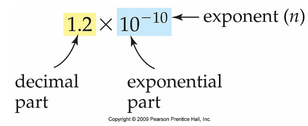
Chapter 2 – Measurement and Problem Solving

Working with numbers

1. Uncertainty
   a. There is a certain amount of doubt in every measurement.
      i. It is important to know the uncertainty when measurements are recorded or read.
      ii. **Accuracy** – the closeness between a measurement and its true value. Since the true value is rarely known, accuracy is based on how close independent studies agree. The closer they agree, the more confidence scientists have in the accuracy of their results.
         1. Affected by ***determinate errors*** – errors due to poor technique or incorrectly calibrated instruments.
      iii. **Precision** – describes the reproducibility of the experiment. How close are the repeated measurements to each other?
         1. Affected by ***indeterminate errors*** – errors due to estimating the last uncertain digit of a measurement.
            a. Random
            b. Cannot be eliminated
   b. Communicating with numbers
      i. How do we communicate with others the uncertainty in our measurements?
         1. The last digit in a measurement involving numbers is uncertain.
         2. When reading measurements, an estimate is made between the smallest divisions on the scale of the ruler, thermometer, or dial.
            a. The last digit on an electronic device will be considered the uncertain digit.
            b. The difference between "2" and "2.0". These are not the same number. By saying "2", we say the number is not 1 and not 3. By saying "2.0", we know the number is not 1.9 or 2.1.
            c. It is the responsibility of the experimenter to write numbers in a way that reflect the precision of the measurement.
      ii. Values reported as results from experiments are in the form of the "mean plus-or-minus one standard deviation" ($X \pm$ s.d.).
         1. This gives a range over which we have confidence that the true value will fall.
            a. Used in scientific journals
            b. Requires a sample of at least three
            c. Most calculators are able to compute the mean and standard deviation.
         2. The standard deviation indicates how much each value differs from the average.
         3. The size of deviation would indicate precision in running a sequence of experiments. The smaller the deviation, the more precise the work was. (Candle Examples)
      iii. **Percent Error (Percent Difference)**
         1. **absolute error** = difference between your experimental value (observed value) and what you are supposed to get (actual value).
         2. **percent error** = **percent difference** = a common way of reporting results.

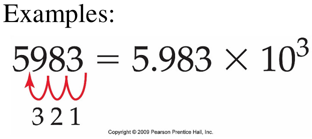
$$\% \text{ error} = \% \text{ diff} = \left| \frac{actual - observed}{actual} \right| x \ 100$$

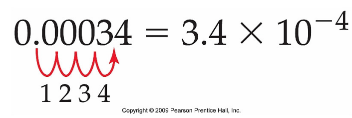
2. Scientific Notation:



a. Used to express large and small numbers by getting rid of the place holding zeros.
b. If the decimal is moved to the <u>left the exponent is positive</u>, and if the decimal is moved to the <u>right the exponent is negative</u>.

Examples:





2. Significant figures

a. Digits in a measurement having value plus one digit having a value that is estimated.
   i. All non-zero digits are significant
   ii. A zero is special. It has four different functions (figure 1)
      1. As a number
      2. **As an indicator of accuracy**
      3. As a place holder
      4. As a cosmetic zero
b. Zeros are significant when they are
   i. Interior zeros (1.**05**, 0.01**01**)
   ii. Trailing zeros after a decimal point and a non-zero digit (0.04**0**, 0.0004**00**,
   iii. Leading zeros (to the left of the first non-zero digit) are not significant. The number 0.0005 has only one significant digit.
   iv. The difference between zeros used for accuracy and those being used as placeholders can only be addressed using scientific notation.
      1. 1, 000, 000 → indicates all zeros are place holders
      2. 1.00×10^{6} → indicates the first two zeros are significant
      3. 1.00000×10^{6} → indicates all the zeros are significant
      4. Examples: (significant digits in bold)
         a. **23.46** mL
         b. 0.00**36** seconds
         c. **854.236** g
         d. **6.02** x 10^{23} molecules
         e. 0.**98** mol
         f. 00**23** m
         g. **2.000** J
         h. **1.00026** x 10^{-3} cm
         i. **824** mg
c. **Exact values**. Some numbers are exact values which involve no uncertainty. There is an exact number of people in the classroom. If there are 20 people, the zero is automatically significant.
   i. Chemistry examples:
      1. There are 4.184 joules in a calorie
      2. There are exactly two hydrogen atoms in a water molecule.
      3. Stoichiometric coefficients and subscripts
      4. Defined quantities. If we know 1 liter = 1000 mL, the number 1000 actually has an infinite number of significant digits.

4. Mathematical operations with significant digits.

a. Rounding to the correct number of significant figures

1. Round down if the last (or leftmost) digit dropped is 4 or less.
2. Round up if the last digit dropped is 5 or more.
3. Be certain to use only the last digit being dropped to decide in which direction to round – ignore all digits to the right of it.

b. Addition and subtraction. The number of places to the right of the decimal point (degree of accuracy), determines the number of significant digits reported.

c. Multiplication and division. In these operations, the number of reported digits is determined by the least accurate value (one with the fewest significant digits).

d. For calculations involving multiple steps, **round only the final answer.**

e. In calculations involving both multiplication/division and addition/subtraction, do the steps in parentheses first.

5. **Derived units are formed from other units**

a. Volume

i. Units of length to the 3rd power become volume.

ii. cm^3 is derived from multiplying the lengths of three dimensions.

iii. Common volume units that we will use are based on liters.

1. A milliliter (mL) is equivalent to 1 cm^3.
2. A liter, since it contains 1000 mL, would be the same as a cube of 1000 cm^3.

b. Density

i. The density of a substance is a ratio of its mass to volume.

$$\text{Density} = \frac{\text{mass}}{\text{volume}} \quad \text{or} \quad d = \frac{m}{V}$$

TABLE 2.4 Densities of Some Common Substances

Substance	Density (g/cm^3)
water	1.0
ice	0.92
ethanol	0.789
lead	11.4
copper	8.96
gold	19.3
aluminum	2.7
platinum	21.4
iron	7.86
titanium	4.50
charcoal, oak	0.57
glass	2.6

1. What happens to water when it freezes?

2. Explain why ice floats?

3. A crown is tested to determine its density. It displaces 10.7 mL of water and has a mass of 206 g. What is it made of?

6. **Converting from one unit to another**
   a. Dimensional analysis (Factor Label) is the process of converting from one unit to another.
   b. Key Ideas:
      i. Always write down your units and numbers.
      ii. Use the "ladder" or "bracket" method to keep things organized.
   c. Example: How old is a 16.5 year old high school student, in seconds? (linear)
   d. **Class activity**: If this room needed to be retiled, how many square feet of tiles should be purchased? (square)
   e. If a vinyl tiles are sold for $15.50/yd^2, how much will the tiles cost for this room? (square)
   f. **Class activity**: How many liters of air can fit in this room? (cubic)
   g. If an indoor running track is 305 meters long, how many laps must be completed in order to run 25 km? (metric)
   h. In 1999, a new class of black holes was discovered, with masses 100 to 10,000 times the mass of our sun, but occupying less space than our moon. One of these holes has a mass of 1×10^3 suns and a radius equal to one-half the radius of our moon. What is the density in g/cm^3? (Sun mass = 2.0×10^{30}kg, moon radius = 2.16×10^3 mi, sphere volume = $4/3\pi r^3$)