Chapter 1 Chemistry and Measurement (Essential Ideas in OpenStax)

Matter
- Matter is anything that has substance and occupies space. Matter also has mass and volume.
- A material is any particular type of matter.
- An atom is the smallest base unit of matter which has a chemical identity.
  It exists as a very tiny particle.
  The basic theory of atoms was developed by John Dalton in early 19th century (chapter 2).
- A molecule is a set of atoms that are bonded together in
  a particular arrangement to form a more complex material.
- A substance is a material with a single chemical identity, and a pure substance
  is composed entirely of one particular atom, molecule, or chemical formula.

Scientific Method
- An experiment is an observation of phenomena that is controlled
  so that rational conclusions can be obtained from results.
- A law is a fundamental relationship or regularity of nature.
  It is stated concisely, such as with an equation.
- A hypothesis is a tentative explanation of natural regularity.
  It can be tested with further experimentation.
- A theory is an explanation for which extensive testing has proven its validity.
  It cannot be proven absolutely, but a well-developed theory would explain
  everything observed, and would not be contradicted by any observations.
- Theories can have limitations, and can be improved upon or replaced.
  This process requires more experiments and more explanations.

Conservation of Mass
- Mass is a definite quantity of matter.
- The Law of Conservation of Matter is that total mass remains constant during a chemical rxn.
- Matter cannot be created or destroyed except by nuclear reaction, where $E = mc^2$.
  This law was first demonstrated by Antoine Lavoisier in late 18th century.
- Weight (w) is not the same as mass (m). Weight is the force of gravity acting upon a mass.
  The general equation for force is $F = ma$, where a is acceleration in m/s^2.
  For weight, the equation is $w = mg$, where $g = 9.81$ m/s^2 (acceleration due to gravity).

Ex 1.01 Matter (or Mass) is Conserved in a Chemical Reaction
- aluminum + oxygen → aluminum oxide (a white odorless crystalline powder)
- Mass before reaction (of aluminum and oxygen) = Mass after reaction (of aluminum oxide)
- (5.40 g aluminum) + (mass oxygen) = 10.20 g
- mass oxygen = 10.20 g – 5.40 g = 4.80 g

Phases

- Solid is matter that is rigid, has a fixed volume and a fixed shape (not a fluid).
- Liquid has a relatively fixed volume, but no fixed shape (fluid).
- Gas also has no fixed shape, but is easily compressible.
  Its volume is a function of pressure and temperature.
- Gases and liquids are fluids. That is, they have no fixed shape.

Physical and Chemical Changes

- A physical change is when matter changes its form, but not its identity,
- Physical changes include vaporization, distillation (separation into components by vaporizing with heat), and dissolving a molecular substance.
- A physical property is a characteristic that is observed without a change to chemical identity.
- A chemical change is when matter changes identity, and it involves a chemical reaction.
  A chemical change occurs when molecules form or decompose.
- Chemical changes include dissolving an ionic substance, which separates the substance into its component ions.
- A chemical property describes a change to chemical identity.
- A substance cannot be converted into components or into another substance by a purely physical change.
- An element is the most basic type of substance.
  It cannot decompose into a simpler substance.
  All of the atoms in an element have the same chemical identity.
- A compound is a substance composed of two or more elements.

Law of Definite Proportions

- A pure compound has constant integer proportions of its elements.
  This law was developed by Joseph Louis Proust (early 19^{th} century).
- For example, ammonia always contains 3 moles of H for every 1 mole of N.
  Each molecule contains three H atoms and one N atom.
  So, its formula is a constant, and is written as NH_3.

Mixtures

- A mixture can generally be separated into two or more substances by a pure physical process.
- A homogeneous mixture is completely uniform down to the atomic or molecular level.
  This is called a solution, and all of the molecules are dissolved together completely.
- A heterogeneous mixture contains two or more physically distinct parts. If the mixture contains particles of one material dispersed within another continuous material, it is called a colloid or a dispersion. Examples of colloids include foams, gels, and aerosols.
- A phase is a single homogeneous material (a solid, liquid, or gas).
  It can be either a pure substance or a homogeneous mixture.

Physical Measurements

- Measurements determine physical quantities and are always expressed with fixed standard units of measurement.
- Precision is the closeness of a set of values to each other for identical measurements (repeatability).
- Accuracy is the closeness of the measurements to the actual value (limits of error).

