, c = 2.99792458 x 10^8 m/s
  - Pi, π = 3.141592654… (no units)
  - Planck's constant, h = 6.62617636 x 10^{-34} J·s
  - Gas constant, R = 8.3144126 J/mol·K

How many significant figures doesn't matter for these numbers because they are known to more sig figs than you will ever need

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Derived (Calculated) Values

Value	Units most often used
Amount of a substance	mol
Energy	Joules, cal, kcal
Solution concentration	Molarity (M) = mol/L
Density	g/mL or g/cm^3
Velocity	m/s
Molar mass	g/mol

How many significant figures matters for these numbers
if you are using these numbers in the next step in a series of calculations

A Note About SI

SI = *Le Système International d'Unités* (International system of units)

- If you use SI units exclusively in a calculation, then the answer will always come out in SI units. This is why people memorize SI units (makes life easier).
- Unfortunately, in many cases, convention (or ease) is to not use SI units, so you might have to convert if you want to do the calculation in SI.

Not SI (must convert)	SI units
Grams (g)	Kilograms (kg)
Centimeters (cm)	Meters (m)

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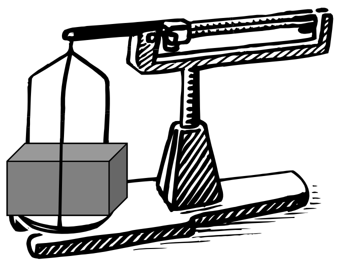
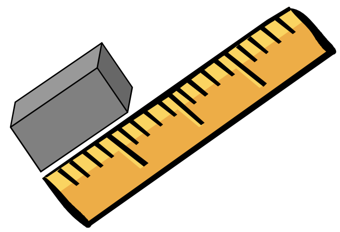
Example

Determining the density of a rectangular block of aluminum



Mass is 32.4030 grams



Dimensions are

2.0 cm x 2.0 cm x 3.0 cm

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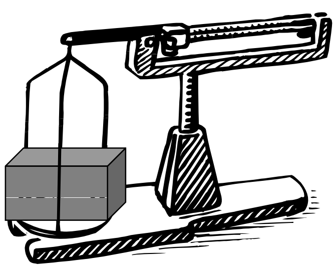


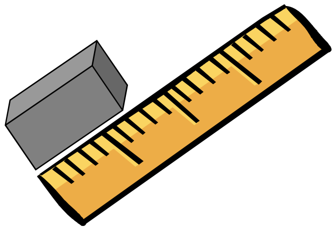
Mass is 32.4030 grams

How many digits are *significant*?

Significant = to have meaning

Mass measurement

$$\underbrace{32.4030}_{\text{Six digits of information}} \text{ grams}$$



How many digits are *significant*?

Length measurement

Dimensions are

2.0 cm x 2.0 cm x 3.0 cm

Need to calculate volume from this

2.0 cm

Two digits of information

Same number of digits of information for the other two length measurements

Rules for Sig Figs

- Generally, count the digits
- Zeroes written to the left don't count

  0.00056 has 2 sig figs
- Zeroes written to the right do count

  81.00 and 0.0008100 both have 4 sig figs
- Convention for numbers not containing a decimal point

  7200 has 2 sig figs, 7200. has 4 sig figs
- See pp. 22-23 in the text for rules

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Keeping Track of Sig Figs

Density calculation:

1. Calculate volume from length measurements
2. Calculate density from mass and volume

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Keeping Track of Sig Figs

Volume calculation

$V = L \times W \times H$

$$= \overset{\textbf{2 sig figs}}{(2.0\text{ cm})} \times \overset{\textbf{2 sig figs}}{(2.0\text{ cm})} \times \overset{\textbf{2 sig figs}}{(3.0\text{ cm})}$$

$$= 12.0\text{ cm}$$

We note at this point that the final calculation of density must be rounded to 2 sig figs, but since this is an intermediary step in the calculation, it is good to keep one more sig fig than necessary for now. We will round to the correct number of sig figs when we get to the last calculation.

