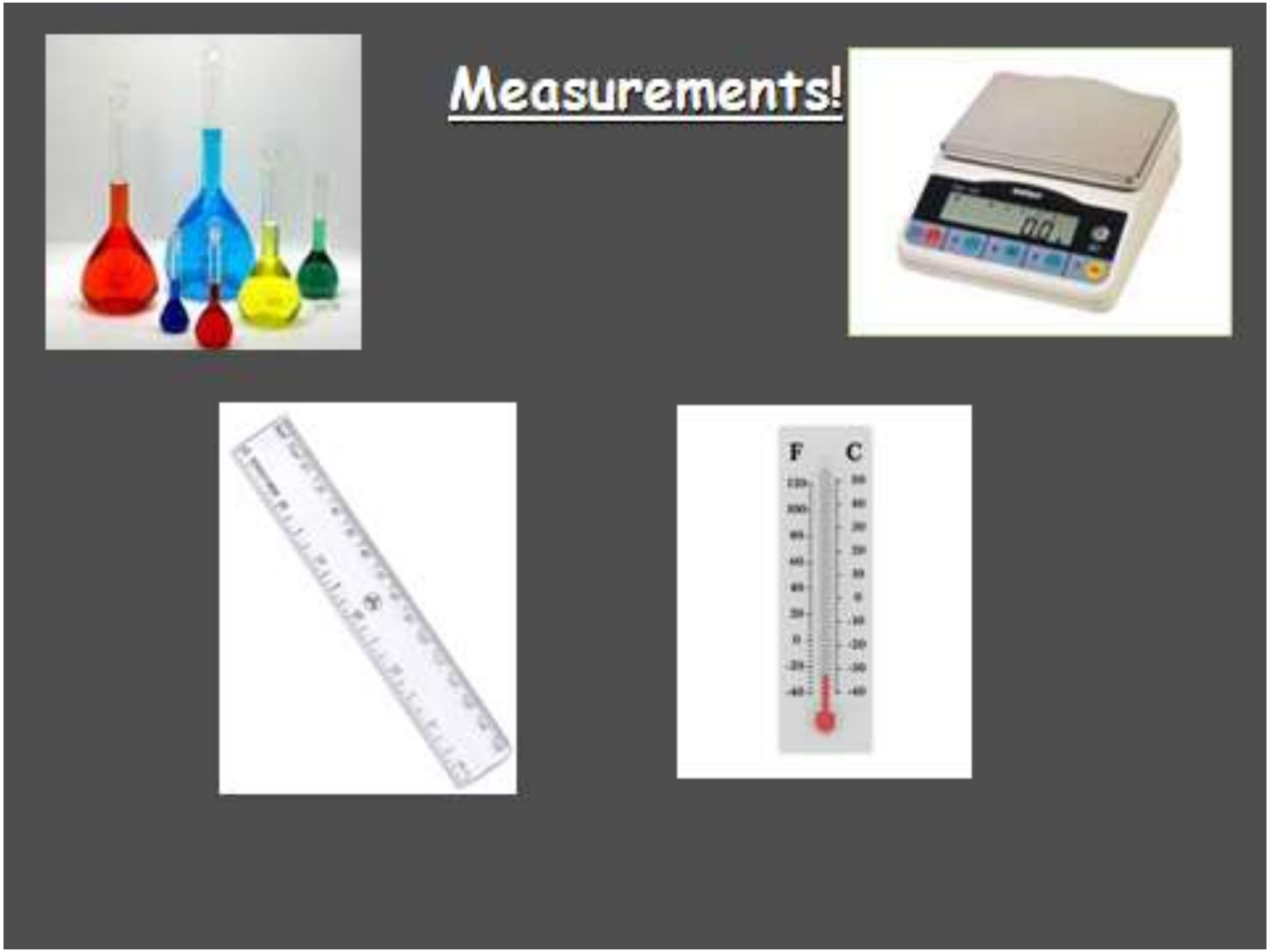


What are significant digits? Why do we need them when measuring?

Qualitative Data Data

Overview:

Deals with descriptions.

Data can be observed but not measured.

Colors, textures, smells, tastes, appearance, beauty, etc.

Qualitative → Quality

Quantitative Data

Overview:

Deals with numbers.

Data which can be measured.

Length, height, area, volume, weight, speed, time, temperature, humidity, sound levels, cost, members, ages, etc.

Quantitative → Quantity

With Experimentation and Observation comes Measurements! (Quantitative DATA)

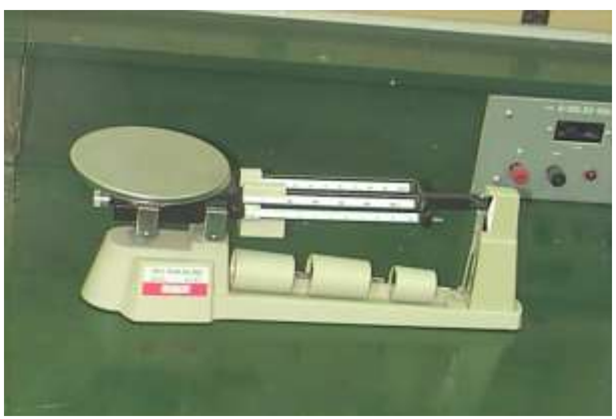
- Uncertainty in Measurement

- A digit that must be estimated is called uncertain. A measurement always has some degree of uncertainty.

Why Is there Uncertainty?

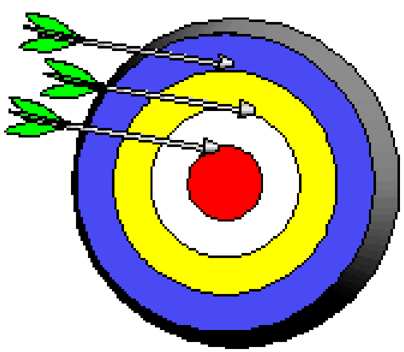
- Measurements are performed with instruments
- No instrument can read to an infinite number of decimal places

Which of these balances has the greatest uncertainty in measurement?

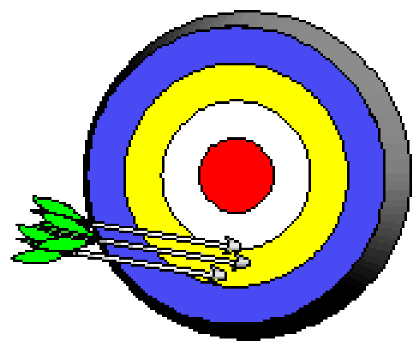


Precision and Accuracy

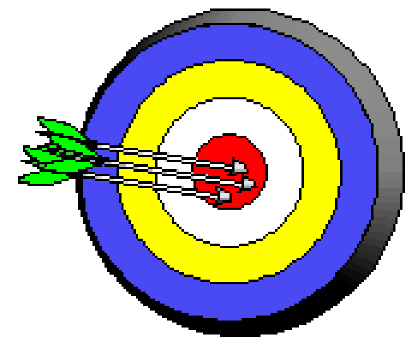
- Accuracy refers to the agreement of a particular value with the true value. (did you hit the bulls eye?)
- Precision refers to the degree of agreement among several measurements made in the same manner. (are your arrows grouped together?)



Neither accurate nor precise



Precise but not accurate



Precise AND accurate

You Try Problem: Who to Hire!

You are in charge of hiring a new chemist at a cola plant. The chemist is in charge of having the correct phosphoric acid concentration in the drink.

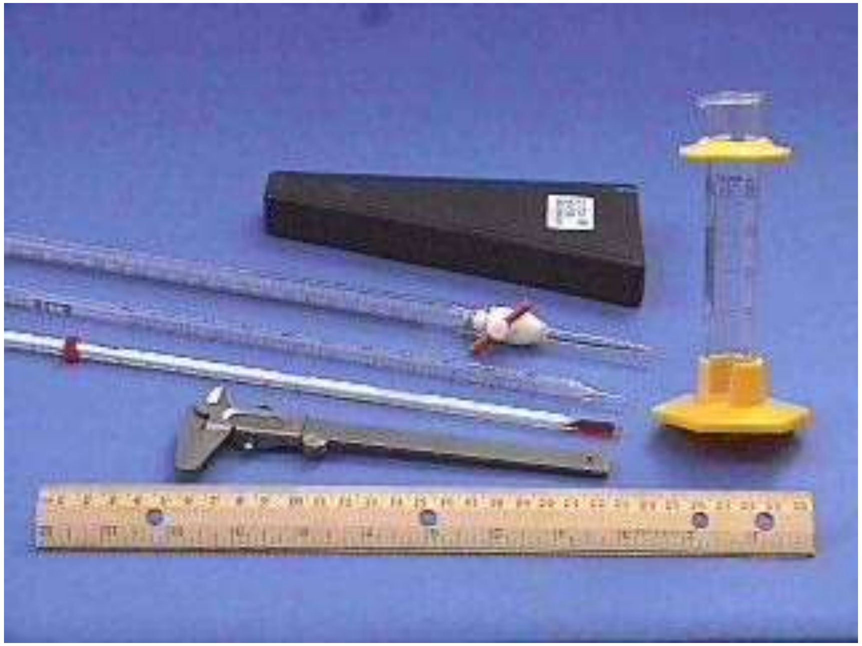
Three Chemists come in and perform an experiment. The data is collected below:

True Value 25.00 g +/- 2.00%

	Amy	Billy	Charlie
Trial 1	25.14	26.00	25.00
Trail 2	24.95	26.20	left
Trial 3	24.77	25.85	left

