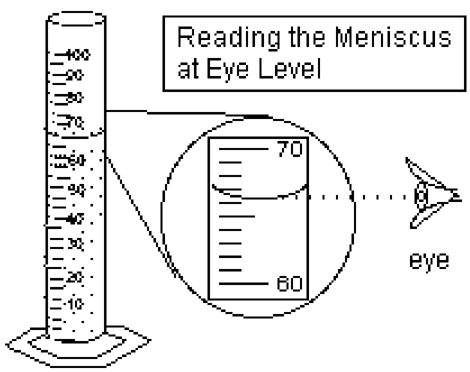
A- Why Significant Figures? **The final answer should be rounded to reflect the accuracy of the measurements.**

Density example: $\frac{10.00 \pm 0.01\, g}{3.00 \pm 0.05\, ml}$ = 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. |
|---|---|---|
| $\frac{9.99g}{3.05\ ml} = 3.28\ g/ml$ | $\frac{10.01g}{2.95\ ml} = 3.39\ g/ml$ | |




**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) $1.25 \text{ X}10^3$

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.

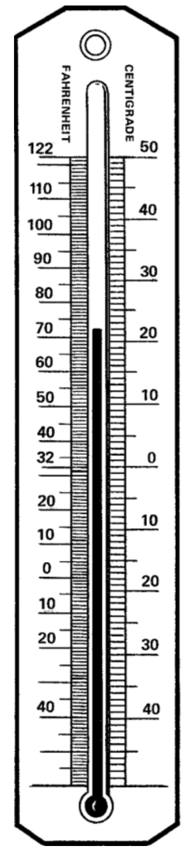
9. **If a calculation only involves adding or subtracting, the answer must have as many decimal places as the measurement with the least decimal places.**

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. 2.00 g
6. 0.02000 g
7. 1.0030 g
8. $2.90 \times 10^{-3}$ g

9. 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:

10. 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 and $E = mc^2$
11. The molar mass of Cl is 35.45 g/mole. What is the mass of 3.001 moles of $Cl_2$?
12. Convert $1.0 \times 10^{-4}$ mol/L of NaOH to ppm.

13. Find the sum of the molar masses of H ( 1.00797) and Cl (35.45) in g/mole