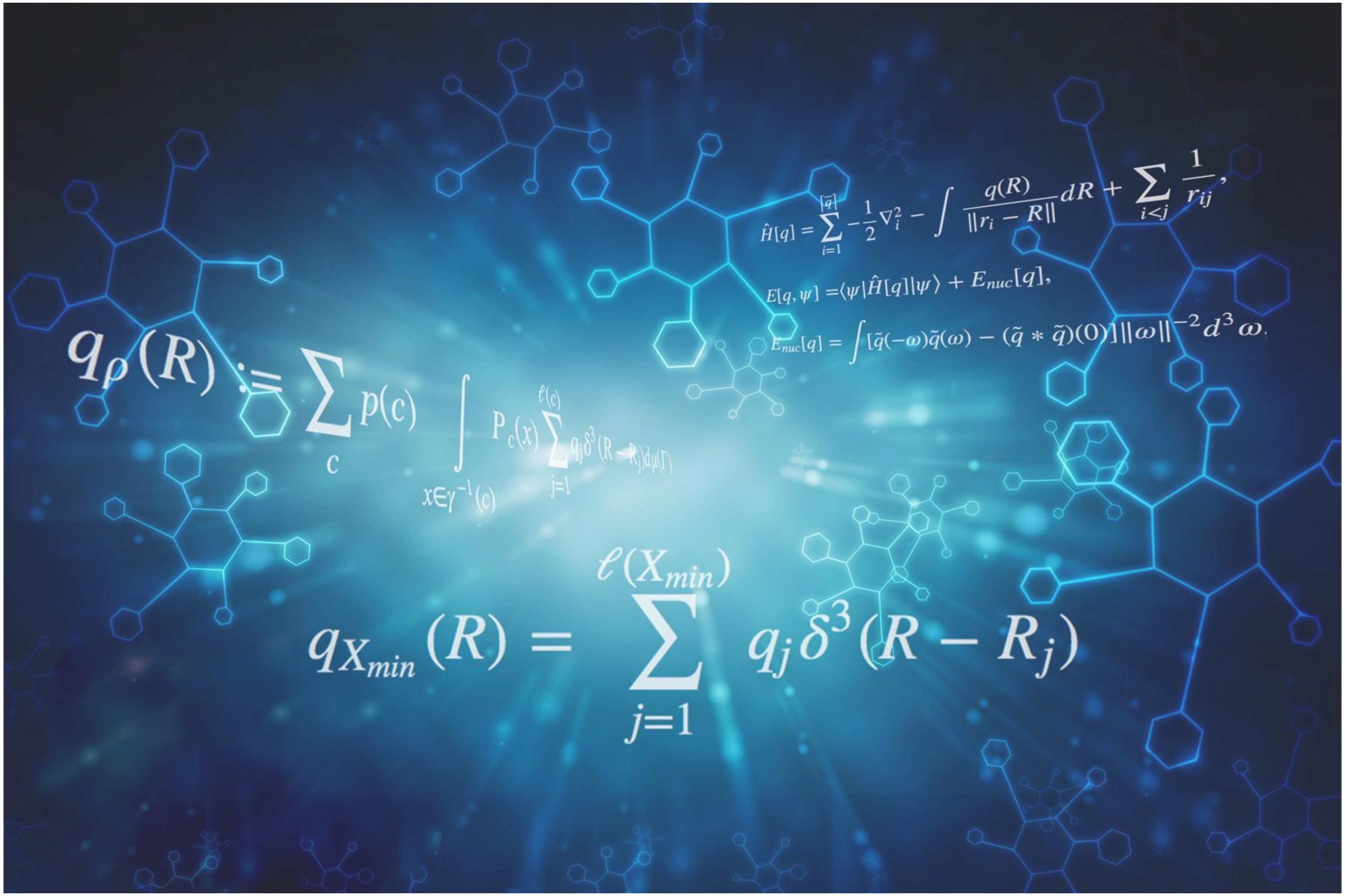


General Chemistry I, CHEM 103

Lecture 2: Measurements, mathematical treatment of measurements and dimensional analysis II

(aka the math you need to know for Gen Chem)

Announcements

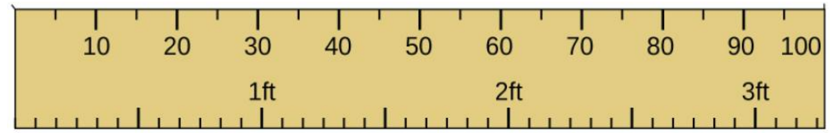
1. Reading from textbook: Chapter 1 from textbook Units 1.4-1.6

2. Achieve Homework will be posted tomorrow morning and will be due Wednesday evening. Announcement will be made on Moodle for you to see.

