Unit 1: Measurement and Calculations

Chemistry Chapter 1 & 3

GA Performance Standards

- SCSh5.a. Trace the source on any large disparity between estimated and calculated answers to problems.
- SCSh7.b. Universal principles are discovered through observation and experimental verification.
- SCSh8.a. Scientific investigators control the conditions of their experiments in order to produce valuable data.
- SCSh8.b. Scientific researchers are expected to critically assess the quality of data including possible sources of bias in their investigations' hypotheses, observations, data analyses, and interpretations.
- SCSh8.c. Scientists use practices such as peer review and publication to reinforce the integrity of scientific activity and reporting.
- SCSh8.d. The merit of a new theory is judged by how well scientific data are explained by the new theory.
- SCSh9.d. Establishing context

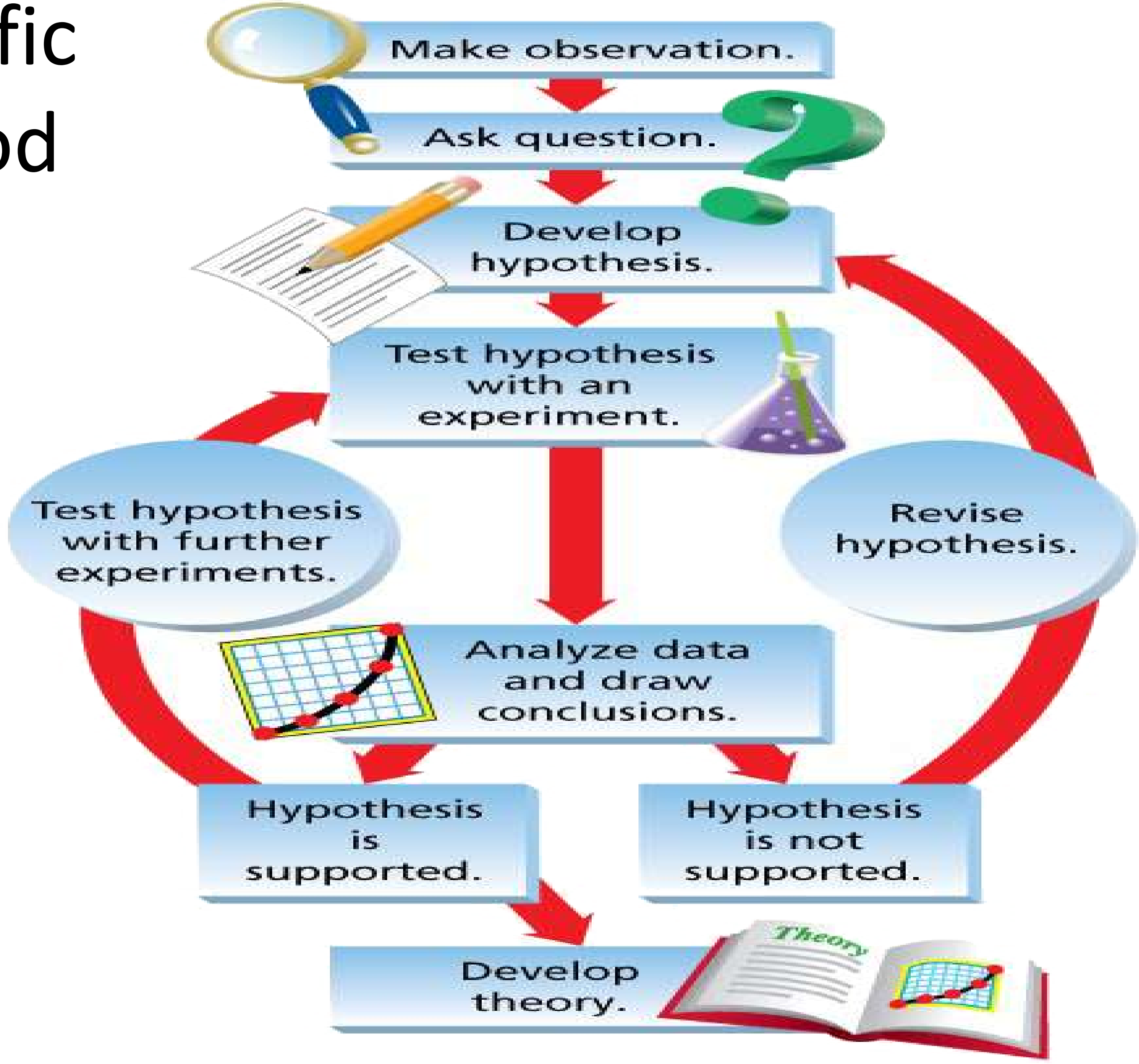
Introduction

- What is Chemistry
  - Study of matter
  - Matter is anything that has mass and occupied space
  - Examples:
  - Things that are NOT matter:

Scientific Method

- Scientific method is a logical, systematic approach to the solution of a scientific problem
  1. Make observation
  2. Ask a question
  3. Form a hypothesis
  4. Experiment
  5. Analyze data
  6. Draw Conclusion