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Keeping Track of Sig Figs

6 sig figs

$$D = \frac{m}{V} = \frac{32.4030\ \text{g}}{12.0\ \text{cm}^3}$$

Remember this number is supposed to be 2 sig figs

Calculator gives 2.70025

$$= 2.7\ \text{g}/\text{cm}^3$$

Must round to 2 sig figs

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Rules for Sig Figs

- Multiplication/division rule
  - The measurement with the least total sig figs wins
- Addition/subtraction rule
  - The measurement with the least decimal places (compared to the decimal point) wins
- Other rules can wait until you need to do more complicated calculations

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Example of Addition Rule

Problem

Find the sum: 28.6 + 8.289 + 0.003 + 1007.56

Solution

Line up the numbers at the decimal point, compare, and cut off at least significant (compared to decimal point)

$$
\begin{array}{r}
28.6 \\
8.289 \\
0.003 \\
+\ 1007.56 \\
\hline
1044.452
\end{array}
\quad \xrightarrow{\text{rounds to}} \quad 1044.5
$$

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Some Measurements and Conversions You Need to Know

Two types of conversions

- Proportional
  - Time
  - Length or distance
  - Volume
  - Mass
- Equations
  - Temperature

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Time Conversions

2.5
hours

?
seconds

Metric Prefix Meanings for Conversions

Prefix	Meaning	Example
Centi (c)	1/100th of	1 cm = 0.01 m 100 cm = 1 m
Milli (m)	1/1000th of	1 mL = 0.001 L 1000 mL = 1 L
Kilo (k)	1000 of	1 kg = 1000 g
Micro (μ)	10^{-6} of	1 μmol = 10^{-6} mol 1,000,000 μmol = 1 mol
Nano (n)	10^{-9} of	1 nm = 10^{-9} m

Conversions Using Metric System

1) How many moles are in 12.2 mmol?

12.2 mmol

2) Red light has a wavelength of 630 nm. How many meters is that?

630 nm

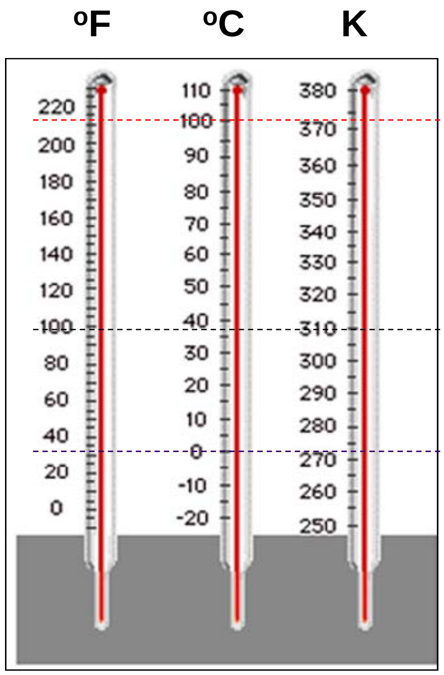
Volume Conversions

Important volume conversion to remember: $1\,\mathrm{mL} = 1\,\mathrm{cm}^3$

A can of soda is marked as having 258 cc of soda in it. How many liters is this?

258 cm^3

Temperature Scales



From Chemistry & Chemical Reactivity 5th edition by Kotz / Treichel. C 2003. Reprinted with permission of Brooks/Cole, a division of Thomson Learning: www.thomsonrights.com. Fax 800-730-2215.

Boiling point of water

212 ºF = 100 º C = 373 K

Human body temperature

98.6 ºF = 37 º C = 310 K

Freezing point of water
Melting point of ice

32 ºF = 0 º C = 273 K

Temperature Conversions Require Equations

Celsius (ºC) ↔ Kelvin (K)

You need to memorize this conversion!!!

$$\mathrm{K} = {}^{o}\mathrm{C} + 273.15 \quad or \quad {}^{o}\mathrm{C} = \mathrm{K} - 273.15$$

Example: A gas has a temperature of 25.8 ºC. What is the temperature in Kelvin?