3. Instructor: Dr. Eugenia Vasileiadou
   - Office: Park 288
   - Office Hours: Tuesdays and Thursday 9:15 AM-10:15AM

   Course TA: Ramona Shekhar

   Office Hours: Thursdays 5PM – 7PM at Park Atrium



Significant figures

- <u>Significant figures are the digits in a number that carry meaningful information about its precision</u>. In science and engineering, no measurement is perfect. Significant figures prevent you from claiming your data is more precise than it actually is.
- Are these numbers the same?

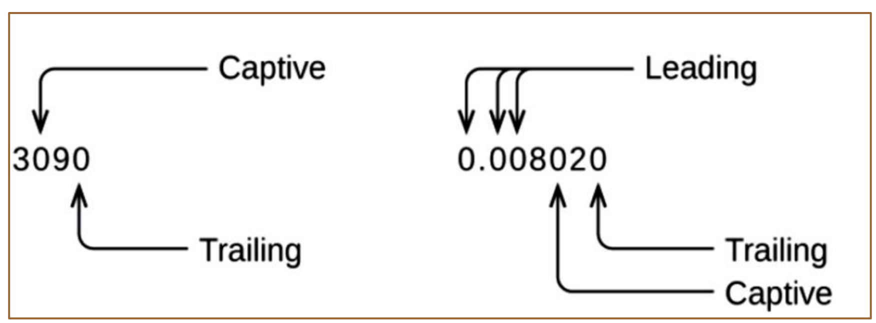
  0.35 g vs. 0.3500 g
- No – 0.3500 g is 100x more precisely known!
  - $0.3\underline{5}$ g: uncertain about the hundredths (10^{-2}) position; certain about all positions > 10^{-2}
- $0.350\underline{0}$ g: uncertain about the 10^{-4} position
- The value conveys knowledge about the instrument with which it was measured

