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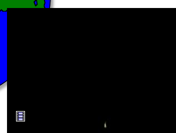
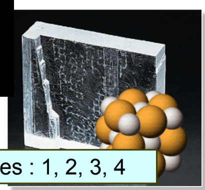
Welcome to the World of Chemistry





Syllabus Learning Outcomes : 1, 2, 3, 4

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Required materials

Needed for

- tests and quizzes
- reading assignments
- homework

1) Kotz, Treichel, Townsend; Chemistry&Chemical Reactivity, 7th ed; Thomson
2) CHM 101 Syllabus for homework and reading assignments
3) Scientific calculator (non-graphing type must be used for exams)
4) Klein, General Chemistry I: Mastering the Fundamental Skills
5) Handouts – constants and periodic table – work problems using these resources as these will be attached to all exams
6) Attendance Required

Syllabus Learning Outcomes : 1, 2, 3, 4

Instructor: Dr. Michael Love - Office Hours L107 (see syllabus)
Email: mlove@sussex.edu
Course: Chemistry 101

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Grading

• Hour Examinations (3)	30%
• Quizzes (about weekly)	10%
• Lab	30%
• Final Exam (comprehensive)	30%

- No Make Up Quizzes
- Mobile phones, PDAs, and graphing calculators not permitted for exams
- Academic Dishonesty Penalty – You fail the assignment and will be reported to the Dean of Students, minimum

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Student Conduct

- Silence mobile phones, pagers and other devices that may disturb others during class
- Harassment of any kind is not permitted. Report incidents immediately to me.

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Get Supplemental Help

- Instructor: Dr. Michael Love
  - Lecture time, office hours L107 (see syllabus)
  - Email: mlove@sussex.edu
- Science Resource Center (B300)
- Students with special needs, contact me.

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How to Succeed in Chemistry

- Take, read and summarize your lecture notes
- Do the homework
- Ask questions and keep up with the material
- Study 12 hours weekly based on 3 class hours
- Prepare, Practice, Repetition, Effort
  - Redo the homework as practice for quizzes
  - Redo quizzes and homework for the exams
  - Work problems forwards and backwards
  - Study lab problems
  - Redo quizzes, homework, and exams for the final
  - Join a study group

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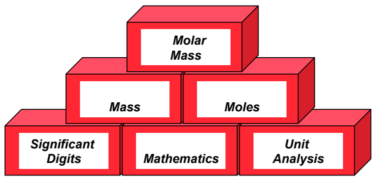
Concepts are Cumulative



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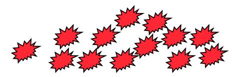
Concepts are Cumulative



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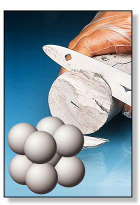
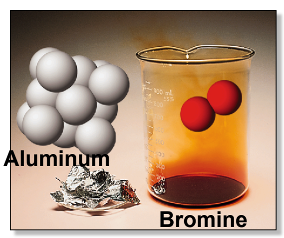
Define Chemical Elements

- **Find on periodic table**
- **pure substances that cannot be decomposed by ordinary means to other substances.**



Sodium



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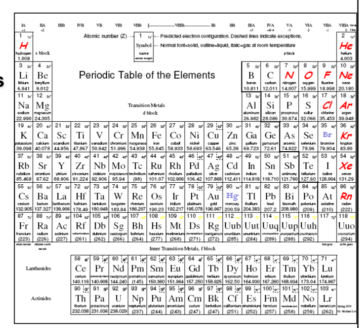
What is the Periodic Table?

- **Lists**
  - **element names**
  - **element symbols**
  - **atomic numbers**
  - **molar masses**
  - **electron configuration**



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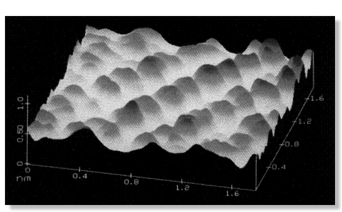
Define Atom

- **smallest particle of an element that has the chemical properties of the element.**



Copper atoms on silica surface.