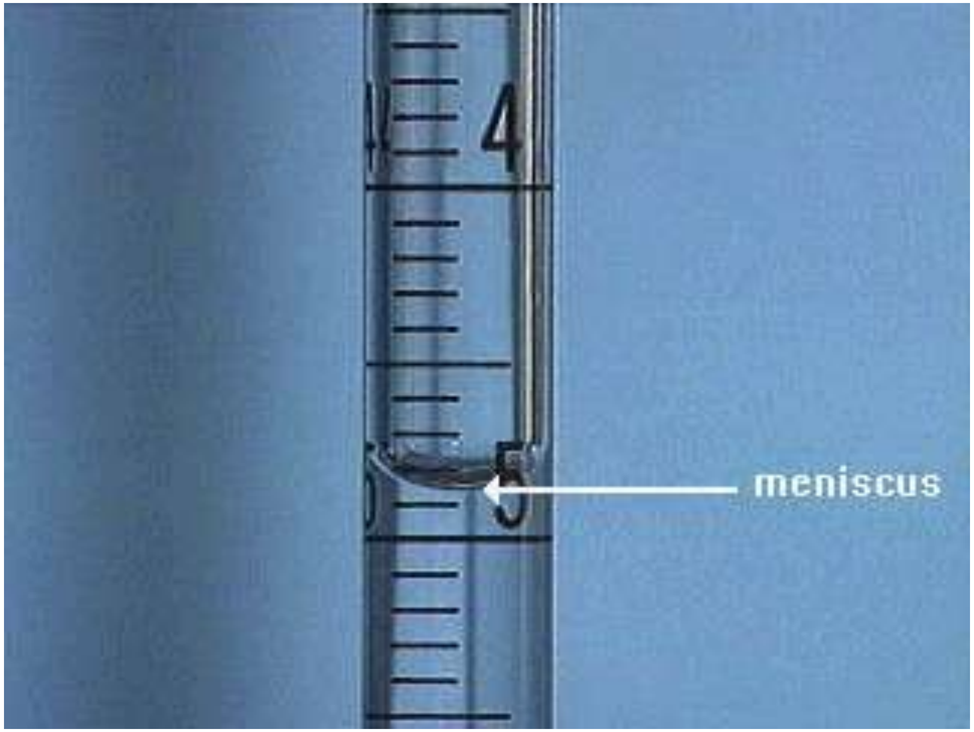
Measuring

- **Volume**
- **Temperature**
- **Mass**



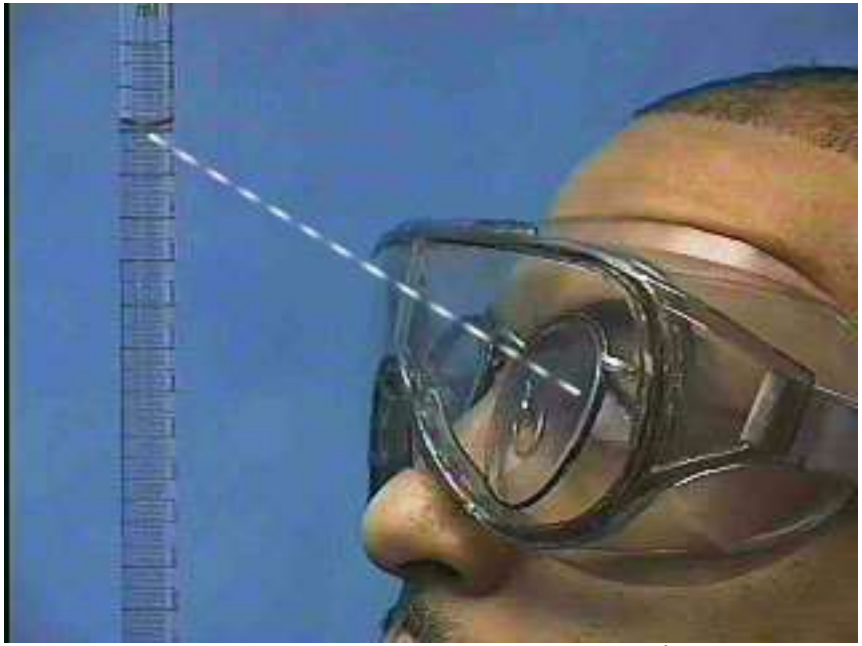
Reading the Meniscus

Always read volume from the bottom of the meniscus. The meniscus is the curved surface of a liquid in a narrow cylindrical container.

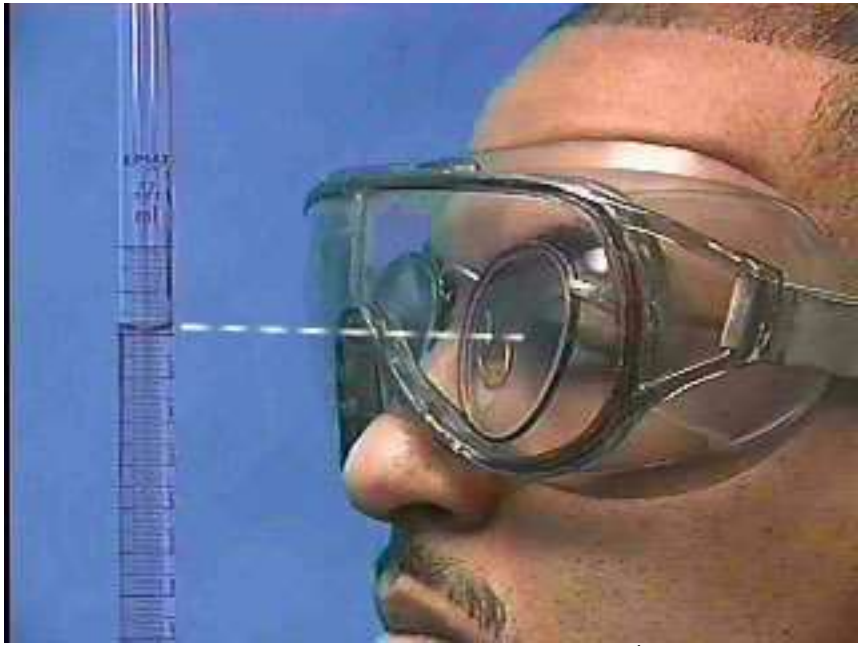


Try to avoid parallax errors.

Parallax errors arise when a meniscus or needle is viewed from an angle rather than from straight-on at eye level.



Incorrect: viewing the meniscus
from an angle



Correct: Viewing the meniscus
at eye level

Graduated Cylinders

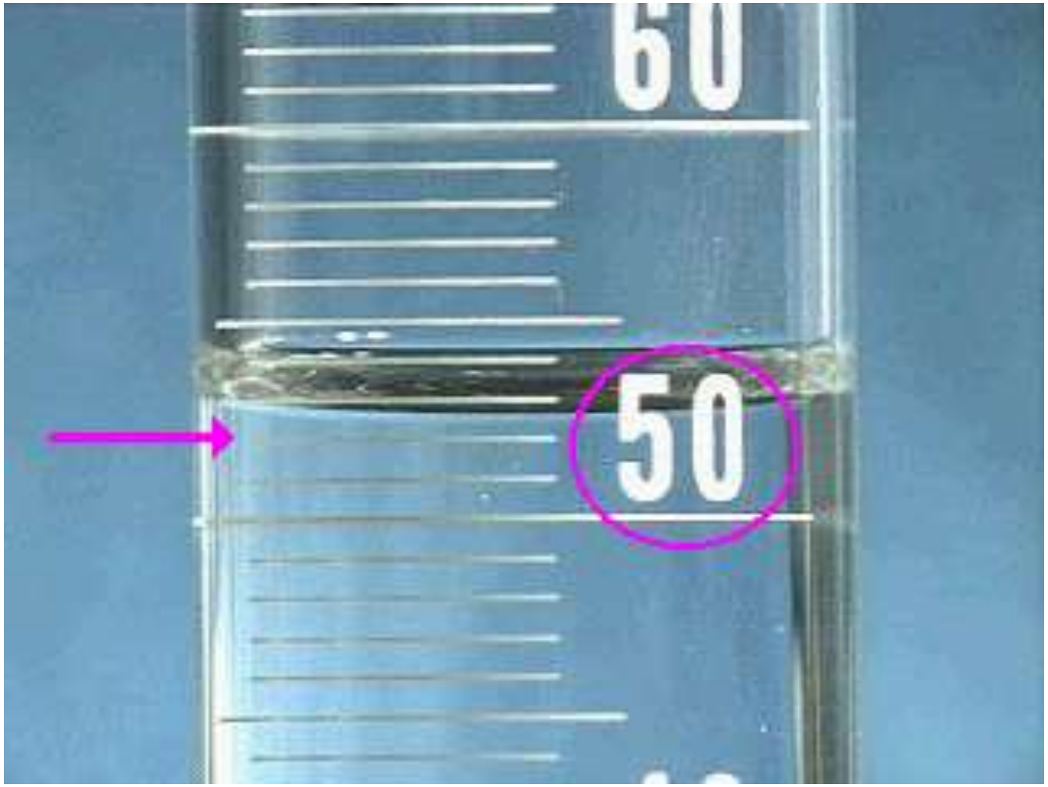
The glass cylinder has etched marks to indicate volumes, a pouring lip, and quite often, a plastic bumper to prevent breakage.

Measuring Volume

- **Determine the volume contained in a graduated cylinder by reading the bottom of the meniscus at eye level.**
- **Read the volume using all certain digits and one uncertain digit.**
  - **Certain digits are determined from the calibration marks on the cylinder.**
  - **The uncertain digit (the last digit of the reading) is estimated.**

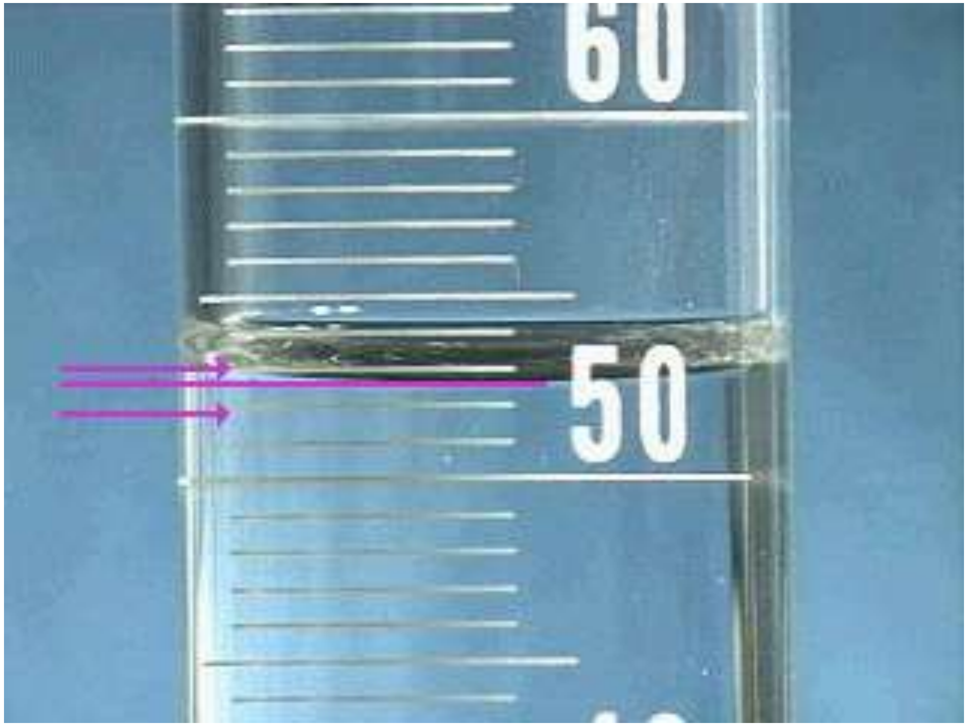
Use the graduations to find all certain digits

There are two unlabeled graduations below the meniscus, and each graduation represents 1 mL, so the certain digits of the reading are… 52 mL.