K = ºC + 273.15

$= \underbrace{25.8 + 273.15}_{\text{Sig figs: One place beyond decimal pt}} = 298.95\ \mathrm{K}$

Rounds to 299.0 K

Really Big Numbers

In a 22.4 liter sample of air at standard conditions, there are approximately this many particles present:

602,204,531,000,000,000,000,000

2 3 3 3 3 3 3 3

= 6.02204531 x 10^{23} particles

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Really Small Numbers

A single snowflake has a mass of approximately

0.000 0030 kg

3 3

=3.0 x 10^{-6} kg

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Calculations Using Scientific Notation

- A typical snowflake has $100=10^2$ ice crystals
- A single ice crystal has 10^{18} water molecules
- A water molecule has a mass of 3.0×10^{-26} kg
- Therefore, a typical snowflake has a mass of approximately

$$10^2\ crystals \times \frac{10^{18}\ water\ molecules}{1\,crystal} \times \frac{3.0 \times 10^{-26} kg}{1\ water\ molecule} = 3.0 \times 10^{-6} kg$$

Data taken from http://hypertextbook.com/facts/2001/JudyMoy.shtml

Scientific notation (aka Exponential notation)

- A nice way to represent big and small numbers
- Makes it easy to indicate significant figures
  9000 written with two sig figs is 9.0×10^3
- Makes it easy to estimate answers
  $(3.0 \times 10^8) \times (2.0 \times 10^{-6}) = 6.0 \times 10^2$
- Scientific notation and your calculator → try the practice problems in the Assignments section on the course website to make sure you are proficient at using scientific notation in your own calculator

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What is an Atom?

B.C.E. – Democritus: an atom is the smallest particle of matter

1800's – Electrons exist and they have some properties (negative charge, very small mass)

Late 1800's-Mid 1900's – Protons and neutrons exist and they have some properties (protons are +, neutrons are neutral, have nearly same mass which is > electron mass)

What do you already know about atoms?

How is the atom organized?

What is the nucleus?

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How small is the nucleus?

Work in pairs. Take one paper as they come around. Each paper contains traces of 10 pennies in a box. The box is 10. cm × 10. cm square.

Goal: find the area of a single penny

- One person closes eyes and makes 50 marks in the box while the other person counts the marks
- If a mark falls outside the box, do over
- Afterward, count how many marks fell inside the pennies
- Use a proportion to figure out what the area of a single penny is

Proportion

$$\frac{\#\ of\ \text{"}hits\text{"}}{total\ \#\ of\ marks} = \frac{area\ of\ 10\ pennies}{total\ area\ of\ the\ box}$$

$$\frac{\#\ of\ \text{"}hits\text{"}}{50} = \frac{area\ of\ 10\ pennies}{100\ \text{cm}^2}$$

What is the area of a single penny?

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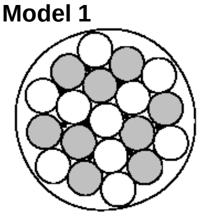
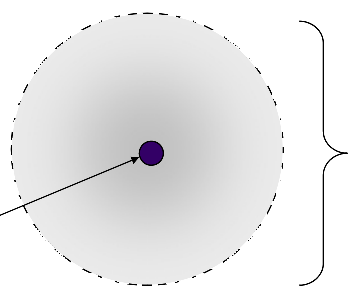
What does an atom look like?

Model 1



Thomson, 1898
has + and - charges

Model 2
Rutherford, 1910
Nucleus (+) is very small,
atom is mostly empty space



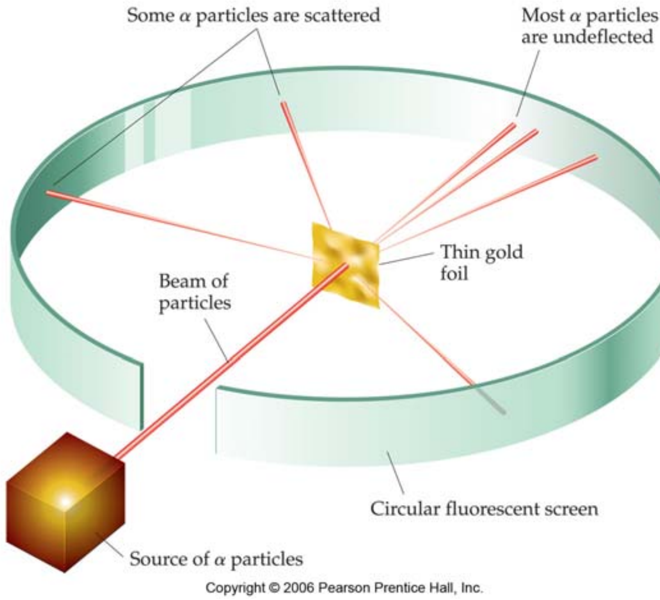
Electrons occupy most of atom's space (though they are tiny) 10^{-8} cm.

