<u>Scientific Notation</u>

Scientific notation is used to express very large or small numbers.

Ex. 6.02×10^{23}

The decimal goes after the first number

The exponent (or power of 10) show shows how far the decimal point moved

A *positive* exponent represents a number GREATER than 1.

5.243×10^{3}

A *negative* exponent represents a number LESS than 1.

5.243×10^{-3}

Scientific Notation Practice

Convert each of the following numbers to standard notation.

- a. 3.4×10^5 b. 2.8×10^3 c. 6.1×10^{-2} d. 7.8×10^{-4}

- Convert each of the following numbers to scientific notation.
- a. 5600 b. 0.00921 c. 980,000 d. 0.0000073

Scientific Notation Practice Ans.

- Convert each of the following numbers to standard notation.
- a. 3.4×10^5 b. 2.8×10^3 c. 6.1×10^{-2} d. 7.8×10^{-4}

340,000 2800 0.061 0.00078

- Convert each of the following numbers to scientific notation.
- a. 5600 b. 0.00921 c. 980,000 d. 0.0000073

5.6×10^3 9.21×10^{-3} 9.8×10^5 7.3×10^{-6}

SI Units - **S**ystème **I**nternational d'Unites, or International System of Units

MOST countries use SI units.

SI units are used in science.

BASE UNITS

Quantity	**Base Unit**	**Abbreviation**
Length	meter	m
Mass	kilogram	kg
Time	seconds	sec
Temperature	Kelvin	K
Amount	moles	mol

SI UNITS

What is Kelvin?

Kelvin = $^{o}C + 273$

Written as K only (no degree "^{o}" sign)

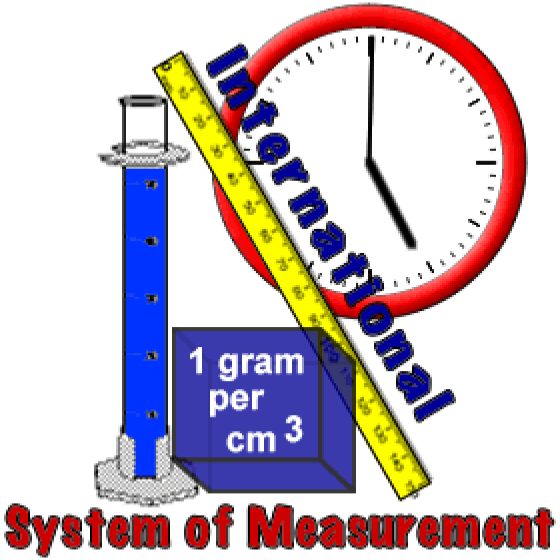
Ex. $15^{o}C + 273 = 288$ K

Derived Units Combination of base units.

Ex. Volume = length x width x height = liters (l)

Density = mass/volume = g/ml

SI UNITS



UNITS MATTER!!

ALL MEASUREMENTS MUST HAVE:

A MAGNITUDE (number)

UNITS (label)

Ex. 25 kg, 2.03×10^{13} m, 3.04×10^{-2} mol

SI UNITS - Prefixes *(Need to memorize!)*

King	Henry	Died	By	Drinking	Chocolate	Milk	Monday	Night
Kilo	Hecto	Deka	BASE	Deci	Centi	Milli	Micro	Nano
1000 $1x10^{3}$	100 1×10^{2}	10	1 How many (g,L,m) in each unit)	0.1	0.01 1×10^{-2}	0.001 $1x10^{-3}$	1×10^{-6}	$1 x10^{-9}$

Density

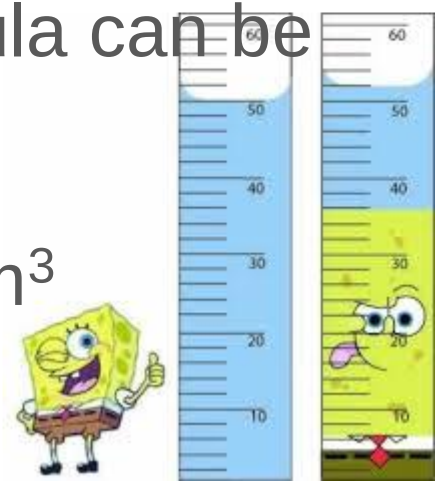
The density of an object will remain the same, regardless of the amount of the object.

Density = $\dfrac{\textbf{Mass}}{\textbf{Volume}}$

Ex. A liquid has a density of 0.87 g/mL and a mass of 0.025 kg. What is its volume?