Estimate the uncertain digit and take a reading

The meniscus is about eight tenths of the way to the next graduation, so the final digit in the reading is 0.8 mL .



The volume in the graduated cylinder is 52.8 mL.

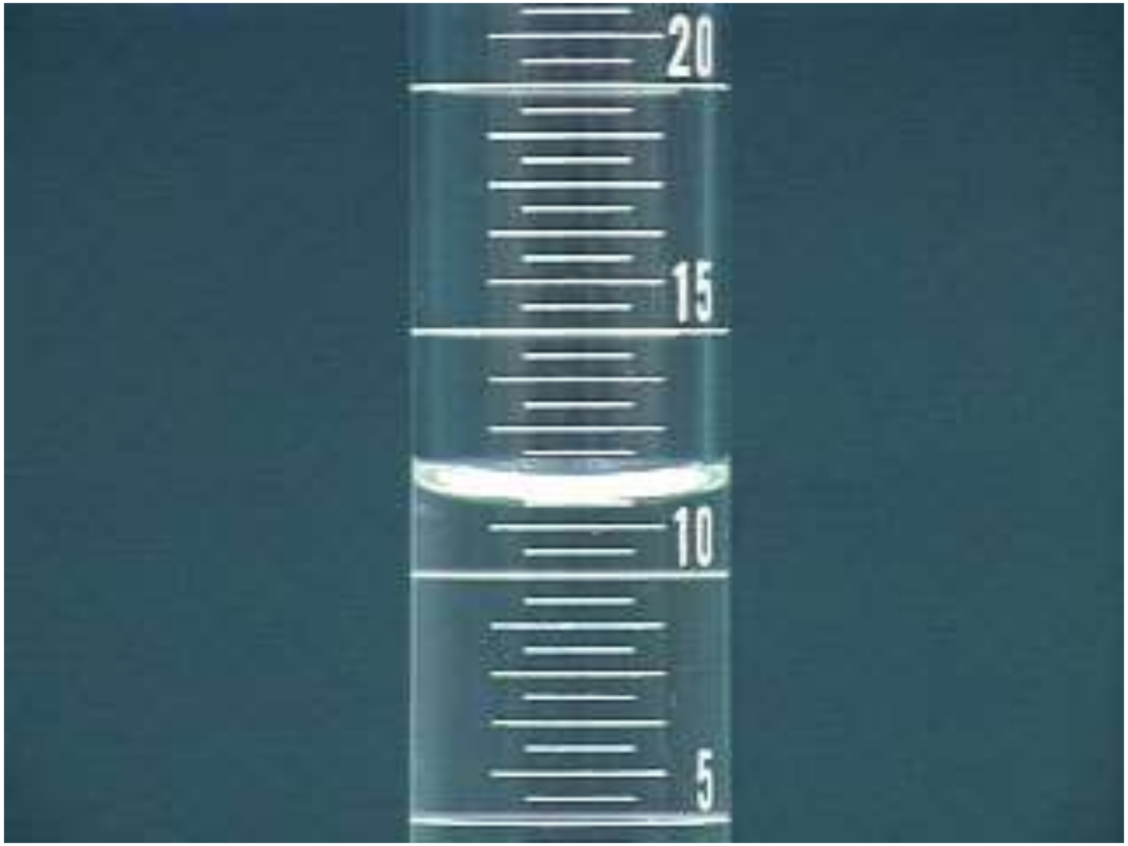
10 mL Graduate

What is the volume of liquid in the graduate?

6 . 6 2 mL

25mL graduated cylinder

What is the volume of liquid in the graduate?

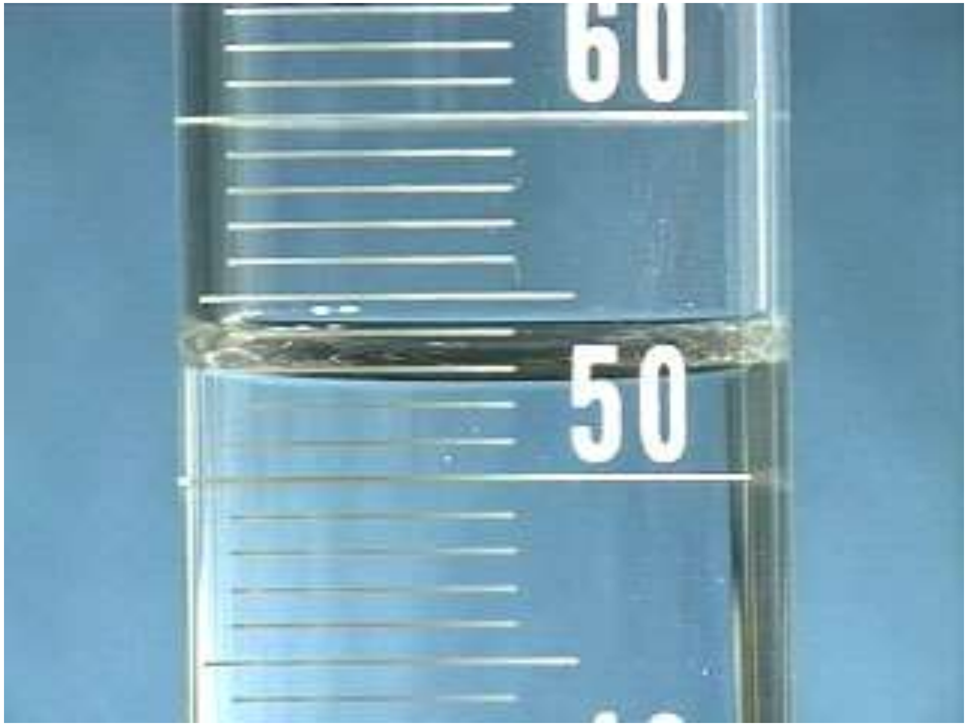


1 1 . 5 mL

100mL graduated cylinder

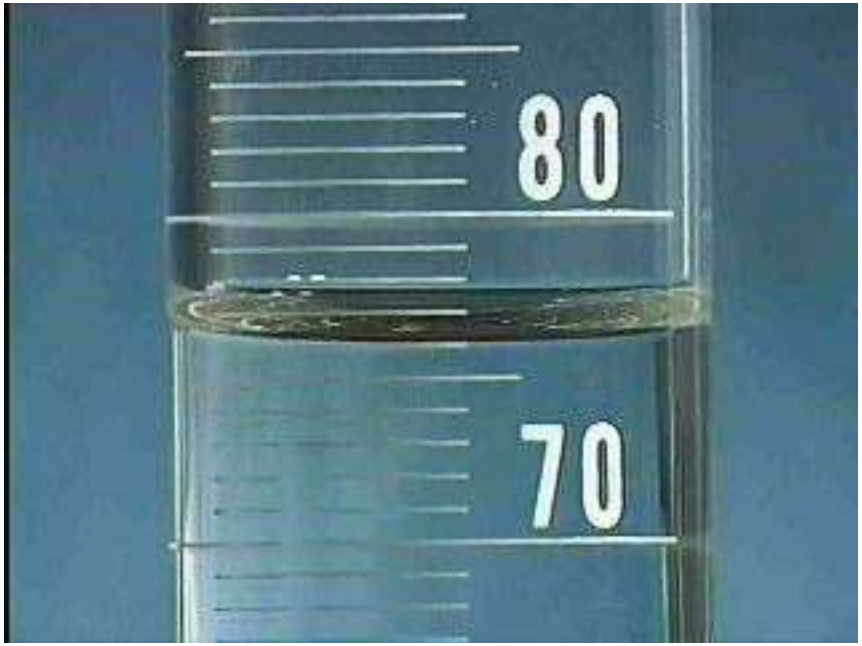
What is the volume of liquid in the graduate?