Distance across = 1.8 nanometer (1.8 x 10^{-9} m)

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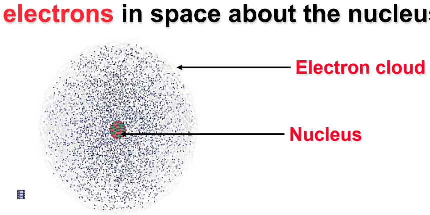
Atom consists of a

- **nucleus that contains**
  - **protons and**
  - **neutrons**
- **electrons in space about the nucleus.**



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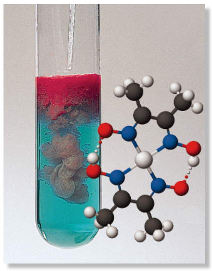
CHEMICAL COMPOUNDS

consist of different atoms and can be decomposed to those atoms.



The red compound is composed of
- nickel (Ni) (silver)
- carbon (C) (black)
- hydrogen (H) (white)
- oxygen (O) (red)
- nitrogen (N) (blue)

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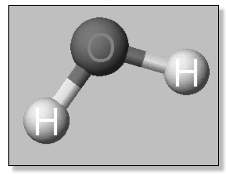
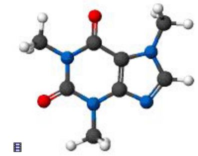
MOLECULE contains 2 or more nonmetal atoms retaining the chemical characteristics of the substance.

Composition of molecules is given by a **MOLECULAR FORMULA**

H_2O Br_2 $C_8H_{10}N_4O_2$ - caffeine





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Elements react to form compounds

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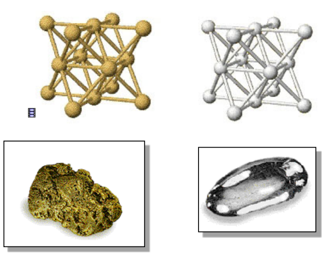
The Nature of Matter

Gold Mercury



Chemists are interested in the nature of matter and how this is related to its atoms and molecules.

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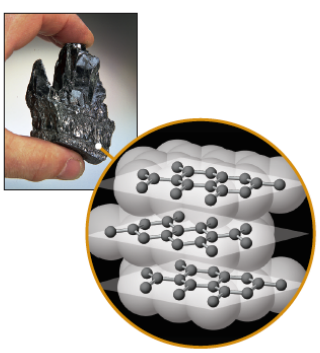

Graphite — layer structure of carbon atoms reflects physical properties.

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Chemistry & Matter

- We can explore the MACROSCOPIC world — what we can see —
- to understand the PARTICULATE worlds we cannot see.
- We write SYMBOLS to describe these worlds.

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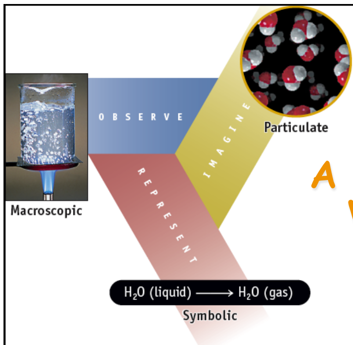
A Chemist's View of Matter



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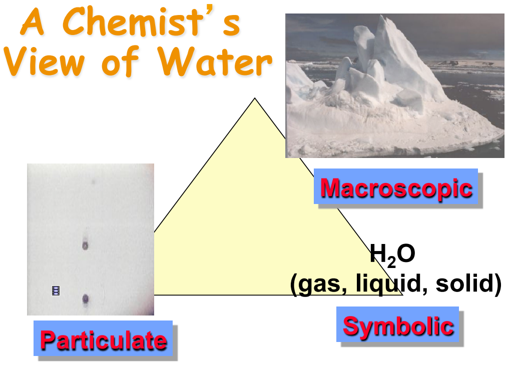
A Chemist's View of Water



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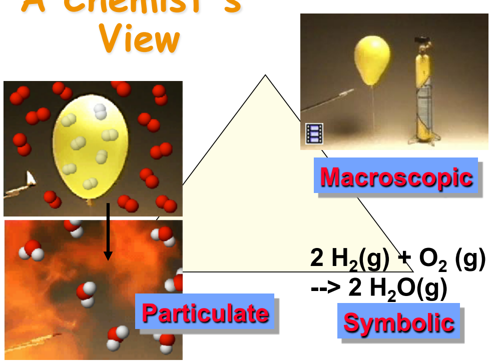
A Chemist's View



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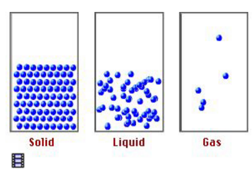
Matter is in Constant Motion

Matter is atoms and molecules in motion.