$200

$5000

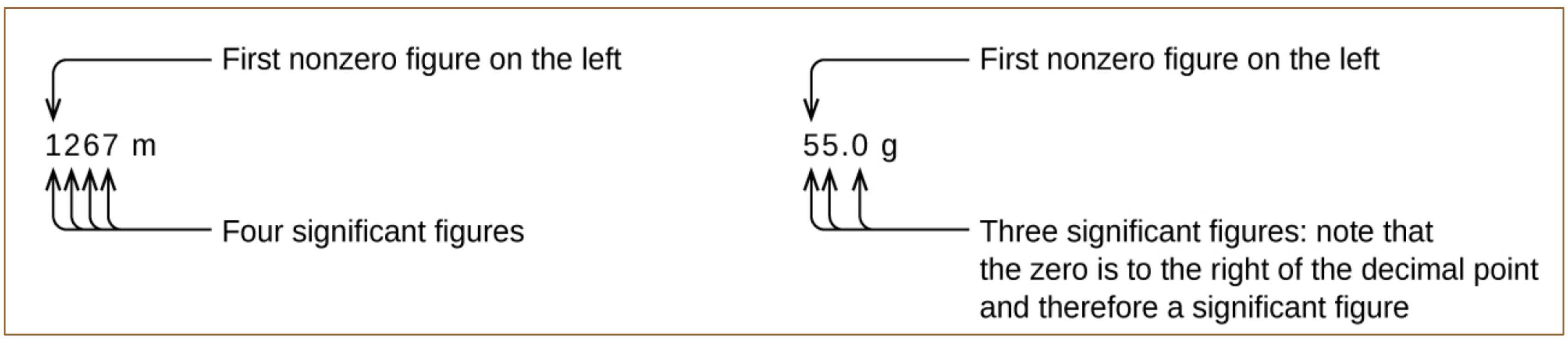
Significant figure rules

- All exact numbers have infinite significant figures
- Rest: all figures are significant, except for leading zeros and trailing zeros when there is no decimal point
  - 45.503
  - 0.406
  - 0.400
  - 0.0047
- Certain ambiguities exist, some of which can only be addressed by scientific notation
  - 7500 m
  - 7500. m
  - 7.50×10^3 m
- When in doubt – convert number to scientific notation
  - Perhaps the easiest way to see the # of SF!



Significant figure rules

1. Non-zero digits are always significant
2. Zeros between non-zero digits are always significant
3. Leading zeros are never significant
4. Trailing zeros are only significant if the number contains a decimal point



Significant figure rules: Let's practice together!

1. Non-zero digits are always significant
2. Zeros between non-zero digits are always significant
3. Leading zeros are never significant
4. Trailing zeros are only significant if the number contains a decimal point

Go to Moodle page, under today's lecture:

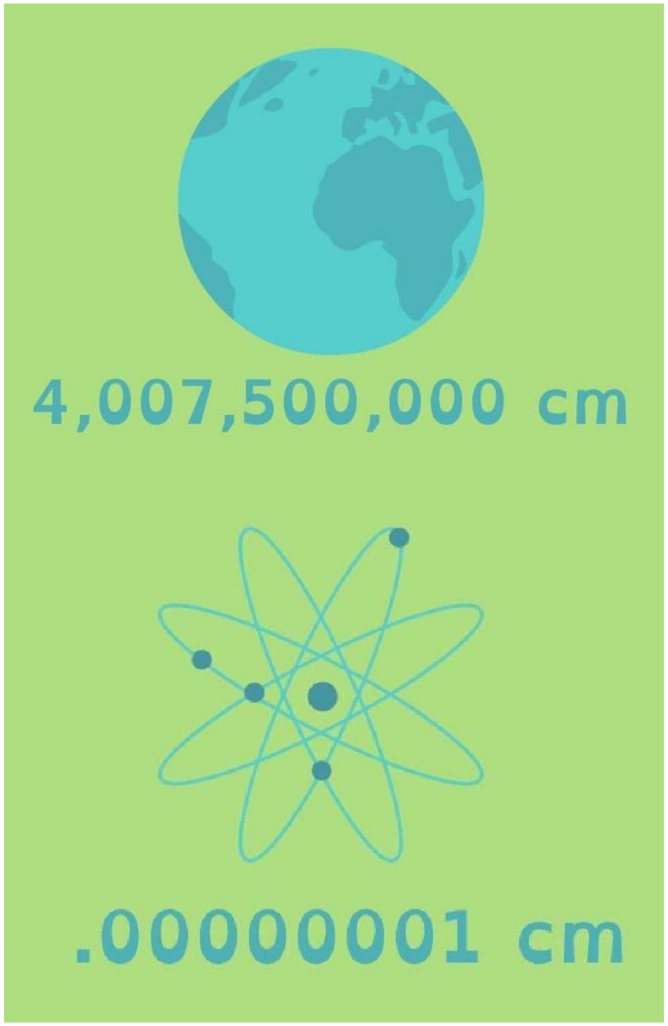
https://docs.google.com/forms/d/e/1FAIpQLSe40UW2-tjVPn_q2wl4cvjR-3D7UWjdPCt5wnzF26TuJuKemg/viewform?usp=header

Find Significant Figures

	How Many Sig Figs?	Which Figures are Significant?
81	2	8, 1
26.2	3	2, 6, 2
0.007	1	7
5200.38	6	5, 2, 0, 0, 3, 8
380.0	4	3, 8, 0, 0
78800	3	7, 8, 8
78800.	5	7, 8, 8, 0, 0