5 2 . 7 mL



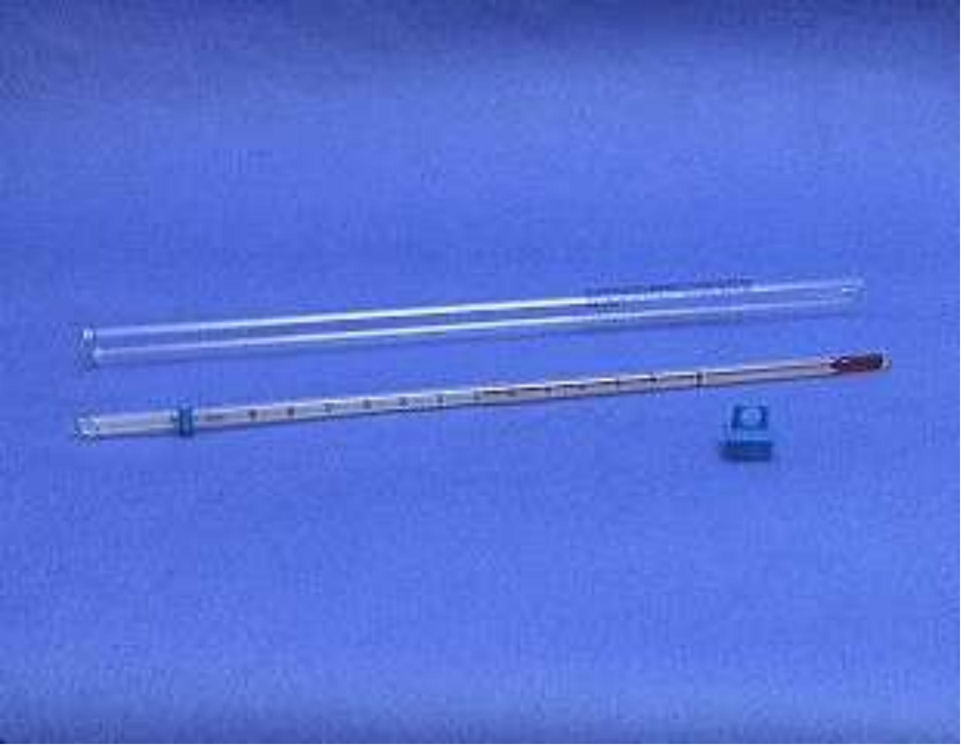
Self Test

Examine the meniscus below and determine the volume of liquid contained in the graduated cylinder.



The cylinder contains:

7 6 . 0 mL

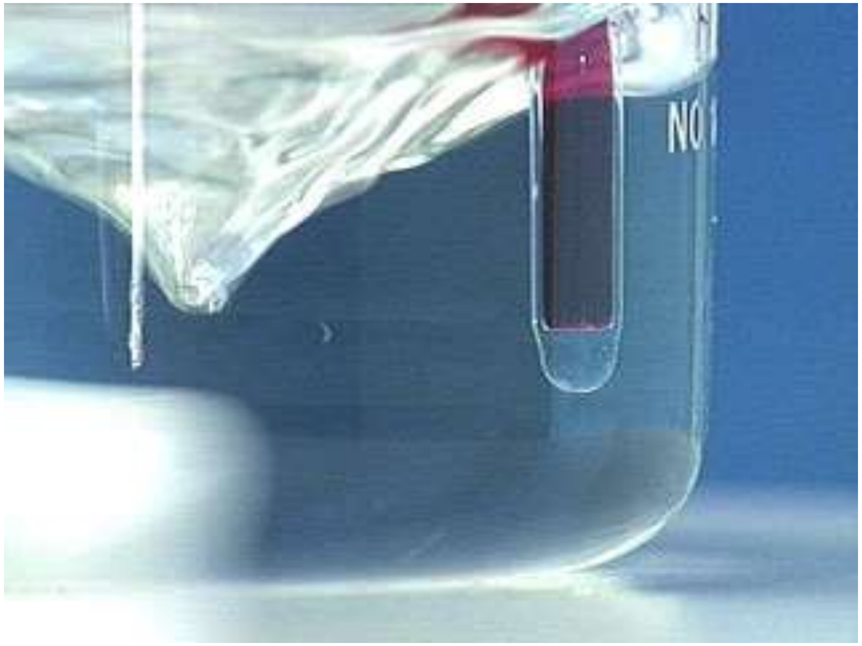


The Thermometer

- Determine the temperature by reading the scale on the thermometer at eye level.
- Read the temperature by using all certain digits and one uncertain digit.
- Certain digits are determined from the calibration marks on the thermometer.
- The uncertain digit (the last digit of the reading) is estimated.
- On most thermometers encountered in a general chemistry lab, the tenths place is the uncertain digit.

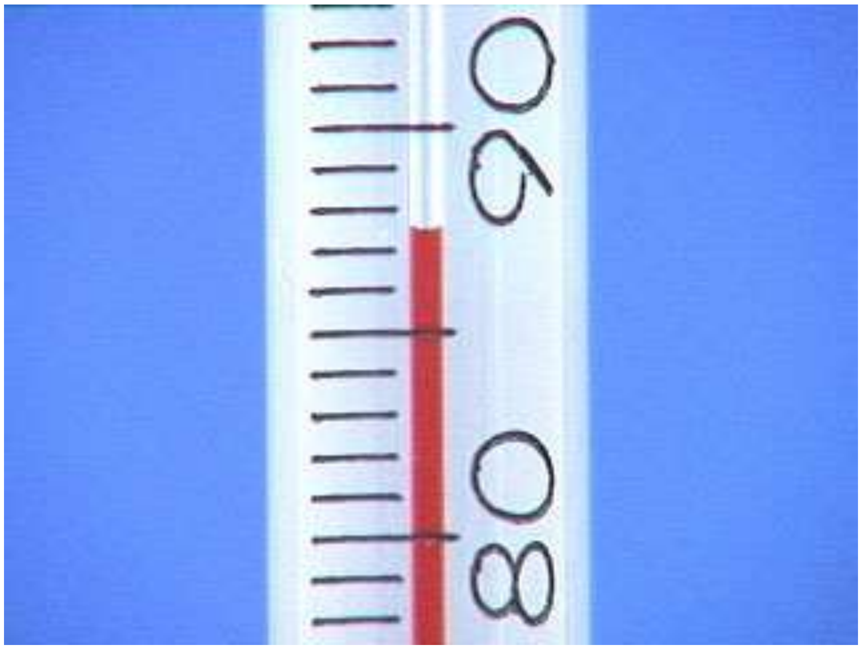
Do not allow the tip to touch the walls or the bottom of the flask.

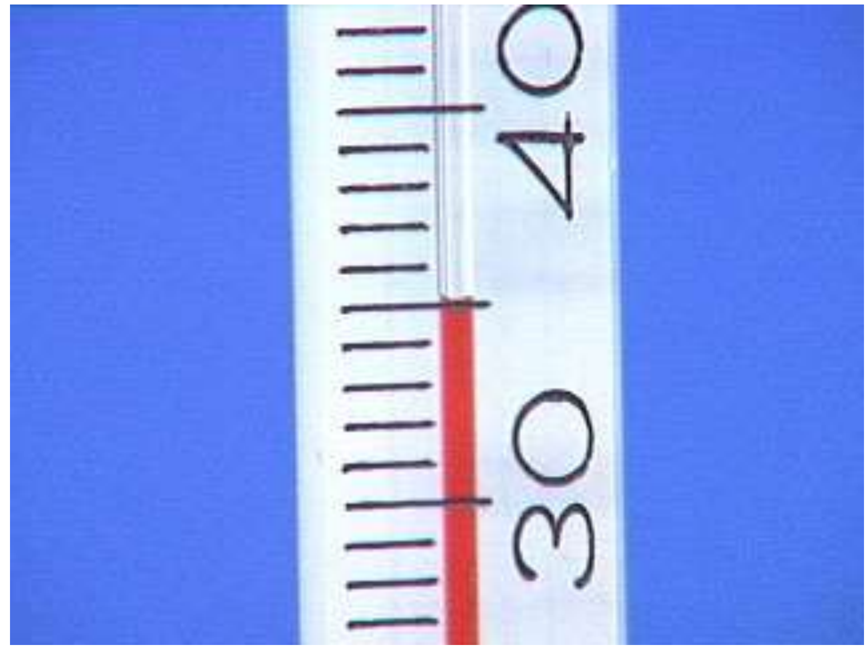
If the thermometer bulb touches the flask, the temperature of the glass will be measured instead of the temperature of the solution. Readings may be incorrect, particularly if the flask is on a hotplate or in an ice bath.



Reading the Thermometer

Determine the readings as shown below on Celsius thermometers:

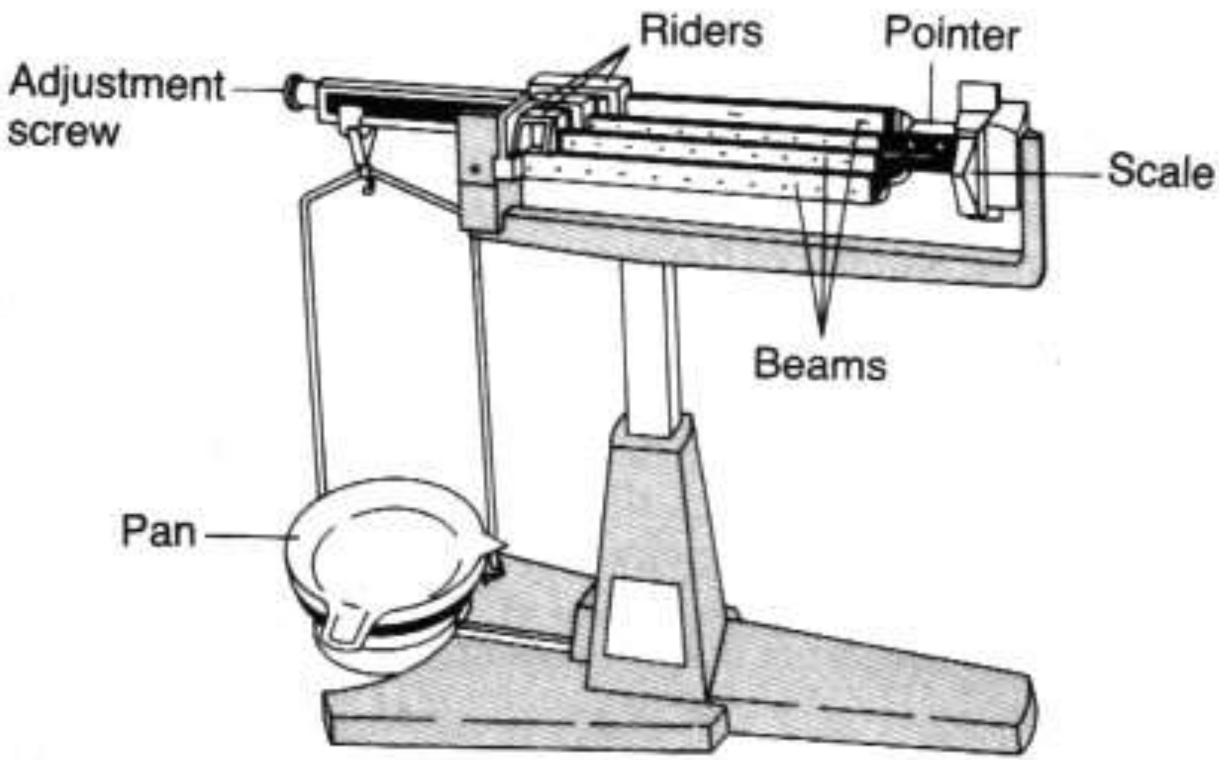




8 7 . 4 °C

3 5 . 0 °C

Measuring Mass - The Beam Balance



Our balances have 4 beams – the uncertain digit is the thousandths place (_ _ _ . _ _ X)