*



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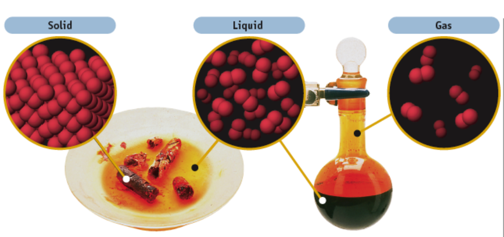
STATES OF MATTER



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Matter has States

- ***SOLIDS*** — have rigid shape, fixed volume. External shape can reflect the atomic and molecular arrangement.
  - Reasonably well understood.
- ***LIQUIDS*** — have no fixed shape and may not fill a container completely.
  - Not well understood.
- ***GASES*** — expand to fill their container.
  - Good theoretical understanding.

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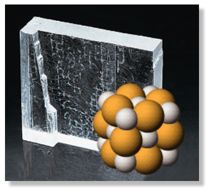
Physical properties describe matter

- color
- liquid
- melting point
- boiling point
- odor
- density
- brittleness



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Physical changes alter the state of matter

- boiling a liquid
- freezing a liquid
- melting a solid
- subliming a solid
- depositing a gas
- condensing a gas
- dissolving a solid in a liquid to give a solution
- crushing a solid

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Chemical properties and chemical change alter the composition of matter

- **Chemical change** or ***chemical reaction*** — transformation of one or more atoms or molecules into one or more different molecules.

Nitrogen Triiodide Detonation

*

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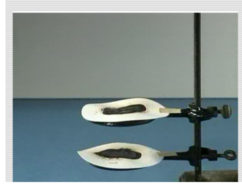
Examples of Chemical Change

- Silver tarnishes
- Gasoline burns
- Battery provides electricity
- Garbage rots
- Kittens breathe
- Plants photosynthesize
- Hard boil an egg
- Iron rusts

Nitrogen Triiodide Detonation



	Physical Properties
Property	**Example**
Qualitative	
Color	Sulfur is *yellow.*
Odor	Hydrogen sulfide *stinks.*
Solubility	Table salt *dissolves in water.*
Hardness	Diamond is *exceptionally hard.*
Electrical conductivity	Copper *conducts electricity.*
Quantitative	
Mass	A nickel has a mass of *5 grams.*
Temperature	Water for the bath is at *40 °C.*
Melting point	Lead melts at *327.5 °C.*
Density	At 20 °C, water has a density of *0.998 grams per milliliter.*
	Chemical Properties
Substance	**Typical Chemical Property**
Iron	*Rusts* (combines with oxygen to form iron oxide)
Carbon	*Undergoes combustion* (combines with oxygen to form carbon dioxide)
Silver	*Tarnishes* (combines with sulfur to form silver sulfide)
Sodium	*Reacts* violently *with water* to form hydrogen gas and a solution of sodium hydroxide.
Nitroglycerin	*Explodes* (decomposes, when detonated, to a mixture of gases)

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Types of Observations and Measurements

- **QUALITATIVE observations** include changes in color and physical state. (What is it?)
- **QUANTITATIVE measurements** involve numbers. (How much?)
- Use **SI units** — based on the metric system

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UNITS OF MEASUREMENT

Use SI units — based on the metric system

Length	**Meter, m**
Mass	**Kilogram, kg**
Time	**Seconds, s**
Temperature	**Celsius degrees, ˚C kelvins, K**

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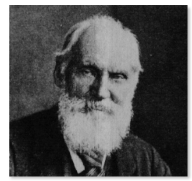
Temperature Scales