Scientific Notation

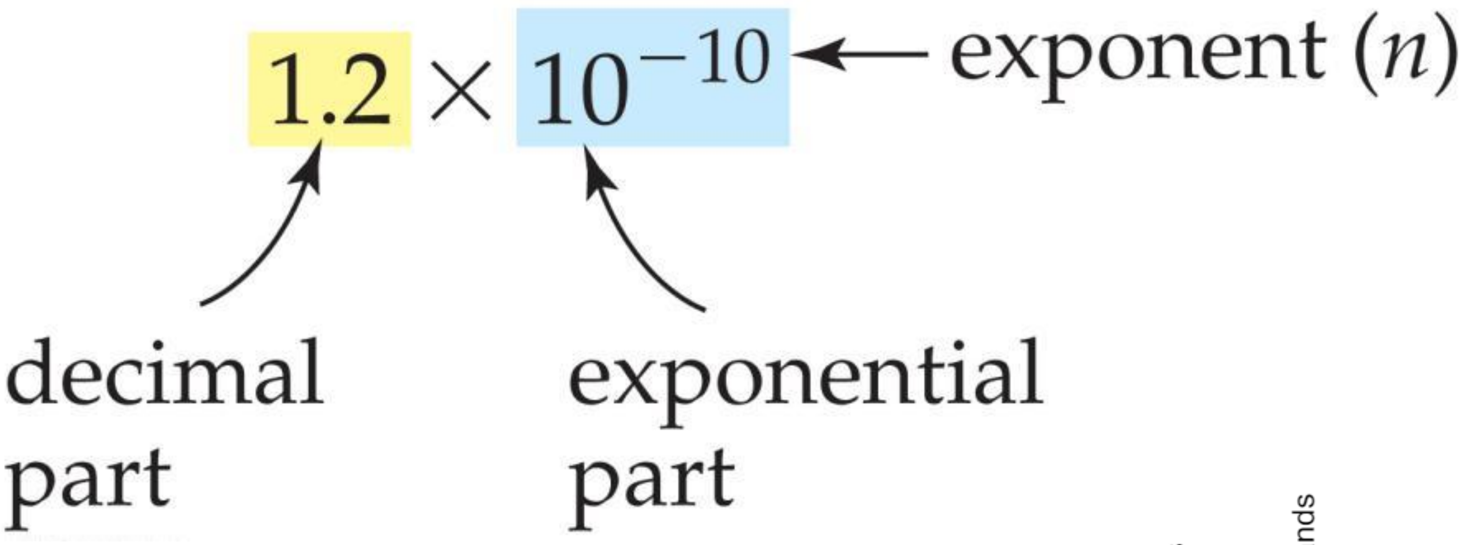
- Best way to:
  - express very big or very small numbers
  - emphasize significant figures
- Format: a.bcd… x 10^x
- 123,000,000 = 1.23 x 10^8
- 0.0004567 = 4.567 x 10^{-4}

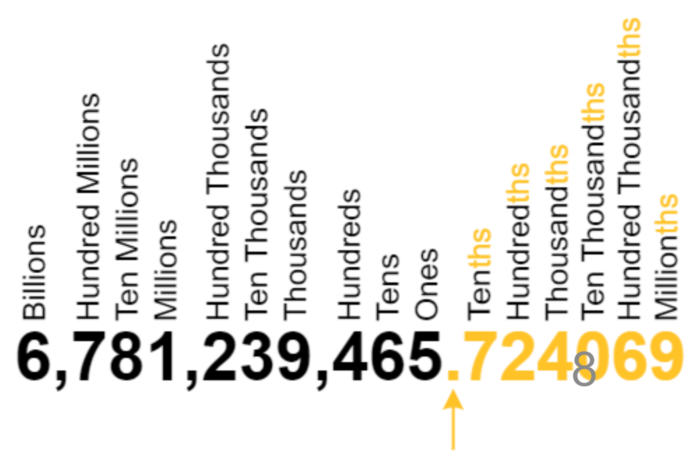


Scientific Notation

A number written in scientific notation has two parts:
The **decimal part**: **a number that is between 1 and 10.**

The **exponential part**: **10 raised to an exponent, *n*.**





Writing in Scientific Notation

A positive exponent means 1 multiplied by 10 *n* times.