Volume:

Objects with a familiar shape → a formula can be used to calculate volume.



Ex. Cube = length X width X height = cm^3

Odd shaped objects--- use **water displacement method…**

Density practice problems - answers

1. A block of aluminum occupies a volume of 15.0 mL and weighs 40.5 g. What is its density?

D = 40.5/15 **d = 2.7 g/mL**

2. What is the weight of the ethyl alcohol that exactly fills a 200.0 mL container? The density of ethyl alcohol is 0.789 g/mL.

.789 = m/200 **m = 157.8 g**

3. A rectangular block of copper metal weighs 1896 g. The dimensions of the block are 8.4 cm by 5.5 cm by 4.6 cm. From this data, what is the density of copper?

V = 8.4 x 5.5 x 4.6 **d = 1896/212.52**

V = 212.52 cm^3 **d = 8.9 g/cm^3**

4. 28.5 g of iron shot is added to a graduated cylinder containing 45.50 mL of water. The water level rises to the 49.10 mL mark, from this information, calculate the density of iron. **V = 49.10-45.5** **d = 28.5/3.6**

V = 3.6 mL **d = 7.9 g/mL**

SI UNITS - Conversion

Ex. *How many grams are in 200 kg?*

To convert from one unit to the other,

start with what you have. **200 kg**

Create a conversion factor:

$$200 \text{ kg} \times \frac{1 \times 10^3 \text{g}}{1 \text{ kg}} = 200{,}000 \text{ g}$$

k	**kilo**	**10^3**
h	**hecto**	**10^2**
d	**deka**	**10^1**
B	**BASE**	**10**
d	**deci**	**10^{-1}**
c	**centi**	**10^{-2}**
m	**milli**	**10^{-3}**
μ	**micro**	**10^{-6}**
n	**nano**	**10^{-9}**

Conversion Methods

FYI - In class, you will see me do conversions similar to the example below.

$$56.78\text{ cm} \times \frac{1 \times 10^{-2}\text{ m}}{1\text{ cm}} \times \frac{1\text{ km}}{1 \times 10^{3}\text{ m}} = 5.678 \times 10^{-4}\text{ km}$$

You can also use the T-bar method.

56.78 cm	1×10^{-2} m	1 km	$= 5.678 \times 10^{-4}$ km
	1 cm	1×10^{3} m	

Both methods yield the same result, it is just a different way to set up your conversions. Both methods are acceptable. It is your preference as to which one to use!

Percent Error - Example

In lab, you measure the mass of your sample to be 6.2 g. Reliable sources indicate that the actual mass of the sample should have been 8.3 g. What was your percent error?

- $$\frac{|6.2 - 8.3|}{8.3} \times 100 = 25.3\% \text{ error}$$

Percent Error

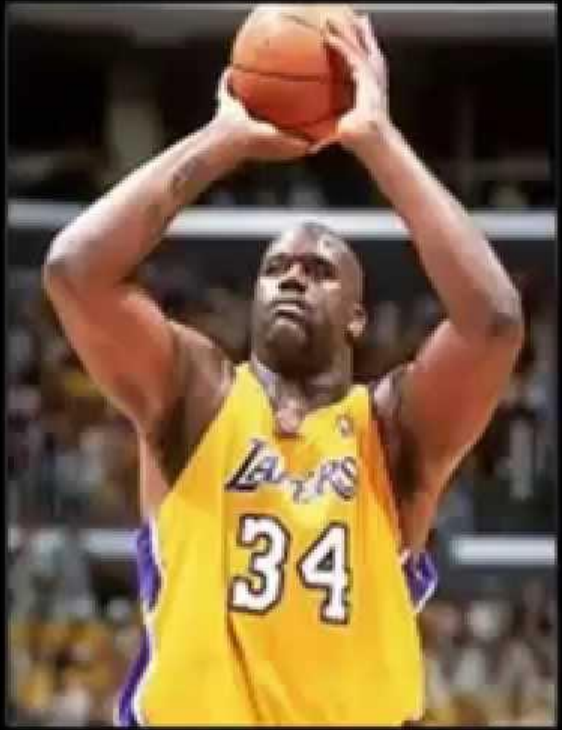
Percent Error – A way of expressing accuracy

$$= \frac{|\text{experimental value} - \text{accepted value}|}{\text{accepted value}} \times 100\%$$

* **Accepted value** (or true value)– correct value based on reliable references
* **Experimental value** – value measured in the lab

Shaq was precise with free throws. He just wasn't accurate.