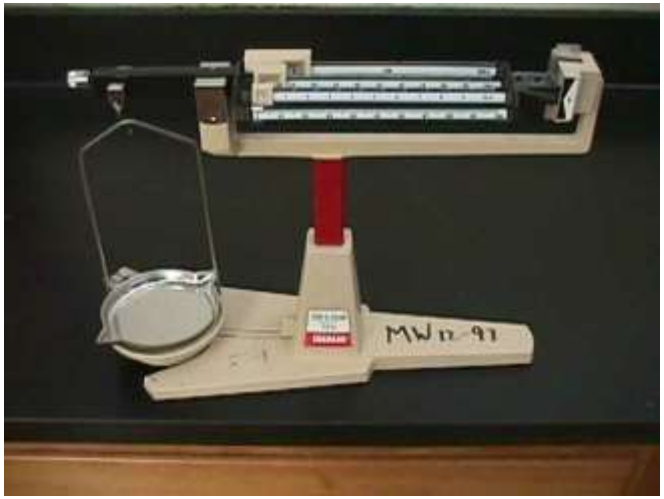
Balance Rules

In order to protect the balances and ensure accurate results, a number of rules should be followed:

- **Always check that the balance is level and zeroed before using it.**
- **Never weigh directly on the balance pan. Always use a piece of weighing paper to protect it.**
- **Do not weigh hot or cold objects.**
- **Clean up any spills around the balance immediately.**

Mass and Significant Figures

- Determine the mass by reading the riders on the beams at eye level.
- Read the mass by using all certain digits and one uncertain digit.

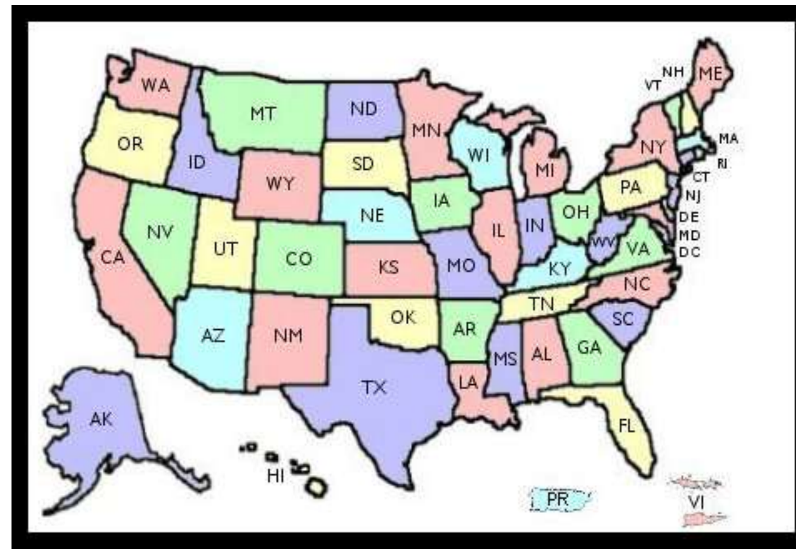


- The uncertain digit (the last digit of the reading) is estimated.
- On our balances, the thousandths place is uncertain.

How do we determine significant digits when we are given a measurement?

- We are going to use the:

Pacific /Atlantic rule



Rule ONE: All non-zero numbers are significant

- When the decimal is <u>A</u>bsent start on the right of the number and slide until you reach a non-zero number then count all the numbers that follow the non-zero number (even "sandwiched" zeros)
- 120,000 has 2 S.F.
- 101,000 has 3 S.F.

- When the decimal is <u>P</u>resent start on the left of the number and slide until you hit a non-zero number and count all the numbers that follow that non-zero number
- 0.000509 has 3 S.F.
- 120.000 has 6 S.F.

How many significant figures in each of the following?

Sig Fig Practice #1

1.0070 m → 5 sig figs

17.10 kg → 4 sig figs

100,890 L → 5 sig figs

3.29×10^3 s → 3 sig figs

0.0054 cm → 2 sig figs

3,200,000 → 2 sig figs

Rules for Significant Figures in Mathematical Operations

- Multiplication and Division: # sig figs in the result equals the number in the least precise measurement used in the calculation. (Don't Report What the Calculator has !!!! Report to the lowest SF)

  - 6.38 x 2.0 = 12.76
    - 12.76 must be rounded to have only 2 numbers →
  - 13 (2 sig figs)

Sig Fig Practice #2

Calculation	Calculator says:	Answer
3.24 m x 7.0 m	22.68 m^2	23 m^2
100.0 g ÷ 23.7 cm^3	4.219409283 g/cm^3	4.22 g/cm^3
0.02 cm x 2.371 cm	0.04742 cm^2	0.05 cm^2
710 m ÷ 3.0 s	236.6666667 m/s	240 m/s
1818.2 lb x 3.23 ft	5872.786 lb·ft	5870 lb·ft
1.030 g ÷ 2.87 mL	2.9561 g/mL	2.96 g/mL

Rules for Significant Figures in Mathematical Operations

- Addition and Subtraction: The number of decimal places in the result equals the number of decimal places in the least precise measurement. Line up decimals and report to the lowest place holder
- $$\begin{array}{rr} & 6.8 \\ + & 11.934 \\ \hline & 18.7311 \end{array}$$

18.7

- (1 number in the first decimal place)

Sig Fig Practice #3

Calculation	Calculator says:	Answer
3.24 m + 7.0 m	10.24 m	10.2 m
100.0 g - 23.73 g	76.27 g	76.3 g
0.02 cm + 2.371 cm	2.391 cm	2.39 cm
713.1 L - 3.872 L	709.228 L	709.2 L
1818.2 lb + 3.37 lb	1821.57 lb	1821.6 lb
2.030 mL - 1.870 mL	0.16 mL	0.160 mL

Scientific Notation

In science, we deal with some very **LARGE** numbers:

1 mole = 602000000000000000000000

In science, we deal with some very **SMALL** numbers:

Mass of an electron =
0.00000000000000000000000000000009 1 kg

Imagine the difficulty of calculating the mass of 1 mole of electrons!

0.0000000000000000000000000000000091 kg

x 6020000000000000000000000

????????????????????????????????????

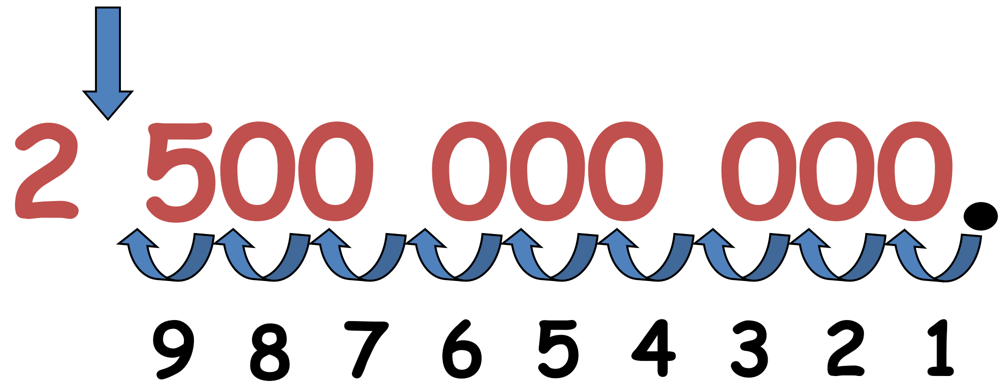
Scientific Notation:

A method of representing very large or very small numbers in the form:

$M \times 10^n$

- M is a number between 1 and 9
- n is an integer

- A positive integer = large numbers
- A negative integer= small numbers



Step #1: Insert an understood decimal point

Step #2: Decide where the decimal must end up so that one number is to its left

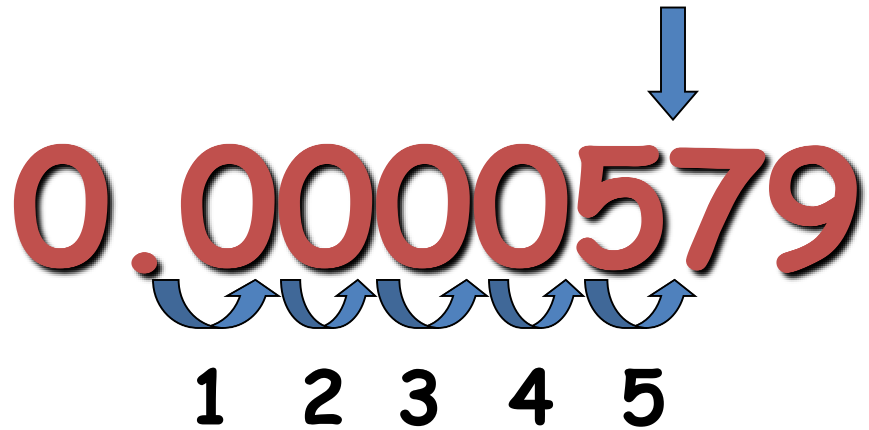
Step #3: Count how many places you bounce the decimal point-

be careful not to introduce more S.F. than original number!

Step #4: Re-write in the form $M \times 10^n$

$$2.5 \times 10^9$$

The exponent is the number of places we moved the decimal.



Step #2: Decide where the decimal must end up so that one number is to its left

Step #3: Count how many places you bounce the decimal point –again be mindful of S.F.

Step #4: Re-write in the form $M \times 10^n$

$$5.79 \times 10^{-5}$$

The exponent is negative because the number we started with was less than 1.