  7. Develop Theory or re-evaluate hypothesis

Scientific Method



Scientific Method in the real world

- Teacher Example:

- Group Example:

Scientific Theory vs Scientific Law

- Scientific Theory: a well tested explanation for observations and/or experimental result
  - Attempts to explain why or how
  - Can not be proven only can get stronger
  - Kinetic Theory of matter stated atoms are in constant motion and explains how they move
- Scientific law: a statement that summarizes the results of many observations and experiment
  - Does NOT try to explain why/how
  - Gravity

Measurement

Units of measure
Significant Figures

Measurement

- Measuring with SI Units
- The metric system units are based on multiples of 10 and can be converted easily
- International System of Units (SI) is a revised version of the metric system
- **The five SI base units commonly used by chemists are the meter, the kilogram, the Kelvin, the second, and the mole.**

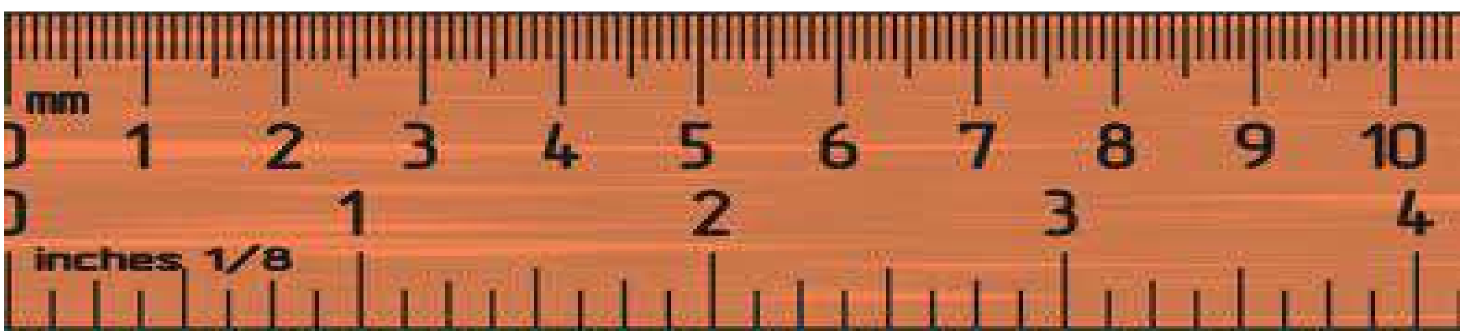
SI Base Unit

Quantity	SI standard unit	Base unit**
length	Meter (m)	Meter (m)
Mass	Kilogram (Kg)	Gram (g)
Temperature	Kelvin (K)	Kelvin (K)
Time	Seconds (s)	Seconds (s)
Volume	Decimeter cubed (dm^3)	Liter (L)
Amount of a substance	Mole (mol)	Mole (mol)
Heat and Energy	Joules (J)	Joules (J)
Force and weight	Newton (N)	Newton (N)