- **Fahrenheit**
- **Celsius**
- **Kelvin**

Anders Celsius 1701-1744



Lord Kelvin (William Thomson) 1824-1907

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Temperature Scales

	Fahrenheit	Celsius	Kelvin
Boiling point of water	212 ˚F	100 ˚C	373 K
	180˚F	100˚C	100 K
Freezing point of water	32 ˚F	0 ˚C	273 K

Notice that 1 kelvin degree = 1 degree Celsius

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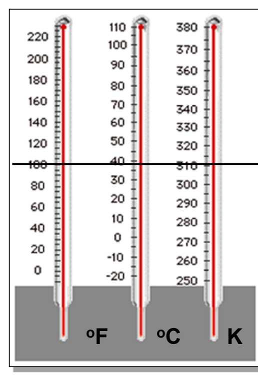

Temperature Scales

100 °F
38 °C
311 K

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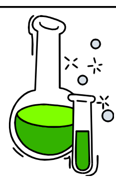

Calculations Using Temperature

- **Generally use temperatures in Kelvin**

  K = ˚C + 273

  Given on help sheet

- **Body temp = 37 ˚C + 273 = 310 K**
- **Liquid nitrogen = -196 ˚C + 273 = 77 K**

Convert Temperatures

- Equation for converting °C to °F.

  °F = 1.8 °C + 32

  Given on help sheet

- Be able to solve this equation for °C.

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UNITS OF MEASUREMENT

Use SI units — based on the metric system

Length	**Meter, m**
Mass	**Kilogram, kg**
Time	**Seconds, s**
Temperature	**Celsius degrees, ˚C** **kelvins, K**

Know Metric Prefixes

Given

Prefix	Abbreviation	Meaning	Value	Example
mega-	M	million	1×10^{6}	megagram
kilo-	k	thousand	1×10^{3}	kilogram
-	-	1	1×10^{0}	gram
deci-	d	tenth	1×10^{-1}	decigram
centi-	c	hundredth	1×10^{-2}	centigram
milli-	m	thousandth	1×10^{-3}	milligram
micro-	μ	millionth	1×10^{-6}	microgram
nano-	n	billionth	1×10^{-9}	nanogram
pico-	p	trillionth	1×10^{-12}	picogram
femto-	f	-	1×10^{-15}	femtogram

BIG

small

The number zero 0 is smaller than all of these

Convert Units of Mass

Q) 1.0cg equals how many mg?

- **1 kilogram (kg) = 1 x 10^{3} grams (g)**
- **1 centigram (cg) = 1 x 10^{-2} grams (g)**
- **1 milligram (mg) = 1 x 10^{-3} grams (g)**
- **1 centigram (cg) = 10 milligram (mg)**

$$1.0cg \times \frac{1x10^{-2}\,g}{1cg} \times \frac{mg}{1x10^{-3}\,g} = 10.mg$$

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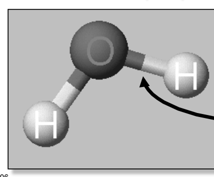
Convert Units of Length

- **1 kilometer (km) = 1 x 10^{3} meters (m)**
- **1 centimeter (cm) = 1 x 10^{-2} meters (m)**
- **1 nanometer (nm) = 1 x 10^{-9} meter (m)**
- **1 centimeter (cm) = 10 millimeters (mm)**



O—H distance =
9.58 x 10^{-11} m
9.58 x 10^{-9} cm
0.0958 nm

Convert Units of Volume

- **1 liter (L) = 1000 ml = 1000 cm^{3}**
- **1 liter (L) = 1 dm^{3}**
- **1 cm^{3} = 0.001 L = 1 mL**

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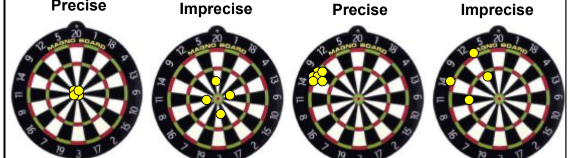
Accuracy and Precision

- **Accuracy – agreement with accepted value**
- **Precision – agreement of measurements with each other**

Accurate & Precise | **Accurate & Imprecise** | **Inaccurate & Precise** | **Inaccurate & Imprecise**