What Significant Figures Are

- Significant figures report known digits only and tell us how well a particular value is known.
- They include all certain digits plus a final digit which has some uncertainty.

Which Zeros Are Significant

1. 0's at the very beginning of the value (left side) are place holders only and are not significant.
2. Terminal 0's to right of the decimal point are significant. Report these if and only if certain: 0.5 has only one significant digit, however 0.500 has three and the terminal zeros are known.
3. Terminal 0's to left of the decimal point may not be significant: 53000 is ambiguous. Use scientific notation to remove the ambiguity: 5.30×10^4 has three digits.

How To Multiply and Divide with Significant Figures

- Use as many significant digits in the final result as in the measurement which has the least number of significant digits. For instance, $0.3\underline{1} \times 18.02 = 5.\underline{6}$ with two digits.

How To Add and Subtract with Significant Figures

- Use as many decimal places in the final result as in the measurement which has the least number of decimal places. For instance, $16.0\underline{0} + 2.016 = 18.0\underline{2}$ with two decimal places.

How To Treat Exact Numbers

- Exact numbers have no uncertainty. They include integers and some conversion factors.
- They do not decrease the number of significant digits or decimal points in a calculation at all.
- Treat them as if they have an infinite number of significant 0's following the value.

How To Round

- Drop the non-significant digits. Then, adjust the last significant digit accordingly.
- If first non-significant digit is < 5, round DOWN by retaining the last significant digit.
- If first non-significant digit is > 5, round UP by adding 1 to the last significant digit.
- If first non-significant digit is = 5, round UP unless there are no digits at all past the 5. If the non-significant 5 is the very last digit, then round to an EVEN digit.

How To Apply Significant Digits In Your Calculations

- Keep the non-significant digits during the intermediate calculation steps.
- Drop the non-significant digits and round last digit to submit the final reported result.
- Report ALL of the significant digits and ONLY the significant digits in the final answer!

Example 1.02 Perform Calculations with Significant Figures

a) For $4.578 \times 6.\underline{8} / 5.8257$, the result has two digits.
   This is because the measurement with the least number of digits (two) is $6.\underline{8}$.

b) For $7.4\underline{4} - 0.299$, the result ($7.1\underline{4}$) has two decimal places, as does $7.4\underline{4}$.

c) For $9.2\underline{8} - 8.3\underline{1}$, the result ($0.9\underline{7}$) has two decimal places, as do both measurements.

d) $86.51 \times (9.2\underline{8} - 8.3\underline{1}) = 86.51 \times 0.9\underline{7} = 8\underline{4}$
   The result ($8\underline{3}.9147$) has two significant digits, which leaves zero decimal places.

e) $72.88 - (8\underline{3}.9147)$ has zero decimal places ($-1\underline{1}$).

International System of Units (SI) – the metric system

- Base units are m, kg, s, Kelvins (K), mole, ampere (A), and candela (cd).
- Prefixes are used to denote exponents of 10.
  milli (m) = 1/1000   centi (c) = 1/100   deci (d) = 1/10   kilo (k) = 1000
  1 mm = 1/1000 m = 0.001 m and 1000 mm = 1 m
  1 km = 1000 m and 1 m = 1/1000 km = 0.001 km
- Angstrom (Å): 1 Å $= 10^{-10}$ m $= 10^{-8}$ cm = 100 pm   (1 pm $= 10^{-12}$ m)

Temperature Conversions

- Scientific temperature measurements are in units of Kelvins, which equal ^{o}C + 273.15.
  25.00 ^{o}C = 298.15 K and 0.00 K = − 273.15 ^{o}C
- Conversions between Celsius and Fahrenheit are based on 0 ^{o}C = 32 ^{o}F (freezing pt of water) and a temperature change of 5 ^{o}C equals a change of 9 ^{o}F. All of those values are exact.
- This gives us two equations: $^{o}\text{C} = \left(\frac{5}{9}\right) \times (^{o}\text{F} - 32)$ and $^{o}\text{F} = (1.8) \times (^{o}\text{C}) + 32$
  For example: $(1.8)(25\,^{o}\text{C}) + 32 = 45 + 32 = 77\,^{o}\text{F}$

Ex 1.03 Convert $83.\underline{0}\,^{o}$F to ^{o}C and K

- $^{o}\text{C} = (5/9) \times (83.\underline{0} - 32.0) = (5/9) \times (51.\underline{0}) = 28.\underline{3}333\,^{o}\text{C} = 28.\underline{3}\,^{o}\text{C}$
- $\text{K} = 28.\underline{3}333 + 273.15 = 301.\underline{4}833\ \text{K} = 301.\underline{5}\ \text{K}$

Derived Units

- Derived units are obtained by multiplying and dividing base units together.
  See the list of derived units in Table 2.
- length^2 = area, so the derived unit for area is m^2
- length^3 = volume, so the derived unit for volume is m^3
- density = mass divided by volume, so the derived units for density are kg/m^3 and g/cm^3.