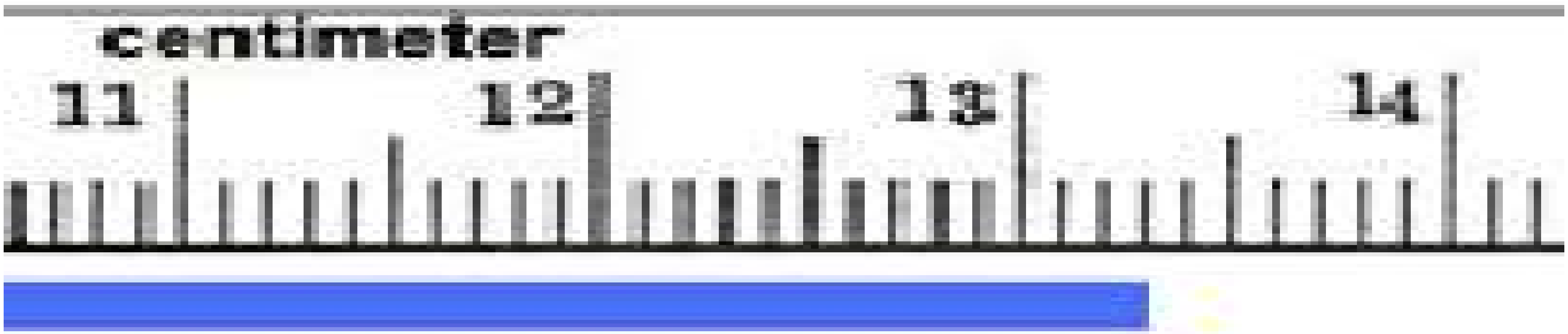
Metric Prefixes

- Added to the base unit to make it larger or smaller
- Changes by powers of 10
- Physical science prefix pneumonic: "King henry died by drinking chocolate milk"
- kilo, hecto, deca, base, deci, centi, milli
- Chemistry has 6 more you may see
- Tera, Giga,mega, kilo, hecto, deca, base, deci, centi, milli, micro, nano, pico
- T, G,M, k,h, da, base, d, c, m, μ, n, p

- **In general, a calculated answer cannot be more precise than the least precise measurement from which it was calculated.**
  - Example: if measuring with a standard ruler and recording the measurements in cm you measurement can only have two decimal places.
  - The line below would be measured at 3.59 cm.



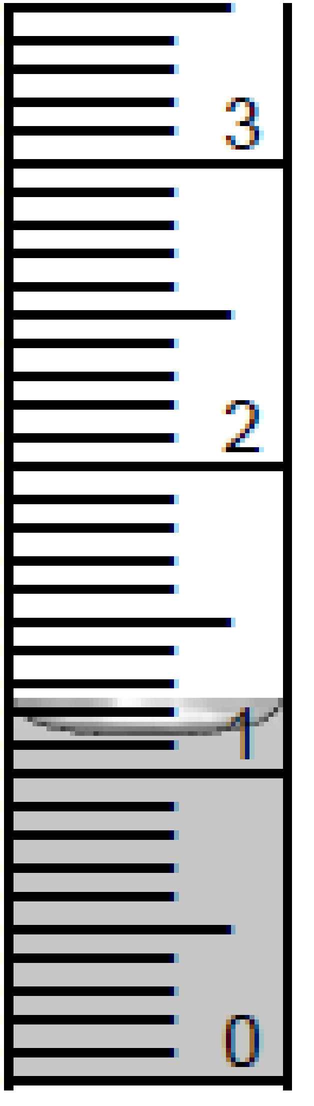
Ruler Example



- The blue line would be recorded to be 13.3_ cm long. With the _ being the estimated digit.
- 13.30 cm, 13.31 cm would both be valid measurements.
- 13.300 cm or 13.310 cm would NOT be valid
- **Read to the unit you are certain of, then estimate one more place.**