Nucleus contains most of mass and all of + charge. Size is 10^{-12} cm.

Note: not drawn to scale!

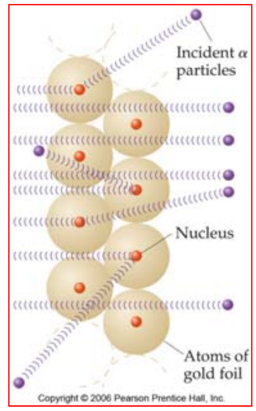
http://micro.magnet.fsu.edu/primer/java/scienceopticsu/powersof10/

More about Rutherford's Experiments



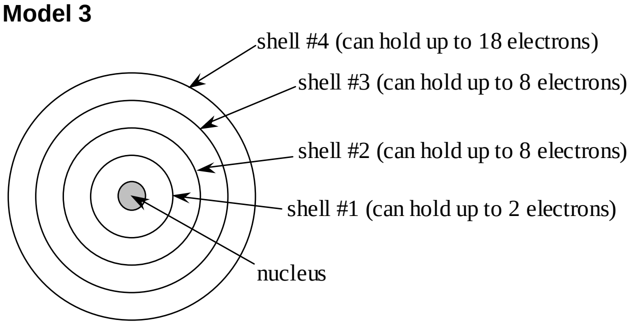
"It was almost as if you fired a 15-inch shell into a piece of tissue paper and it came back and hit you."

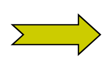
– Ernest Rutherford



What does an atom look like?

Model 3





Bohr, 1912

Note: not drawn to scale!

Model explains the hydrogen spectrum (stay tuned until chapter 6…)

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What does an atom look like?

Model 4 Schrodinger, Heisenberg, Bohr, de Broglie and some others, 1912-1924

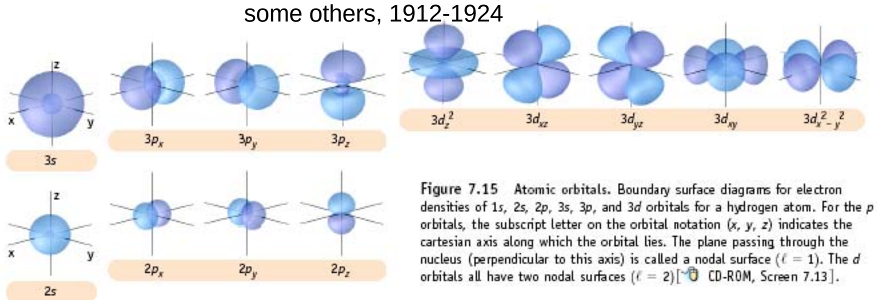
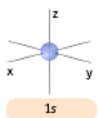


Figure 7.15 Atomic orbitals. Boundary surface diagrams for electron densities of 1*s*, 2*s*, 2*p*, 3*s*, 3*p*, and 3*d* orbitals for a hydrogen atom. For the *p* orbitals, the subscript letter on the orbital notation (*x*, *y*, *z*) indicates the cartesian axis along which the orbital lies. The plane passing through the nucleus (perpendicular to this axis) is called a nodal surface ($\ell = 1$). The *d* orbitals all have two nodal surfaces ($\ell = 2$) [CD-ROM, Screen 7.13].



Quantum Mechanical Model = current working model (stay tuned until chapter 7 for more information…)

What we've learned so far

- Dimensional analysis is a way to figure out problems by putting the units in the right places so it works out
- You have to keep track of significant digits while doing a calculation so that you will report an answer only to the limit of its significance
  - Multiplication/division rule goes with least sig figs
  - Addition/subtraction rule line up the decimals
  - Obey proper order of operations
- Scientific (exponential) notation is a convenient way to report sig figs and also helps with doing calculations
- Atoms are really tiny
- The nucleus of an atom takes up only one-ten thousandth $\left(\frac{1}{10,000}\right)^{th}$ the space that an atom occupies
- There are lots of models for the atom, ranging in sophistication, and each has limits and usefulness

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