PERFORMING CALCULATIONS IN SCIENTIFIC NOTATION

ADDITION AND SUBTRACTION

Review:

Scientific notation expresses a number in the form:

$$M \times 10^n$$

$1 \leq M < 10$

n is an integer

$$
\begin{array}{r}
4.0 \times 10^{6} \\
\times\ 3.0 \times 10^{8} \\
\hline
1.2 \times 10^{15}
\end{array}
$$

1. **Find your EE Exp button on your calculator!!**

2. Plug in 4.0 HIT EE/EXP then the exponent 6 (DO NOT PLUG IN THE X 10!!!)

4. Do the multiplication operation

3. Plug in 3.0 hit EE/Exp then the exponent 8

5. Then Equals VIOLA!!!

You Try IT!.

$$\begin{array}{r} 4.0 \times 10^{6} \\ \times\ 3.0 \times 10^{5} \\ \hline 1.2 \times 10^{12} \end{array}$$

A Problem for you…

$$
\begin{array}{r}
2.37 \times 10^{-6} \\
+\ 3.48 \times 10^{-4} \\
\hline
\end{array}
$$

All Measurement we make will be Using Metric System!

The Fundamental SI Units

(le Système International, SI)

Physical Quantity	**Name**	**Abbreviation**
Volume	liter	L
Mass	kilogram	kg
Length	meter	m
Time	second	s
Temperature	Kelvin	K
Electric Current	Ampere	A
Amount of Substance	mole	mol
Luminous Intensity	candela	cd

Common Metric Prefixes

Kilo	Hecto	Deca	Liter Meter Gram	deci	centi	milli
1,000	100	10	1	0.1	0.01	0.001
10^3	10^2	10^1		10^-	10^{-2}	10^{-3}
King	Hersey's	Daughter	Likes, Makes, Gulps	Delicious	Chocolate	Milk

We can easily change one metric unit into another by sliding the decimal! Be careful to watch S.F.

King	Hersey's	Daughter	Likes, Makes, Gulps	Delicious	Chocolate	Milk

- 1000 ml = 1 liter
- .005 kM = 5 meter
- 250,000 cm = 2.5 kilometer

You Try!

1. A snake slithers 0.000250 Km. How far is that in cm?

25.0 cm

2. A boy drinks 0.000500 kiloliters of water at practice. How many milliliters is that?

500. ml

9×10^{4} Kg

3. Doggie has a mass of 9×10^{7} grams. How many kilograms is that?

It’s not that easy to convert English measurements into Metric

- Need unit conversions:

- for length: 2.54 cm = 1 inch
- for mass: 454 g = 1 lb
- for volume 0.943 l = 1 quart

Metric Conversion Practice

In practice a conversion factor is used to convert between units.
Example We know that 1dollar = 4 quarters

How many quarters in 20 dollars?

$$\text{Number} \cdot (\cancel{\text{old unit}}) \times \left[\frac{\text{new unit}}{\cancel{\text{old unit}}}\right] = \text{New number} \cdot (\text{new unit})$$

Conversion factor

$$20 \text{ dollars} \quad \times \quad \frac{4 \text{ quarters}}{1 \text{ dollar}} \quad = \quad 80 \text{ quarters}$$

Type 1: Conversion of Distance
(always convert to metric)

- Example 1: Sammy the sail slithers 5.05 in how far is that in cm?

•5.05 inches = ? Cm

$$5.050\text{in} \left| \frac{2.54\text{ cm}}{1\text{ inch}} \right| = 12.8\text{ cm}$$

Example 2 : Bob the bunny hops 6.63 yards. How far is that in meters?

$$6.63\text{ yds} \quad \frac{3.00\text{ft}}{1\text{ yd}} \left| \frac{12.00\text{in}}{1\text{ ft}} \right| \frac{2.54\text{ cm}}{1\text{ in}} = 606\text{ cm} = 6.06\text{ meters}$$

Type 2: Volume Conversions

- 3. Mrs. Gleavy drank 1.55 gallons of water in a day. How many liters did she drink that day?

- $1.54 \text{ gal} \left| \frac{4 \text{ qts}}{1 \text{ gal}} \right| \frac{0.943 \text{ l}}{1 \text{ qt}} = 5.81 \text{ l}$

Type 3: Conversion of Mass

- 4. A child's chair can hold 150 kilograms. A person that weighs 195.0 pounds sits on the chair will it break?

- $$195.0\text{lbs} \times \frac{454\ \text{g}}{1\ \text{lb}} = 88500\ \text{g} = 88.5\ \text{kilograms}$$

NO! 88.5kg < 150 kg

Type Four: Two Units !!!!

- 5. A speed limit sign reads 40 km/hour. You are traveling 73.3 ft/min. Should you get a ticket?

$$\frac{73.3\ \text{ft}}{\text{Min}} \left| \frac{60.0\text{min}}{1\text{hr}} \right| \frac{12\ \text{in}}{1\ \text{ft}} \left| \frac{2.54\ \text{cm}}{1\ \text{in}} \right. = 134051.9\ \text{cm/hr}$$

$= 1.34\text{km/hr}$ NO < 40 km/Hr

You Try:

- 1. A polar bear with a weight of 275 pounds sat on a chair that can hold 98.0 kilograms. Will the chair break?

#2.

- A runner needs to complete a 5k road race. He is running 2.20 miles to see his predicted time. IS he running the correct distance?

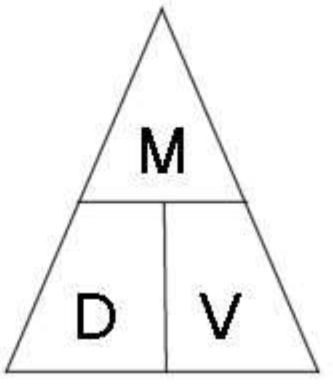
#3

- A speed limit sign reads 30km/hr. Your are traveling 25.0 ft/sec. Will you get a tiecket

Density- the amount of matter in a unit of volume-

can be used for identification purposes!

- Using the density triangle – any variable equation can be found by covering the unknown-



What can you conclude about the density of rubber, glycerol, oil, paraffin and cork?

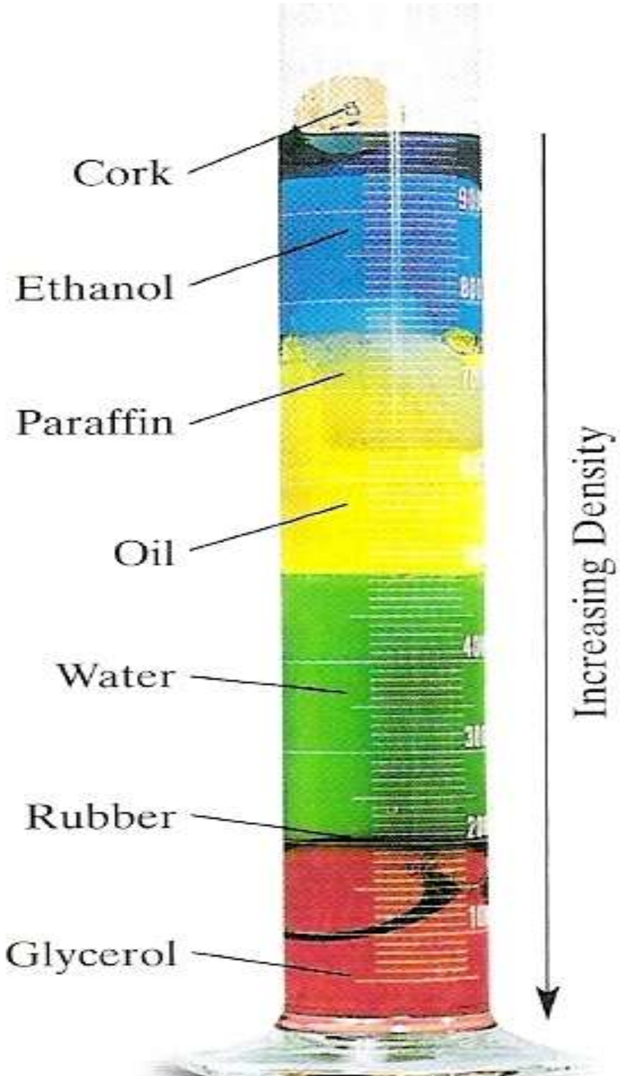


Table 4 Densities of Various Substances

Substance	Density (g/cm^3) at 25°C
Hydrogen gas, H_2*	0.000 082 4
Carbon dioxide gas, CO_2*	0.001 80
Ethanol (ethyl alcohol), C_2H_5OH	0.789
Water, H_2O	0.997
Sucrose (table sugar), $C_{12}H_{22}O_{11}$	1.587
Sodium chloride, NaCl	2.164
Aluminum, Al	2.699
Iron, Fe	7.86
Copper, Cu	8.94
Silver, Ag	10.5
Gold, Au	19.3
Osmium, Os	22.6

*at 1 atm

USEFUL INFORMATION!

1 cm^3 = 1 ml

Density of water = 1 g/ml therefore

50 g of water= 50 ml !

Finding Density

- *Grab your calculators folks!*

1. What is the density of a cube of material that has a mass of 25.00 grams and a side dimension of 2.0 cm?

2. A material has a mass of 45.8 grams and a volume of 7.15 ml. What is the density?

Volume can be determined two ways:

- Example One direct volume measurement.
- 3. Silver has a density of 10.5 g/cm^3. A cube with a side dimension of 2.0 cm is found. It has a mass of 84.0 grams. Could the cube be silver?

- Example two indirect volume measurement:
- 4. A necklace is found with a mass of 21.5 grams. When it is placed in 50.0 ml of water the water rises to 51.7 ml. Is the necklace silver?

Calculations using density and conversions with metric

- Example One direct volume measurement.

- 5. Copper has a density of 8.89 g/cm^3. What would the mass of a cube with a side dimension of 1.25 inches be?

- Example two indirect volume measurement:

- 6. Silver has a density of 10.5 g/ml. A necklace is found with a mass of 2.64×10^{-2} lbs. When it is placed in 50.0 ml of water the water rises to 51.7 ml. Is the necklace silver?