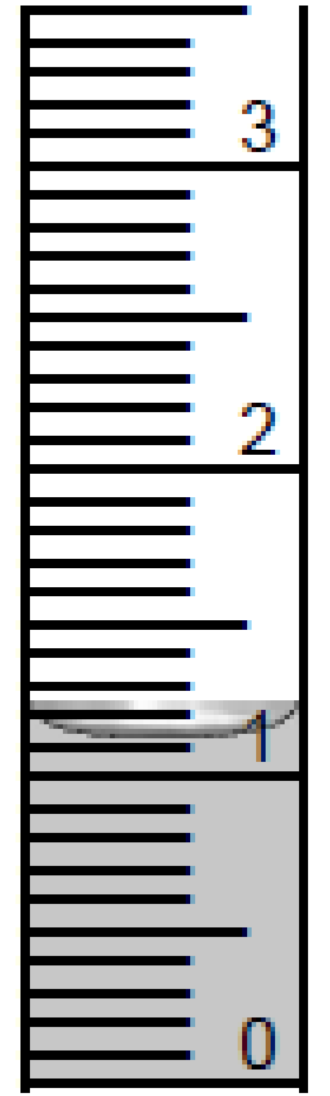
Graduated Cylinder

- In order to read the graduated cylinder correctly, it must be placed on a stable surface such as the desk top of the work area
- And you MUST be at eye level with the meniscus
- To determine the volume of liquid use the number that is directly at or below the bottom of the meniscus



Graduated Cylinder

- You must estimate one more digit that you can precisely measure.
- The graduated cylinder pictured measured in mL and 10th of a mL.
- The blue liquid would have a volume of 1.11 mL or 1.12 mL.
- A measurement of 1.110 mL or 1.1120 mL is more precise than the tool allows.
- **Read to the unit you are certain of, then estimate one more place.**



Significant Figures and Calculations

Complete Significant Figure activity to identify the significant figure rules

Significant Figures Rules

- Rules for determining whether a digit in a measured value is significant
  1. Nonzero digits are significant. 5.23 has 3 significant figures
  2. Zeros between nonzero digits are significant. 5001 has 4 significant figures
  3. Zeros at the end of a number **and** to the right of a decimal place are significant. 1.0100 has 5 significant figures
  4. Zeros in front of nonzero digits are not significant, they are only place holders. In general start counting at the 1st NON zero number 0.000099 has 2 significant figures
  5. Zeros to the left of an understood decimal point are not significant, they are only place holders. 55000 has 2 significant figures
  6. Defined quantities and counted quantities have unlimited number of significant figures. 1 ft = 12 in has ∞ sigfigs.

Significant Figures Examples

a) 2.03
b) 1.0
c) 0.00860
d) 4.50×10^{12}
e) 5.1020
f) 780
g) 780,000
h) 0.78000
i) 50.

a) 3
b) 2
c) 3
d) 3
e) 5
f) 2
g) 2
h) 5
i) 2

- When rounding first decide how many significant figures the **answer should have.**
- Next round to that number of digits , counting from the left.
- If the number to right of the last significant digit is 4 or less round down, if it is 5 or up round up.
- Make sure you don't significantly change the value of the original number. Can't round 556 to 6 must be 600
- Example: 5,274.827
  - 6 significant figures:
    - 5,274.83
  - 4 significant figures:
    - 5,275
  - 2 significant figures:
    - 5300

- Practice

A. Round 2.3567 to 3 significant figures
B. Round 56913 to 4 significant figures
C. Round 2.0132 to 2 significant figures
D. Round 5678 to 1 significant figure

- Answers

E. 2.36
F. 56910
G. 2.0
H. 6000

Significant Figures and Calculations

- With multiplication and division the calculation should be rounded to the same number of significant figures as the measurement with the LEAST number of significant figures
- Example: $2.100 \times \frac{5.32}{12} =$
- Calculator give 0.931
- 12 has only 2 significant figures so the answer must have only 2 significant figures
- Answer MUST BE 0.93

Significant Figures and Calculations

- With addition and subtraction the answer must be rounded to the same number of DECIMAL places as the value with the lease number of decimal places.
- Example: 2.450 – 14.2
- Calculator gives: -11.75
- But must be rounded to 1 decimal place so answer is -11.8

Practice

- Perform the following calculations and round to the correct number of significant figures.

