CALCULATIONS USING SIGNIFICANT FIGURES

Name Key

When multiplying and dividing, limit and round to the least number of significant figures in any of the factors.

Example 1: 23.0 cm x 432 cm x 19 cm = 188,784 cm^3
The answer is expressed as 190,000 cm^3 since 19 cm has only two significant figures.

When adding and subtracting, limit and round your answer to the least number of decimal places in any of the numbers that make up your answer.

Example 2: 123.25 mL + 46.0 mL + 86.257 mL = 255.507 mL
The answer is expressed as 255.5 mL since 46.0 mL has only one decimal place.

Perform the following operations expressing the answer in the correct number of significant figures. Write units.

1. 1.35 m x 2.467 m = 3.33 m^2
2. 1,035 m^2 ÷ 42 m = 25 m
3. 12.01 mL + 35.2 mL + 6 mL = 53 mL
4. 55.46 g − 28.9 g = 26.6 g
5. .021 cm x 3.2 cm x 100.1 cm = 6.7 cm^3
6. 0.15 cm + 1.15 cm + 2.051 cm = 3.35 cm
7. 150 L^3 ÷ 4 L = 37.5 L^2 = 40 L^2
8. 505 kg − 450.25 kg = 55 kg
9. 1.252 mm x 0.115 mm x 0.012 mm = 1.7×10^{-3} mm^3 or 0.0017 mm^3
10. 1.278×10^3 m^2 ÷ 1.4267×10^2 m = 8.958 m

SIGNIFICANT FIGURES

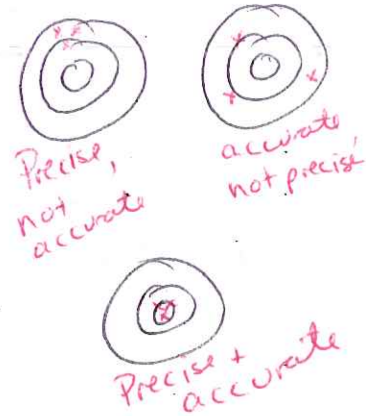
Name Key

A measurement can only be as accurate and precise as the instrument that produced it. A scientist must be able to express the accuracy of a number, not just its numerical value. We can determine the accuracy of a number by the number of significant figures it contains.

precision - the closeness of a set of measurements to each other

accuracy - the closeness of measurements to the correct or accepted value of the quantity measured

1) All digits 1-9 inclusive are significant.
   Example: 129 has 3 significant figures.
2) Zeros between significant digits are always significant.
   Example: 5,007 has 4 significant figures.
3) Trailing zeros in a number are significant <u>only</u> if the number contains a decimal point.
   Example: 100.0 has 4 significant figures.
   100 has 1 significant figure.
4) Zeros in the beginning of a number whose only function is to place the decimal point are not significant.
   Example: 0.0025 has 2 significant figures.
5) Zeros following a decimal significant figure are significant.
   Example: 0.000<u>470</u> has 3 significant figures.
   0.<u>47000</u> has 5 significant figures.



Determine the number of significant figures in the following numbers.

1. 0.02 _1_
2. 0.020 _2_
3. 501 _3_
4. 501.0 _4_
5. 5,000 _1_
6. 5,000. _4_
7. 6,051.00 _6_
8. 0.0005 _1_
9. 0.1020 _4_
10. 10,001 _5_

Determine the location of the last significant place value by placing a bar over the digit. (Example: 1.70$\bar{0}$)

1. 80$\bar{4}$0 _3_
2. 0.030$\bar{0}$ _3_
3. 699.$\bar{5}$ _4_
4. 2.00$\bar{0}$ x 10^2 _4_
5. .9010$\bar{0}$ _5_
6. 90,$\bar{1}$00 _3_
7. 4.$\bar{7}$ x 10^{-8} _2_
8. 10,800,00$\bar{0}$. _8_
9. 3.0$\bar{1}$ x 10^{21} _3_
10. 0.00041$\bar{0}$ _3_

Dimensional Analysis-Grams, Moles, and Molecules   Name: Key

Background:

- *A mole is defined as 6.02 x 10^{23} particles (aka Avogadro's number). This is similar to a dozen being defined as 12 particles. In chemistry these "particles" are most often atoms, molecules, or ions.*

  1 mole = 6.02 x 10^{23} particles

- *Molar mass is the mass (usually in grams) of 1 mole of particles. The molar mass of a substance is determined by its chemical formula. We can find the molar mass of a substance by using the atomic masses on the periodic table.*

  Examples include: the molar mass of Na is 23.0 g/mol; the molar mass of Br_2 is (2 x 79.9 =) 159.8 g/mol; the molar mass of H_2O is (2 x 1.0 + 16.0=) 18.0 g/mol.

  1 mole = (Molar Mass from the Periodic Table) grams

Remember that in the process of dimensional analysis, we use equivalent statements to write ratios (fractions) that can be used to convert from one unit to another.

Example 1: What is the mass of 0.875 moles of CO_2?

$$0.875\ moles\ CO_2 \times \frac{44.0\ grams}{1\ mole} = 38.5\ grams\ CO_2$$

Example 2: How many atoms of Al are in a 255 gram sample of Al?

$$255\ grams\ Al \times \frac{1\ mole\ Al}{27.0\ grams\ Al} \times \frac{6.02\ x\ 10^{23}\ atoms\ Al}{1\ mole\ Al} = 5.69\ x\ 10^{24}\ atoms\ Al$$

Try These: Show your work for each problem using dimensional analysis. Put units on every number that you write down (in your work AND on your answer).

1. What is the mass of 5.07 x 10^{22} ~~molecules~~ atoms of C?

$$5.07 \times 10^{22}\ atoms \times \frac{1\ mol}{6.02 \times 10^{23}\ atoms} \times \frac{12.0\ g}{1\ mol} = 1.01\ g\ C$$

2. How many ~~molecules~~ atoms of S would be in 4.75 moles of S?

$$4.75\ mol\ S \times \frac{6.02 \times 10^{23}\ atoms}{1\ mol\ S} = 2.86 \times 10^{24}\ atoms$$

3. How many atoms of gold are in a 1.00 kg gold bar?

$$1.00\ kg\ Au \times \frac{1000\ g}{1\ kg} \times \frac{1\ mol}{197.0\ g} \times \frac{6.02 \times 10^{23}\ atom}{1\ mol} = 3.06 \times 10^{24}\ atoms$$

4. How many moles of silver are in a sample containing 8.0 x 10^{25} atoms of silver?

$$8.0 \times 10^{25}\ atoms\ Ag \times \frac{1\ mol}{6.02 \times 10^{23}\ atoms} = 133\ mol = 130\ mol$$ ← 2 sig figs

5. A typical banana contains 422 mg of potassium ~~ions~~ atoms. How many ~~ions~~ atoms is this?

$$422\ mg\ K \times \frac{0.001\ g}{1\ mg} \times \frac{1\ mol}{39.1\ g} \times \frac{6.02 \times 10^{23}\ atoms}{1\ mol} = 6.50 \times 10^{21}\ atoms\ K$$