Accurate = Correct
Precise = Consistent



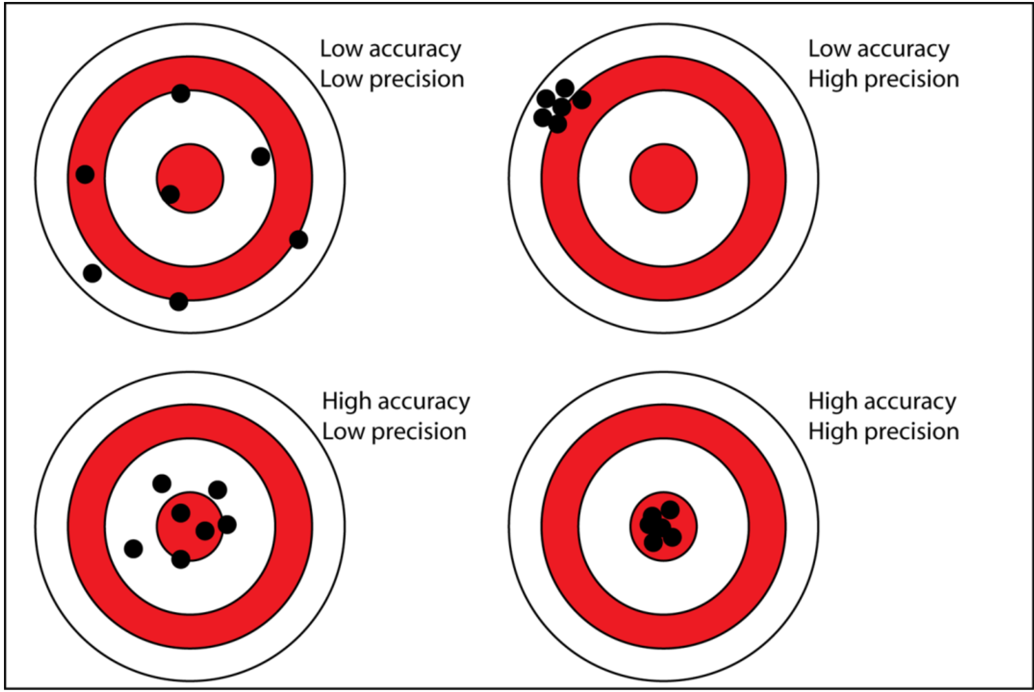
Precision vs Accuracy

Accuracy:

How close a measurement comes to the correct (true) value

Precision:

How close a series of measurements are to one another



The Mole

What is a mole??

A unit of measurement

How many molecules are in a substance

1 mol = 6.022×10^{23}

This is also called **Avogadro's Number**

Moles - Answers

1. How many molecules are in 3.2 moles of water

$$3.2 \text{ mol} \times \frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = \mathbf{1.92 \times 10^{24} \text{ molecules}}$$

2.) How many moles are in 5.34×10^{24} atoms of Copper?

$$5.34 \times 10^{24} \text{ atoms} \times \frac{1 \text{ mol}}{6.02 \times 10^{23}} = \mathbf{8.87 \text{ moles}}$$

3.) How many molecules are in 2.43 mol of CO_2?

1.46×10^{24} molecules

4.) Calculate the number of moles in 7.12×10^{22} atoms of Mg.

0.118 moles

Significant Figures

- Counting Sig Figs
  - Count all numbers EXCEPT:
    - Leading zeros -- 0.0025
    - Trailing zeros *without* a decimal point -- 2,500

Significant Figures

Counting Sig Fig Examples

1. 23.50	4 sig figs
2. 402	3 sig figs
3. 5,280	3 sig figs
4. 0.080	2 sig figs

Significant Figures

- Calculating with Sig Figs
  - <u>Multiply/Divide</u> - Multiply or divide as usual, then the # with the fewest sig figs determines the # of sig figs in the answer.

$(13.91 g/cm^3)(23.3 cm^3) =$ **3 SF**

4 SF **3 SF**

324 g

324.103g

Significant Figures

- Calculating with Sig Figs (con't)
  - <u>Add/Subtract</u> - You would add (or subtract) the numbers as usual, but then you would round the answer to the same decimal place as the least-accurate number.

$$\begin{array}{r} 3.75 \text{ mL} \\ + \; 4.1 \text{ mL} \\ \hline 7.85 \text{ mL} \end{array} \rightarrow 7.9 \text{ mL} \qquad \begin{array}{r} 224 \text{ g} \\ + \; 130 \text{ g} \\ \hline 354 \text{ g} \end{array} \rightarrow 350 \text{ g}$$