	Calculator	Rounded
❑ 2.680 x 0.0051	= 0.013668	= 0.014
❑ 3.120 / 6	= 0.52	= 0.5
❑ 2.45 + 550.9	= 553.35	= 553.4
❑ 9.056 – 4.25	= 4.806	= 4.81
❑ 450 x (2.60 – 2.4865)	= 51.075	= 51
❑ (525 + 4.67)/329.68	= 1.6066185	= 1.61

Scientific Notation

- When writing very large or very small numbers, scientists use a kind of shorthand called scientific notation.
- This is a way of writing a number without so many zeros.
- Example 1: The speed of light is about 300,000,000 m/s
  - Or 3.0×10^8
- Example 2: The mass of a proton is 0.00000000000000000000000001673
  - Or $1.673 \text{ X } 10^{-24}$

All you do is move the decimal so that you **only have one** number before the decimal.

- 850,000,000.0

8 5 0 0 0 0 0 0 0 . 0

$= 8.5 \times 10^{8}$

For large numbers the exponent is positive!!

- 0.000,000,025

0 . 0 0 0 0 0 0 0 2 5

$= 2.5 \times 10^{-8}$

For small numbers the exponent is negative!!

Scientific Notation Examples

- 0.007899 = ?
  - Small number = - exponent 7.899×10^{-3}
- 898745.30 = ?
  - Large number = + exponent 8.9874530×10^{5}
- 0.00003657= ?
  - Small number = - exponent 3.657×10^{-5}
- 531120 = ?
  - Large number = + exponent 5.31120×10^{5}

Getting numbers **out of** Scientific Notation

- Look at the exponent of the number to determine if it needs to get smaller or larger
  - Positive exponent means the number get larger so the decimal moves to the right
  - Negative exponent means the number gets smaller so the decimal moves to the left
- Add zeros to fill in any “BLANK” spaces

- Example 1: 2.35×10^5
  - The exponent is positive so the number needs to get larger
  - 2 3 5 .
  - 2 3 5 0 0 0. or 235000
- Example 2: 8.68×10^{-4}
  - The exponent is negative so the number needs to get smaller
  - . 8 6 8
  - 0. 0 0 0 8 6 8 or 0.000868

Scientific Notation Examples

- 3.256×10^4
  - positive exponent = large number 3256
- 9.78×10^9
  - positive exponent = large number 978000000000
- 5.24×10^{-3}
  - Negative exponent = small number 0.00524
- 2.41×10^{-7}
  - Negative exponent = small number 0.000000241

Measurement and Density

Density

- Density is a unit of mass per unit of volume
  - SI Units of density: g/mL or g/cm^3 or Kg/m^3

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

$$d = \frac{m}{v}$$

Solving word problems

- Example 1: Robin measured the mass of a metal cube to be 25.48 g and the cube measures 3.0 cm on each side. What is the cube density?

- A block of work has a volume of 28.5 m^3 and a mass of 14.05 Kg. What is it's density?

Given	Equation	Solve
v= 28.5 m^3 m= 14.05 Kg D = ?	$D = \frac{m}{v}$	$D = \frac{14.05\ Kg}{28.5\ m^3} = 0.493 \frac{Kg}{m^3}$

- A marble has a mass of 12.48 grams and when placed in a graduated cylinder with 20.0 mL the volume increased to 24.5 mL. What is the marbles density?
  - Given: m= 12.48g d = ?

    v initial= 20.0 mL

    v final= 24.5 mL
  - Equation: $d = m/v$ $v = v_f – v_i$
  - Solve: v = 24.5 ml – 20.0 mL

    d = (12.48 g / 4.5 mL) =2.7733 g/mL

    d = 2.77 g/mL

Using Density

- Rearranging the density equation
  - First get it in a liner format by multiplying by volume
    - Density x Volume = mass
  - If wanting volume then divide by density

$$\text{Volume} = \frac{\text{mass}}{\text{volume}}$$

  - These equations can be used to find information using known density values

- The density of copper is $8.920\ g/cm^3$ if you have $52.75cm^3$ sample of copper how much does it weigh?
  - Given: $d = 8.920\ g/cm^3$

