Measurements & Calculations

Metric System

Conversions within the metric system that are most common are largely created by moving 1000 times (exception is cm which is 100x)

Terra – giga – mega – Kilo – (meter, liter, gram etc.) – milli – micro – nano – pico
Largest ----------smallest

Units:

Length – meter (m)	Time – seconds (s)	Pressure – pascal (pa)
Mass – gram (g)	Energy – joules (j)	Temperature - kelvin
Volume – liter (l)	quantity – mole (mol)	

Derived Units:

Area = m^2	Molar Mass = g/mol	Density = g/cm^3
Volume = m^3	Concentration = mol/l	Velocity – m/s

Significant Figures

I. Significant figures
- A. General Rule – match your answer to the lowest number of significant figures provided in a question.
- B. All digits other than 0 are always significant
- C. 0 between digits other than 0 are significant
  1. 203 has 3 significant digits
- D. 0 to the right of the decimal are significant if they are after a number other than 0
  1. indicates the degree of accuracy in a measurement
     a) 2.300 has 4 significant figures
- E. 0 to the right of a whole number are not significant
  1. 2300 has 2 significant figures
  2. a bar over a whole number indicates significant figures
     a) 2300 has 3 significant figures
- F. 0 to the left of a number are not significant
  1. 0.0023 has 2 significant figures
- G. Think counter intuitively
  1. If you need the zeroes to make the number, they are not significant. (2400, .00045)
  2. If the zeroes are not needed to make the number, they are significant. (2.400)

II. Adding & Subtracting Significant Figures
- A. Round the sum or difference to the same number of decimal places as the quantity with the least number of decimal places.
  1. 5.34 + 10.4 + 1.82 = 10.36 = 10.4

III. Multiplying & Dividing Significant Figures

A. The answer will contain the same number of significant figures as the least significant number in the problem, rounded off.
1. 2 X 2.6 = 5.2 = 5
2. 2.2 X 2.0 = 4.8

Sig Figs Tutorial

Common Calculations in Chemistry

I. Density = mass/volume
A. Water has a density of 1g/cm^3 (mL)

II. Percent Error = Observed value - True value
(i) True value

III. Thermometry - temperature is a measure of the average kinetic energy of the particles in a system
A. Heat flows spontaneously from a system at higher temperature to a system of lower temperature
B. Thermometer - an instrument uses to measure the heat energy in a system
1. Celsius (C) - has fixed points of 0°C for ice to water 100°C for steam to water
2. Kelvin (K) - absolute temperature scale
a) 0°K = absolute zero - the temperature at which all motion ceases
b) 273°K = ice to water
c) 373°K = steam to water
(1) Celsius to Kelvin Conversions = °C + 273°

Exponents

I. Exponential Notation
A. Move the decimal place until there is only 1 number to the left of the decimal point
B. Count the number of places the decimal is moved
1. Use that number as your exponent
a. moved left - + exponent
b. moved right - - exponent

II. Adding and Subtracting Exponents
A. Make sure that the exponents being added have the same power value
1. 2.1 X 10^3 + 1.8 X 10^3 = 3.9 X 10^3
2. If not, convert one of the exponents until they both have the same power
a. it is usually easier to reduce the larger exponent or move both toward the power of 1
b. 2.1 X 10^3 + 1.8 X 10^2 =
c. 21.0 X 10^2 + 1.8 X 10^2 = 22.8 X 10^2 = 2.28 X 10^3
d. Add the number as normal, keep the exponent

III. Multiplying Exponents
A. Multiply the numbers present
B. add the exponents
1. 2.1 X 10^3 * 2.0 X 10^4 = 4.2 X 10^7

IV. Dividing Exponents
A. Divide the numbers present

B. Subtract the exponents

1. $4.2 \times 10^6 / 1.4 \times 10^4 = 3.0 \times 10^2$

Explanation of Exponents

Practice Writing Scientific Notation