A- Why Significant Figures? **The final answer should be rounded to reflect the accuracy of the measurements.**

Density example: $\frac{10.00 \pm 0.05 \ g}{3.00 \pm 0.05 \ ml}$ = 3.333….?

Answer could be as low as:	Answer could be as high as:	With the rules that we will learn, we will round it to 3.33, knowing that the last decimal is an estimate.
$\frac{9.95g}{3.05 \ ml} = 3.262g/ml$	$\frac{10.05g}{2.95 \ ml} = 3.407g/ml$	

Reading the Meniscus at Eye Level
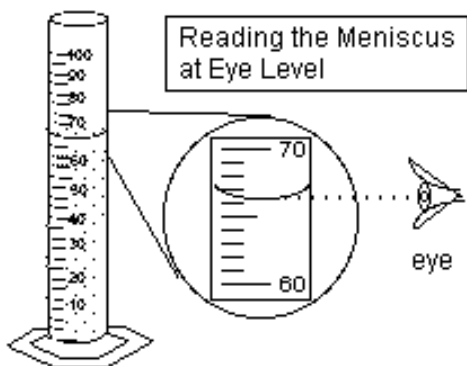
B-   **Rules for Significant Figures**

1. Measurements should always include one estimated figure. That figure is considered significant.
      2.   Non-zero digits and captive zeros are always significant.

Example 1   Report the measurement with the correct number of sig figs. Also report the measurement if the bottom of the meniscus was exactly on the 60 ml mark.

Example 2   How many significant figures in the following?

a) 30.004
b) 1.25 X10^3

3. Leading zeros are never significant.
4. Trailing zeros are only significant in the presence of a decimal.

Example3   The police estimated a crowd of 300 000 fans at the Bruins Stanley Cup parade. What is the # of sig figs?

Example 4   High resolution photos from rooftops estimated the crowd to be 277 000. What's the # of sig figs?

Example5   0.0005 ml

5. **Exact numbers have an unlimited number of sig figs.**

Example 6:   What exact numbers are used in chemistry?

6. **When multiplication and division are involved in a series of calculations, the final answer must have as many sig figs as the measurement with the least number of sig figs. (22.4 L/mole, molar masses and 8.31 kPaL/kmole are all measurements.)**
7. **When using molar masses, use *at least as many sig figs* as there are in the other measurements in the problem.**
8. **Only apply rule number 6 in the last step. Keep all decimal places on your calculator in between calculations.**

Example 7:   Convert 90.0 g of H_2 to moles.

9. **If a calculation only involves adding or subtracting, the answer must have as many decimal places as the measurement with the least decimal places.**

Example 8:   0.00003g + 10.15 g = ????

What is the logic behind this rule?

Exercises

How many significant figures?

1. 0.0004 ml
2. 3.0005 g
3. 900 kg
4. 900. Kg

5. **a) Read the thermometer and report the measurement with the correct number of sig figs.**

   b) What if the line was right on the zero?

   ### Apply the rules of sig figs for the following problems:

   1. If 0.010 g of mass are destroyed in a fission reaction, how much energy will be released? $c = 3.00 \times 10^8$ m / s
   2. The molar mass of Cl is 35.45 g/mole. What is the mass of 3.001 moles of Cl_2?
   3. Convert 1.0×10^{-4} mol/L of NaOH to ppm.