    $v = 52.75cm^3$

    $m = ?$
  - Equation: $d = m/v$ or $d(v) = m$
  - Solve: mass = $(8.920\ g/cm^3)(52.75cm^3) =$

    mass = 470.5 g

- A 250.0 g sample of lead occupied what volume? [density of lead is 11.340 g/cm^3]
  - Given: m = 250 g

    d = 11.340 g/cm^3

    v = ?
  - Equation: d = m/v or v = m/d
  - Solve: v = 250.0 g / (11.340 g/cm^3)

    v = 22.05 cm^3

Conversion Factors and Equality Statements

- Many quantities can usually be expressed different several different units
- **<u>Equality Statement</u>** shows how two (or more) different units are related
  - Example: 1 dollar = 4 quarters
- **<u>Conversion factor</u>** is a ratio of equivalent measurements.
  - Example: $\frac{100\ pennies}{1\ dollar}\ or\ \frac{1\ dollar}{100\ pennies}$
- Whenever two measurements are equivalent, a ratio of the their measurement will equal 1

- **When a measurement is multiplied by a conversion factor, the number changes, but the actual size of the quantity measured remains the same.**
  - Example: 2.0 hours = 120 minuets = 7200 seconds
- **when using conversion factors the final answer has the same number of significant figures as the starting number**

Equality Statements that you should know.

1 min =	60	seconds
1 hour =	60	minuets
1 day =	24	hours
1 week =	7	days
1 year =	52	weeks
1 year =	365	days
1 foot =	12	inches
1 yard =	3	feet

- **Dimensional Analysis** is a way to analyze and solve problems using the units of the measurements.
  - It is converting one thing to another without changing its value
  - Requires equality statements and conversion factors.
- **The key to dimensional analysis is to set it up so that the UNITS cancel.**
- **All numbers must have a unit!**
  No Naked Numbers!!!!

Steps for using dimensional analysis.

1. Write equality statement for units needed in problem
2. Write given number and unit then a fraction bar.
3. The unit you are getting rid of goes on bottom
4. The unit you are going to goes on top
5. Fill in the fraction with the values from the equality statement and solve

Example 1

- If a move is 1.48 hours long how many minuets are you in the theater?

Step 1: 60 minuets = 1 hour

Step 2: 1.48 hours -----------

Step 3: $1.48\ hours\ \frac{\quad}{Hours}$

Step 4: $1.48\ hours\ \frac{minuets}{Hours}$

Step 5: $1.48\ hours\ \frac{60\ minuets}{1\ Hour} 88.8\ minuets$

Example 2

- If you exchanged 50 nickels for quarters how many would you receive?

  step 1: 1 quarter = 5 nickels

  step 2-4: $50\ nickels\ \dfrac{quarter}{nickels} =$

  step 5: $50\ nickels\ \dfrac{1\ quarter}{5\ nickels} = 10\ quarters$

Example 3 ~~ two step problem

- If a movie is 1.75 hours long how many seconds are you in the theater.
  - We don't have one equality statement that relates seconds and hours so we used two

Step 1: 1 hour= 60 minuets, 1 minuet = 60 seconds

Step 2-3: $1.75\ Hours\,\frac{\qquad}{Hour}\,\frac{\qquad}{\qquad}$

Step 2-4 $1.75\ Hours\,\frac{minuet}{Hour}\,\frac{seconds}{minuet}$

Step 5: $1.75\ Hours\,\frac{60\ minuet}{1\ Hour}\,\frac{60\ seconds}{1\ minuet} = 6300\ seconds$

Example 4 ~~ two step problem

- An child just turned 5 years old, how many hours old is the child?