Finding Volume

- 7. Gold has a density of 19.34 g/cm^3. A nugget is found with a mass of 5.60 grams. What should 50.0 ml of water rise to if the nugget is gold?

Finding Mass

8. Copper has a density of 8.89 g/ml. A cube of copper with a side dimension of 3.0 cm is found. What will the mass be?

Finding the Identity of a Material

- 9. Charlie finds a cube of silvery colored metal. It has a mass of 57.12 grams and a side dimension of 2.0 cm. Can you identify the material?

Specific Heat

- Physical Property that is unique to the material

- Amount of energy required to heat 1 gram of a substance by 1 degree Celsius

Specific Heat

- The amount of heat energy required to raise the temperature of one grams of a substance by 1 ºc

Substance	C = J/gºc	Substance	C = J/gºc
Lead	0.129	Aluminum	0.897
Iron	0.449	Ethanol	2.44
Copper	0.385	Water (l)	4.184

Why do you suppose the bottom some aluminum pans are coated with copper?

Heat- sum of the kinetic energy of all particles in a system (Q)

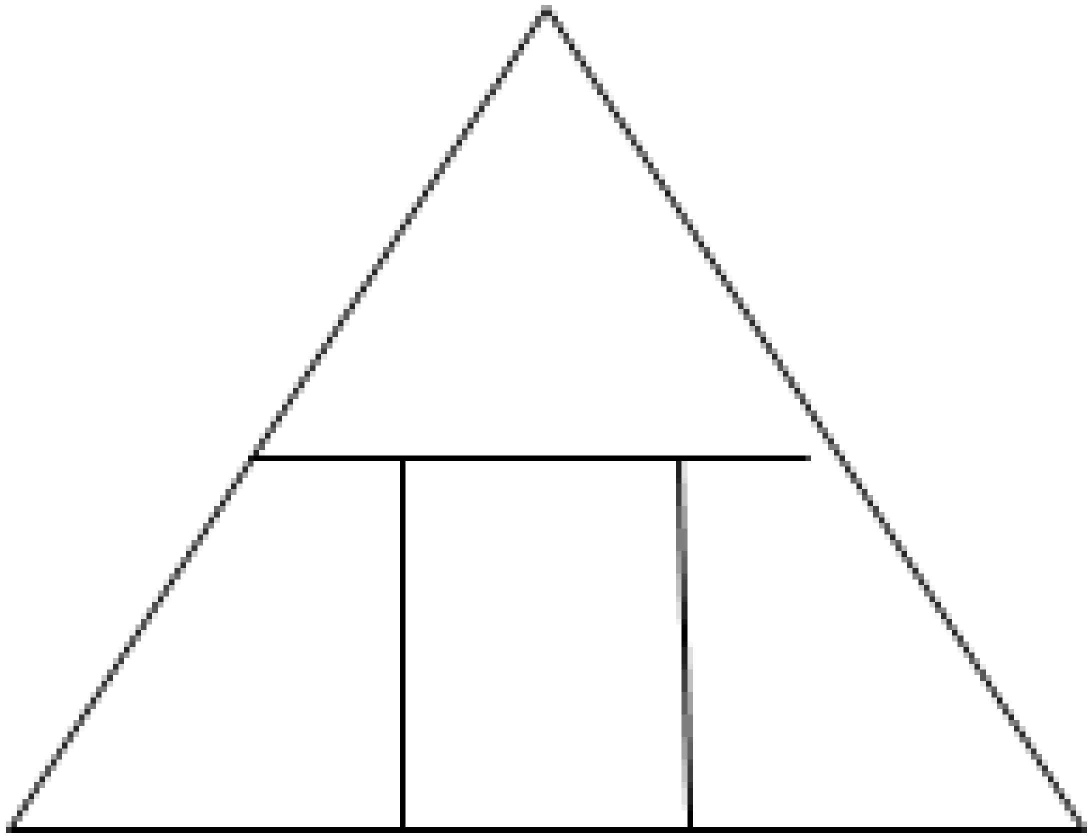
- Heat always flows from hot to cold!
- So why do we add ice cubes to a drink?

Heat Capacity

- Amount of energy required to change a given sample by a given amount
  - $Q = m\ C\ \Delta\ T$
- Q= Heat= Joules
- C= specific heat (table value) J/g^0c

  (unique to material)
- $\Delta\ T = T_{Final} - T_{Initial}$

Create a conversion triangle



Problems Finding Energy

- 1. a. How much energy is required to warm 5.00 grams of copper from 22.00c to 40.00c?

Copper 0.385J/g^{o}c

- b. How much energy is lost when 2.00 grams of lead is cooled from 25.00c to 15.00c?

Finding Mass

- 2. a. How many grams of water are in a sample if it required 166 joules of energy to be warmed from 20.00c to 40.00c?

Water (l) 4.184J/g^{o}c

- b. A sample of iron lost 66.6 joules of energy when cooled from 50.00c to 35.00c. What was the sample mass?

Finding Temperature

- 3. a.What is the final temperature if 25.0 grams of gold absorbs 32.25 joules of energy at 25.00c?

Gold 0.129J/g^{o}c

- 3b. What was the initial temperature of a 12.0 gram sample of iron if it absorbs 107. joules of energy ending at 31.00C?

Finding Identity

- 4. a. What is the identity of a material if 25.0 grams of the sample will absorb 59.3 joules of energy when warmed from 20.00c to 30.00c?

- b. A cube with a mass of 15.00 grams of “gold” colored material absorbs 38.7 joules of energy when warmed from 20.0oc to 40.00c. Is it gold?

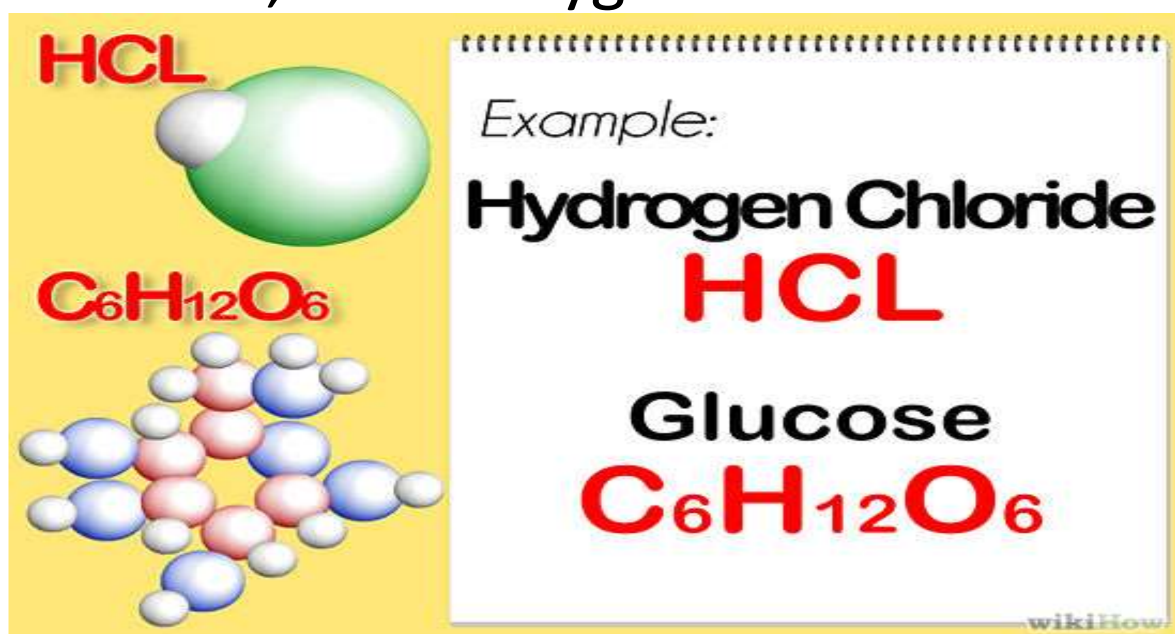
Last ONE!

- c. What material will gain 111 joules of energy if 25.0 grams are warmed from 20.00c to 30.00C?

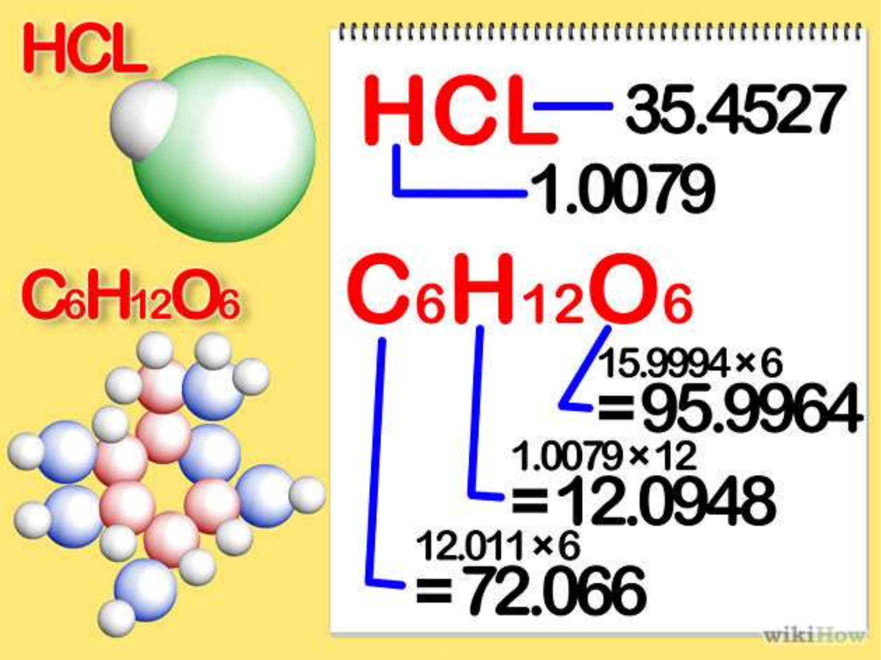
Calculating the Molar Mass of a Compound

1 Find the chemical formula for the compound. This is the number of atoms in each element that makes up the compound. For example, the formula for hydrogen chloride (hydrochloric acid) is HCl; for glucose, it is $C_6H_{12}O_6$. This means that glucose contains 6 carbon atoms, 12 hydrogen atoms, and 6 oxygen atoms.