A negative exponent (–*n*) means 1 divided by 10 *n* times.

$$10^0 = 1$$

$$10^1 = 1 \times 10 = 10$$

$$10^2 = 1 \times 10 \times 10 = 100$$

$$10^3 = 1 \times 10 \times 10 \times 10 = 1000$$

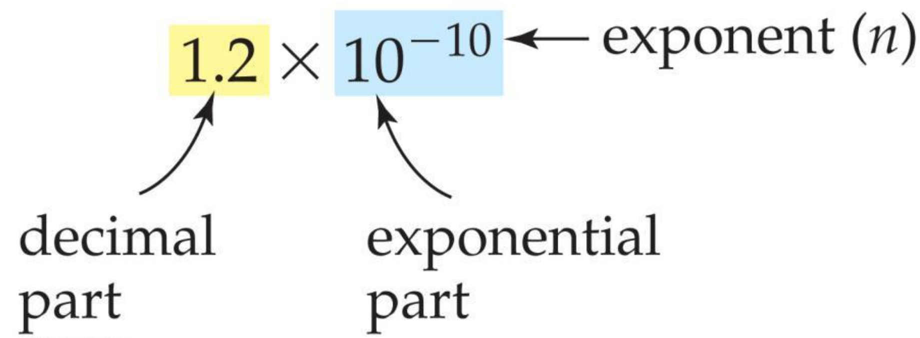
$$10^{-1} = \frac{1}{10} = 0.1$$

$$10^{-2} = \frac{1}{10 \times 10} = 0.01$$

$$10^{-3} = \frac{1}{10 \times 10 \times 10} = 0.001$$

Converting a Number to Scientific Notation

Find the decimal part. Find the exponent.



0.00000000012

Move the decimal point n times to obtain a number between 1 and 10 (the decimal part).

Multiply the decimal part by 10^n or 10^{-n} as appropriate.

Converting a Number to Scientific Notation

$5983 = 5.983 * 10^{3}$

3 2 1

$0.00034 = 3.4 * 10^{-4}$

1 2 3 4

If the decimal point is moved to the left, the exponent (n) is positive.

If the decimal point is moved to the right, the exponent (n) is negative.

Operations with significant figures: three basic rules

- Addition/subtraction
  - *Round the result to the uncertain order of magnitude of the quantity in which uncertainty begins at the highest order of magnitude.*
  - *(More typical, but less comprehensive way the rule is usually expressed): round the result to the last decimal place of the quantity with fewest decimal places.*

- Multiplication/division
  - *The number of significant figures in the answer is the same as the number of significant figures in the least precisely known number used in the calculation.*
  - Example: 34.0 g of sugar are dissolved in water to make 100.0 mL of solution. What is the solution's concentration in g/mL?

- Conversion factors used in calculations should be chosen so as not to limit precision
  - Should have more significant figures than rest of values
  - 1 lb = 453.592 g

Problem: What is the difference in mass between a nickel with mass 4.6837 g and a nickel with mass 4.6 g?

Addition and Subtraction: *Round the result to the decimal place of the least precisely known quantity (to the last common s.f.).*

significant for both numbers

significant for 4.6837 but not 4.6

$$\begin{array}{r} 4.6837\text{g} \\ -4.6\phantom{837\text{g}} \\ \hline 0.0837\text{g} \end{array}$$

10^{-4} is uncertain

10^{-1} is uncertain

round off to 10^{-1}

the last common s.f.

Answer = 0.1g

What if one nickel has mass 4.6837 g and another nickel had mass 4.6 g?

More than one operation? Follow the order of operations!

- Apply the appropriate SF rule as you perform each operation
- Underline last SF in each result to keep track of what you are doing
- Don't round (in your calculator or Excel spreadsheet) until the end to avoid roundoff error!
- Example