step 1: 1 year = 365 days, 1 day = 24 hours

step 2-4:

$$5\ years\ \frac{days}{year}\ \frac{hours}{day} =$$

step 5:

$$5years\ \frac{365\ days}{1\ year}\ \frac{24\ hours}{1\ day} = 43800\ hours$$

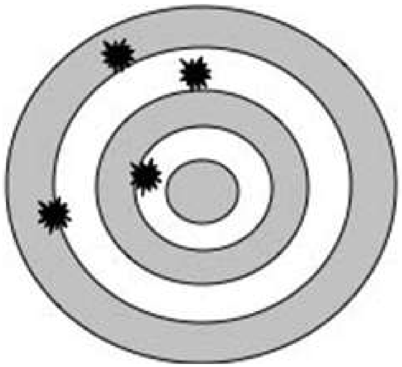
Remember
NO NAKED NUMBERS!!!!

Show ALL units
at every step.
Round at the end.

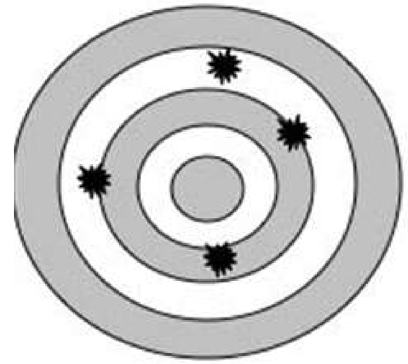
Limits of Measurement

- **Precision** is a gauge of how exact a measurement is.
- Precise measurements are close to each other
- MUST have more than one measurement

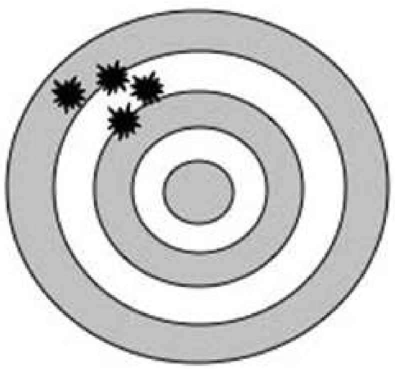
- **Accuracy** is the closeness of a measurement to the actual value of what is being measured
- An accurate measure is close to the true or expected value
- MUST have true or expected value



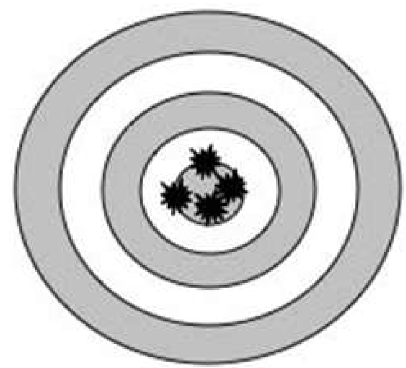
NOT Accurate (not near center)
NOT precise (not near each other)



Accurate (closer to center)
NOT precise (not near each other)



NOT Accurate (not near center)
Precise (close to each other)



Accurate (Near center)
Precise (close to each other)

Sally	Annie	Travis	Jeff
1.95 g/cm^3	2.69 g/cm^3	3.12 g/cm^3	2.71 g/cm^3
1.89 g/cm^3	2.73 g/cm^3	2.70 g/cm^3	
1.92 g/cm^3	2.65 g/cm^3	2.25 g/cm^3	

To the right is the data collected by students during a lab.
Actual Density of Aluminum is 2.70 g/cm^3

1. Which students data is accurate and precise?
2. Which students data is accurate but NOT precise?
3. Which students data is NOT accurate but IS precise?
4. Which students data is NEITHER accurate nor precise?

1. Annie
2. Jeff
3. Sally
4. Travis

Types of measurement

- Qualitative measurement based on some quality or characteristic
  - Deals with descriptions.
  - Data can be observed but not measured.
    - Colors, textures, smells, tastes, appearance, beauty, etc.
- **Qualit**ative → **Qualit**y
    - Blue liquid, soft fabric, cold room

Types of measurement

- Quantitative measurement is something that is measurable in quantity
  - Deals with numbers.
  - Data which can be measured.
  - distance, volume, mass, speed, time, temperature, cost, ages, etc.
- **Quantit**ative → **Quantit**y
    - 25.0 g, 48 mL, 3 days, 45 miles