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Percent Error – measure of accuracy

$$\text{Percent Error} = \frac{\text{Experimental value} - \text{Accepted value}}{\text{Accepted Value}} * 100\%$$

Accepted (certified) mass = 10.0 g
Experimental (measured) mass = 11.0 g

Percent Error = (11.0 g – 10.0 g) / 10.0 g * 100% = 10.0%

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Standard Deviation – measure of precision

$$s = \sqrt{\sum_{1=1}^{n}\left[\frac{\left(X_i - \overline{X}\right)^2}{n-1}\right]}$$

s: standard deviation
Xi: measurement
i: measurement number
$\overline{X}$: average of all measurements
n: total number of measurements

Explain Significant Digits

- **Each number in a properly recorded measurement is a *significant digit* (or significant figure).**
- **The significant digits tell the precision (uncertainty, or the ±) of a measurement, where the uncertainty of the last significant digit is presumed to be ±1.**
- **When you count significant digits, start counting from the left with the first non-zero number.**
- **Placeholder zeros are presumed not to be significant.**
- **How many significant digits?**
- **7 6000 3001 2000.00 200.**
- **0.001 0.0013 100.001 1111**

Sort Numbers by Uncertainty

Step 1: Determine the ± based on SD. Step 2: Sort by ±.

Number	±	Most Uncertain (big ±)
1) 7		
2) 0.0013		
3) 6000		
4) 0.001		
5) 2000.00		
6) 200.	1	
		Least Uncertain (little ±)

Explain Exact Numbers

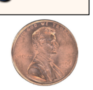
- **When we count something, it is an *exact number*.**
- **Metric conversions are *exact numbers*.**
- **Definitions are exact numbers**
- **Exact numbers have as many significant digits as are needed, so they don't affect significant digits in calculations.**
- **Example of exact numbers:**

  3 coins on this slide (counting)

  1 kg = 1000 g (metric)

  1 mL = 1 cm^3 (metric)

  1 in = 2.54 cm (definition)



A 1¢ coin

A 10¢ coin

A 25¢ coin

Rounding Numbers

- **All numbers from a measurement are significant. However, we often generate *nonsignificant digits* when performing calculations.**
- **We get rid of nonsignificant digits by *rounding off* numbers.**
- **There are three rules for rounding off numbers.**

Rules for Rounding Numbers

1. **If the first nonsignificant digit is less than 5, drop all nonsignificant digits.**
2. **If the first nonsignificant digit is greater than or equal to 5, increase the last significant digit by 1 and drop all nonsignificant digits.**
3. **If a calculation has two or more operations, retain all nonsignificant digits until the final operation and then round off the answer.**

Rounding Numbers

- A calculator displays 12.846239 and 3 significant digits are justified.
- The first nonsignificant digit is a 4, so we drop all nonsignificant digits and get 12.8 as the answer.
- A calculator display 12.856239 and 3 significant digits are justified.
- The first nonsignificant digit is a 5, so the last significant digit is increased by one to 9, all the nonsignificant digits are dropped, and we get 12.9 as the answer.

Rules for Significant Digits

- Addition and subtraction rule
- Multiplication and division rule
- Significant digits come from measurements (or data in a problem). Keep track of SD during calculations by applying each rule as you need it.

Add & Subtract Measurements

- **When adding or subtracting measurements, the answer is limited by the value with the *most uncertainty*.**
- **Lets add three mass measurements.**
- **The measurement 5 g has the greatest uncertainty (±1 g).**
- **The correct answer is** 15 g.
- **2 significant digits in the answer**

Significant Digits	*Nonsignificant Digits*
5	g
5	.1 g
+ 5	.34 g
15	.44 g

Multiply & Divide Measurements

- **When multiplying or dividing measurements, the answer is limited by the measurement with the *fewest significant digits.***
- **Lets multiply two length measurements.**
  - **5.15 cm × 2.3 cm = 11.845 cm^2**
- **The measurement 2.3 cm has the fewest significant digits, two.**
- **The correct answer is** 12 cm^2.