•



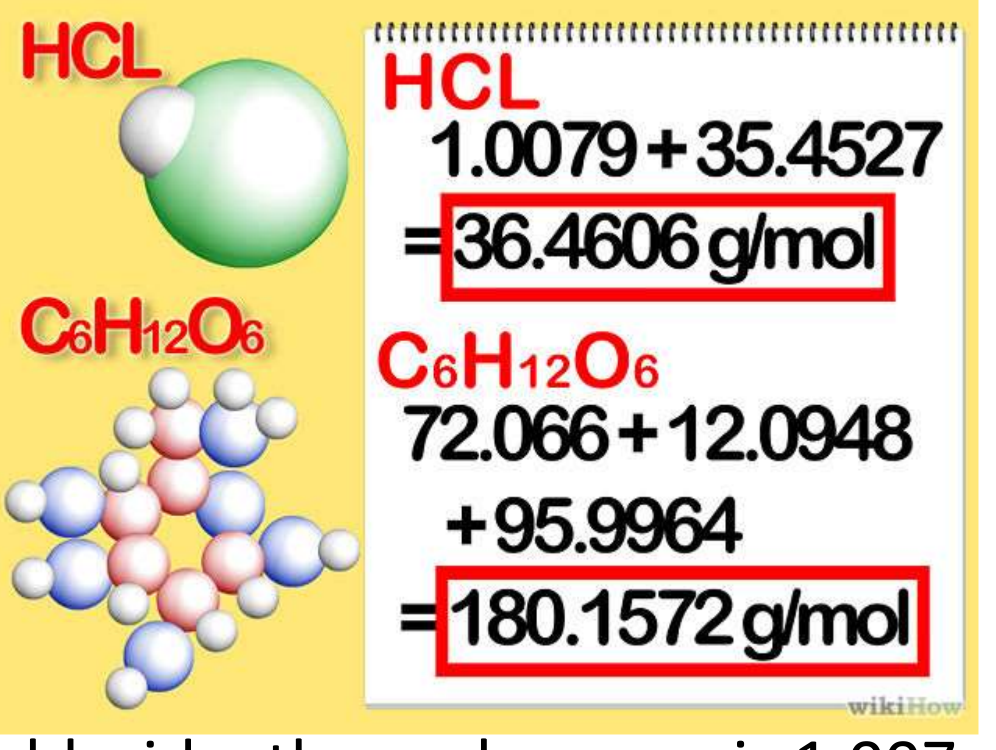
- Find the molar mass of each element in the compound. Multiply the element's atomic mass by the molar mass constant by the number of atoms of that element in the compound. Here's how you do it:



- For hydrogen chloride, HCl, the molar mass of each element is 1.007 grams per mole for hydrogen and 35.453 grams per mole for chlorine.
- For glucose, $C_6H_{12}O_6$, the molar mass of each element is 12.0107 times 6, or 72.0642 grams per mole for carbon; 1.007 times 12, or 12.084 grams per mole for hydrogen; and 15.9994 times 6, or 95.9964 grams per mole for oxygen.

Molar Mass

- Add the molar masses of each element in the compound. This determines the molar mass for the compound. Here's how you do it:



- For hydrogen chloride, the molar mass is 1.007 + 35.453, or 36.460 grams per mole.
- For glucose, the molar mass is 72.0642 + 12.084 + 95.9964, or 180.1446 grams per mole.

Chemical Quantities-The Mole

1 dozen = 12

1 gross = 144

1 ream = 500

1 mole = 6.02×10^{23}

There are <u>exactly</u> 12 grams of carbon-12 in one mole of carbon-12.

A mole is the atomic mass taken in grams of a substance

Avogradro's Number

6.02×10^{23} is called "Avogadro's Number" in honor of the Italian chemist Amadeo Avogadro (1776-1855).

I didn't discover it. Its just named after me!

Amadeo Avogadro

Moles

Remember that: Molar Mass has units of g/mol SO

Molar Mass = $\frac{\text{gram}}{\text{Mole}}$

Solve for Grams

Grams = **Moles =**

How many Moles in 5 grams of Carbon?

Avogadro's Number = 6.022×10^{23} atoms/mole SO

Avogadro's Number = $\frac{\text{Atoms}}{\text{Moles}}$

Solve for Atoms = **Moles =**

How many moles do I have if I have 1.8×10^{24} atoms

Diatomic Elements

In nature these elements exist in pairs.

Therefore the atomic mass is doubled

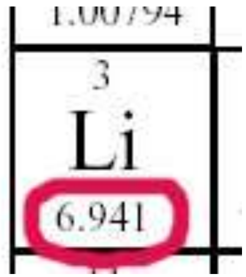
The SUPER SEVEN- There are seven of them, It starts with element 7- nitrogen- forms a seven and has a superhero hat of hydrogen!

Diatomic elements	
H_2	Hydrogen
N_2	Nitrogen
O_2	Oxygen
F_2	Fluorine
Cl_2	Chlorine
Br_2	Bromine
I_2	Iodine

Calculations with Moles:

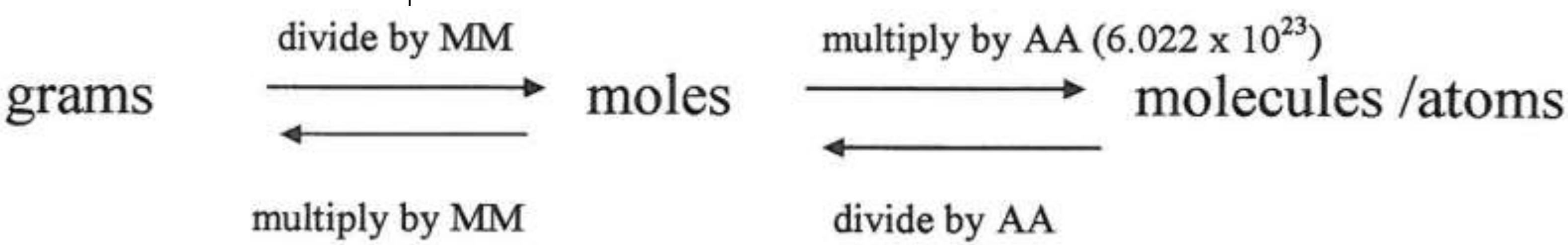
Converting moles to grams

How many grams of lithium are in 3.50 moles of lithium?



3.50 ~~mol~~ Li	6.94 g Li
	1 ~~mol~~ Li

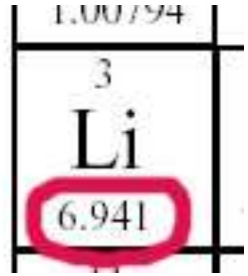
= 24.3 g Li



Calculations with Moles:

Converting grams to moles

How many moles of lithium are in 18.2 grams of lithium?



18.2 ~~g~~ Li	1 mol Li
	6.94 ~~g~~ Li

= 2.62 mol Li

Calculations with Moles:

Using Avogadro's Number

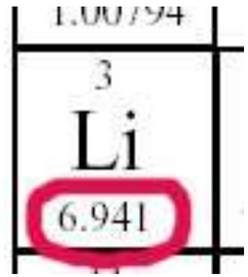
How many atoms of lithium are in 3.50 moles of lithium?

3.50 ~~mol~~ Li	6.022×10^{23} atoms Li	= 2.11×10^{24} atoms Li
	1 ~~mol~~ Li	

Calculations with Moles:

Using Avogadro's Number

How many <u>atoms</u> of lithium are in 18.2 *g* of lithium?



18.2 ~~g~~ Li	1 ~~mol~~ Li	6.022×10^{23} atoms Li
	6.94 ~~g~~ Li	1 ~~mol~~ Li

$(18.2)(6.022 \times 10^{23})/6.94 = 1.58 \times 10^{24}$ atoms Li

Finding Molar Mass of a Compound

- First decide how many of each type of atom you have. (*Remember to multiply a subscript outside a parenthesis to the atoms within)*
- Look up the individual masses on the P.T.
- Multiply the number of atoms by the mass
- Add all parts

What is the molar mass of copper II phosophate? ($Cu_3(PO_4)_2$)

Cu 3 x 63.55 +
P 2 X 30.97 +
O 8 x 16.00 =

380.59 g/mol

Converting to Moles with a Compound

Cindy masses 205.3 grams of $Cu_3(PO_4)_2$, how many moles does she have?

$$205.3 \text{ grams} \left| \frac{1 \text{ mole } Cu_3(PO_4)_2}{380.59 \text{ grams}} \right| = 0.539 \text{ moles}$$

Using A

- How many moles are used for an experiment if 2.57×10^{23} molecules of $Cu_3(PO_4)_2$ are consumed?
  - (again ignore MM and just divide by A#)

$$2.57 \times 10^{23} \text{ molecules of } Cu_3(PO_4)_2 \times \frac{1 \text{ mole}}{6.022 \times 10^{23}} = .427 \text{ moles } Cu_3(PO_4)_2$$

Converting to Grams

- Charlie needs to use 2.50×10^{-4} moles of $Cu_3(PO_4)_2$ for an experiment. How many milligrams should she mass out?

- 2.50×10^{-4} moles | 380.59 grams / 1 mole $Cu_3(PO_4)_2$ | = 9.51×10^{-2} g

$$2.50 \times 10^{-4} \text{ moles} \times \frac{380.59 \text{ grams}}{1 \text{ mole } Cu_3(PO_4)_2} = 9.51 \times 10^{-2} \text{ g}$$

Therefore: 95.1 milligrams

Using A#

- How many kilograms are consumed in a reaction if 2.45×10^{24} molecules of $Cu_3(PO_4)_2$ are used?

$$2.45 \times 10^{24} \text{ molecules of } Cu_3(PO_4)_2 \left| \frac{1 \text{ mole}}{6.022 \times 10^{23}} \right| \frac{380.59 \text{ grams}}{1 \text{ mole } Cu_3(PO_4)_2} = 1.55 \times 10^{3} \text{ g}$$

Therefore: 1.55 kilograms