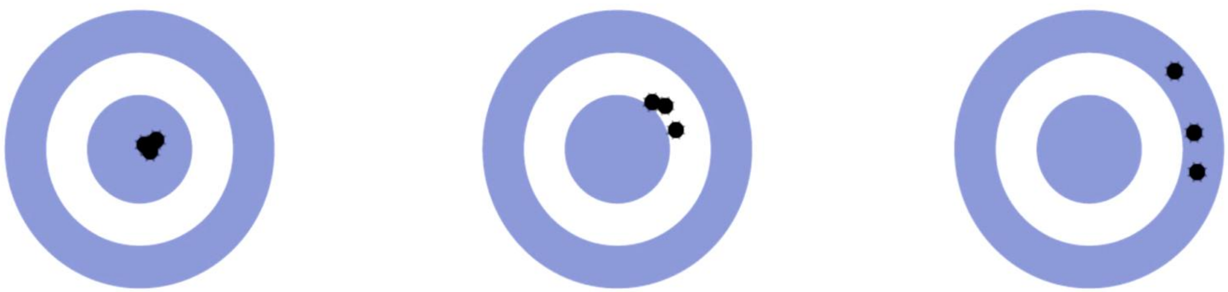
  $(1.646 \times 10^{3}) - (2.18 \times 10^{2}) + (1.36 \times 10^{4}) \times (5.17 \times 10^{-2})$

Point of confusion: significant figures and decimal places

- Both SF and DP inform us about precision (generally, the more SF/DP, the higher the precision)
- …but they do it in different ways
  - SF are affected by multiplication/division
  - DP are affected by addition/subtraction
- …and are independent of each other
  - 0.0403 g
    - 3 SF
    - 4 DP
- Don’t mix them up!

Accuracy and Precision

- Accuracy = agreement with “recognized” value
- Precision = reproducibility, knowing a value to many significant figures



Accurate
and precise
(a)

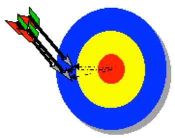
Precise,
not accurate
(b)

Not accurate,
not precise
(c)

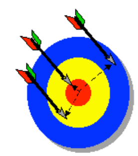
(a) These arrows are close to both the bull’s eye and one another, so they are both accurate and precise. (b) These arrows are close to one another but not on target, so they are precise but not accurate. (c) These arrows are neither on target nor close to one another, so they are neither accurate nor precise.

Additional uncertainty: systematic and random errors

Error



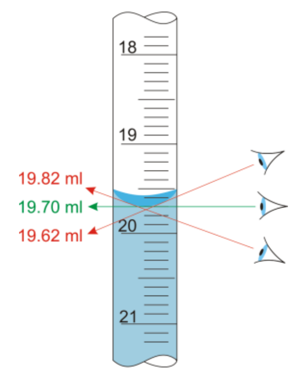
Systematic Error



Random Error

Example:
inaccurately calibrated or not "zeroed" instrument

Example:
technical errors (e.g., parallax error)



Dimensional analysis and unit conversion

- First, some definitions
- Dimension = physical quantity
  - Length, temperature, energy…
- For each dimension, there can be multiple units used to measure it
  - Some units used to measure energy (SI unit bolded): **joules (J)**, calories (cal), electron-volts (eV)…
  - Length: **meters (m)**, centimeters (cm), inches, feet…
- Dimensional analysis: process of converting from one dimension or unit to another

How to perform dimensional analysis

How much volume, in L, is occupied by 1.32 kg of mercury (Hg)? (density = 13.56 g/cm³)

- Put units on every number!
- Identify the unit/dimension you need to change (original • final)
- List all appropriate unit conversion factors and formulas that allow dimensional analysis
  - Recognize that if a = b, then $\frac{?}{?} = \frac{?}{?} = 1$ and multiplying by 1 doesn't change the number's value
  - A value divided by itself is canceled (becomes a factor of 1)
- Arrange factors appropriately so that all unwanted units and dimensions cancel
- Calculate (and check SF)

How to perform dimensional analysis

How much volume, in L, is occupied by 1.32 kg of mercury (Hg)? (density = 13.56 g/cm^3)

Practice problem

- A 2.67 g piece of gold with density = 19.30 g/cm^3 is hammered into a sheet whose area is 41.9 ft^2. What is the height of the sheet in μm? (2.54 cm = 1 inch, 12 in = 1 ft)

- Let's spend a few minutes for you to work out your solution and upload it to Moodle page, under today's lecture.

Upload a picture of your work for this problem. This is counted only for completion, not the outcome of your response!