Powers of Ten

- **A *power of 10* is a number that results when 10 is raised to an exponential power.**
- **The power can be positive (number greater than 1) or negative (number less than 1).**

Powers of 10

Exponential Number		Ordinary Number
1×10^6	$= 10 \times 10 \times 10 \times 10 \times 10 \times 10$	1,000,000
1×10^3	$= 10 \times 10 \times 10$	1,000
1×10^2	$= 10 \times 10$	100
1×10^1	$= 10$	10
1×10^0	$= 1$	1
1×10^{-1}	$= \frac{1}{10}$	0.1
1×10^{-2}	$= \frac{1}{10} \times \frac{1}{10}$	0.01
1×10^{-3}	$= \frac{1}{10} \times \frac{1}{10} \times \frac{1}{10}$	0.001
1×10^{-6}	$= \frac{1}{10} \times \frac{1}{10} \times \frac{1}{10} \times \frac{1}{10} \times \frac{1}{10} \times \frac{1}{10}$	0.000 001

Bigger

Smaller

Scientific Notation

- Numbers in science are often very large or very small. To avoid confusion, we use *scientific notation*.
- Scientific notation uses the significant digits in a measurement followed by a power of ten.

power of 10

$$D.DD \times 10^{n}$$

significant digits

Applying Scientific Notation

- To use scientific notation, first place a decimal after the first nonzero digit in the number followed by the remaining significant digits.
- Indicate how many places the decimal is moved by the power of 10.

Scientific Notation Example

- There are 26,800,000,000,000,000,000,000 helium atoms in 1.00 L of helium gas. Express the number in scientific notation.
- Place the decimal after the 2, followed by the other significant digits.
- Count the number of places the decimal has moved to the left (22). Use this as the power of 10 to complete the scientific notation.

2.68×10^{22} atoms

- Note 3 significant digits in both numbers

Scientific Notation Example

- The typical length between two carbon atoms in a molecule of benzene is 0.000000140 m. What is the length expressed in scientific notation?
- Place the decimal after the 1, followed by the other significant digits.
- Count the number of places the decimal has moved to the right (7). Use this as a negative power of 10 to complete the scientific notation.

1.40×10^{-7} m

- Note 3 significant digits in both numbers

Scientific Notation and Calculators

- Be sure that you know how to enter scientific notation on your calculator!
- 1.40×10^{-7} is often entered

  1.40 EXP 7 +/-

- The display might read 1.4^{-07} to represent the number 1.40×10^{-7}.
- Calculators usually do not track significant digits. You need to do it.

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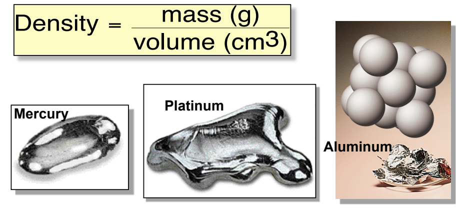
DENSITY - an important and useful physical property

$$\text{Density} = \frac{\text{mass (g)}}{\text{volume (cm3)}}$$



13.6 g/cm^3 21.5 g/cm^3 2.7 g/cm^3

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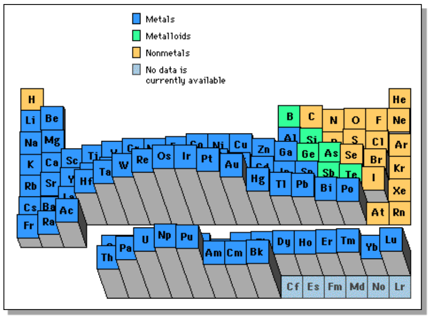
Relative Densities of the Elements



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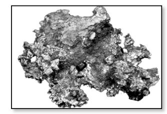
Problem A bar of copper has a mass of 57.54 g. It is 9.36 cm long, 7.23 cm wide, and 0.95 mm thick. Calculate density (g/cm^3).

$$\text{Density} = \frac{\text{mass (g)}}{\text{volume (cm3)}}$$

?

?



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Strategy

1. Get dimensions in common units.

2. Calculate volume in cubic centimeters.

3. Calculate the density.

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SOLUTION

1. Get dimensions in common units.

$$0.95 \text{ mm} \cdot \frac{1\text{cm}}{10 \text{ mm}} = 0.095 \text{ cm}$$

2. Calculate volume in cubic centimeters.

(9.36 cm)(7.23 cm)(0.095 cm) = 6.4289 cm^3

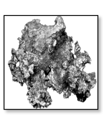
Note 2 significant digits

3. Calculate the density.

57.54 g / 6.4289 cm^3 = 8.9502 g/cm^3 = 9.0 g/cm^3

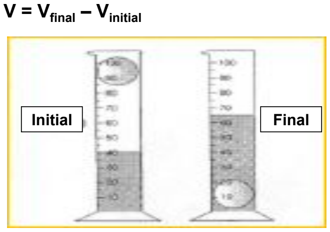
Carry the extra digits through and round the final answer

Determine Volume by Water Displacement



A solid will displace a volume of water, V, when submersed in a liquid.

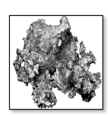
$V = V_{final} - V_{initial}$



How to Calculate Volumes

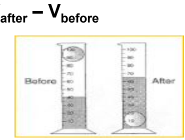
$V_{Cylinder} = \pi r^2 h$
π = 3.14159, r = radius of cylinder
h = height of cylinder, same units as r.

$V_{rectangular\ solid}$ = L x W x H
L = length, W = Width, H = Height
Use same units for L, W and H



$V_{irregular\ solid} = V_{after} - V_{before}$

determined by water displacement



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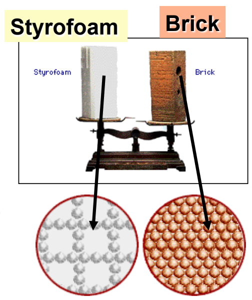
DENSITY

- Density is an **INTENSIVE** property of matter.
  - does NOT depend on quantity of matter.
  - Temperature and density
- Contrast with **EXTENSIVE**
  - depends on quantity of matter.
  - mass and volume.



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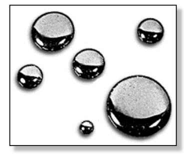
PROBLEM: Mercury (Hg) has a density of 13.6 g/cm^3. What is the mass of 95 mL of Hg in grams? In pounds?



Solve the problem using **DIMENSIONAL ANALYSIS.**

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PROBLEM: Mercury (Hg) has a density of 13.6 g/cm^3. What is the mass of 95 mL of Hg?

First, note that $1\ cm^3 = 1\ mL$

Strategy

1. Use density to calc. mass (g) from volume.
2. Convert mass (g) to mass (lb)

   Need to know conversion factor

   = 454 g / 1 lb

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PROBLEM: Mercury (Hg) has a density of 13.6 g/cm^3. What is the mass of 95 mL of Hg?

1. Convert volume to mass

$$95\ cm^3 \cdot \frac{13.6\ g}{cm^3} = 1.3 \times 10^3\ g$$

2. Convert mass (g) to mass (lb)

$$1.3 \times 10^3\ g \cdot \frac{1\ lb}{454\ g} = 2.8\ lb$$

Conversion Factors

- You should know the metric conversion factors and how to use them (handout) because these are often not provided.
- Conversion factors between British and metric units are usually provided (See Kotz Page A12 and A13)

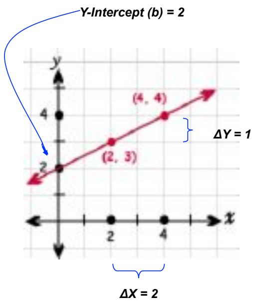
Standard Deviation

- Standard Deviation is a measure of precision for a series of measurements (X_i)
- Measured Values (X_i)
- Average Value (X_{bar})
- n = number of measurements
- Standard Deviation =

  Square Root {Sum $(X_i - X_{bar})^2$ / n}

Your calculator can often calculate standard deviation for you.

Graphing



- Use graph paper
- Each axis must have equal spacing
- Use ruler to draw best line through points
- The Y-intercept is the value where the line hits the y axis, b=2.
- $\Delta Y = Y_2 - Y_1 = 4 - 3 = 1$
- $\Delta X = X_2 - X_1 = 4 - 2 = 2$
- Slope (m) = $\Delta Y / \Delta X = 0.5$
- Use Excel or calculator

End