Dimensional Analysis: Using Conversion Factors

- $1\text{ L} = 1000\text{ cm}^3$ is a unit equation, which gives two conversion factors:

$$\left(\frac{1\text{L}}{1000\text{ cm}^3}\right) \quad \text{and} \quad \left(\frac{1000\text{ cm}^3}{1\text{L}}\right)$$

- Since the two values are always equivalent, a conversion factor is always equal to 1.
- The volume in cm^3 for 2 L of water = $(2\text{ L})\left(\frac{1000\text{ cm}^3}{1\text{L}}\right) = 2000\text{ cm}^3$
- Exponents can be used in the values and the units as well.
  10^2 cm = 1 m can be cubed for volume, and $(10^2\text{ cm})^3 = (1\text{ m})^3$ simplifies to $10^6\text{ cm}^3 = 1\text{ m}^3$

Conversions with Derived Units

- $1\text{ L} = 1\text{ dm}^3 = (1\text{ dm}^3)\left(\frac{10^{-1}\text{m}}{1\text{ dm}}\right)^3 = (1\text{ dm}^3)\left(\frac{10^{-3}\text{ m}^3}{1\text{ dm}^3}\right) = 10^{-3}\text{ m}^3$
- $1\text{ L} = 1\text{ dm}^3 = (1\text{ dm}^3)\left(\frac{10\text{ cm}}{1\text{ dm}}\right)^3 = (1\text{ dm}^3)\left(\frac{10^3\text{cm}^3}{1\text{ dm}^3}\right) = 10^3\text{ cm}^3$
- $1\text{ L} = 1000\text{ ml}$ and $1\text{ ml} = \frac{1}{1000}\text{ L}$
- $1\text{ ml} = \left(\frac{1}{1000}\text{ L}\right)\left(\frac{10^{-3}\text{m}^3}{1\text{ L}}\right) = (10^{-3}\text{ L})\left(\frac{10^{-3}\text{m}^3}{1\text{ L}}\right) = 10^{(-3)+(-3)}\text{ m}^3 = 10^{-6}\text{ m}^3$
- $1\text{ ml} = \left(\frac{1}{1000}\text{ L}\right)\left(\frac{10^{+3}\text{cm}^3}{1\text{ L}}\right) = (10^{-3}\text{ L})\left(\frac{10^{+3}\text{cm}^3}{1\text{ L}}\right) = 10^{(-3)+(+3)}\text{ cm}^3 = 10^0\text{ cm}^3 = 1\text{ cm}^3$

Density

- Density is mass per unit volume, and the equation is $d = m/V$.
  The units are g/ml for liquids, which are equivalent to g/cm^3 for solids.
- The equation can be rearranged to $V = m/d$ and $m = d \times V$.
- The maximum density of water is 0.99997 g/ml at 4 $^{\text{o}}$C, just slightly less than exactly 1 g/ml.

Ex 1.04 Determine Density from Volume and Mass

- 8.10 ml of a clear liquid sample has a mass of 6.367 g.
- $d = m/V = \frac{(6.367\text{ g})}{(8.10\text{ ml})} = 0.786\text{ g/ml}$
- The result does not match the density of water.
  But, it does match the density of isopropyl alcohol.

Ex 1.05 Use Density to Determine Volume from Mass

- An isopropyl alcohol sample has a mass of 37.4 g.
- $V = m/d = \frac{(37.4\text{ g})}{\left(0.786\,\frac{\text{g}}{\text{ml}}\right)} = 47.6\text{ ml}$
- Density is being used here as a conversion factor.

Ex 1.06 Metric Conversion Factors

- Convert 25.4 g to mg and kg in scientific notation.
- $(25.\underline{4}\text{ g})\left(\frac{10^3\text{mg}}{1\text{ g}}\right) = (2.5\underline{4} \times 10^1\text{ g})\left(\frac{10^3\text{mg}}{1\text{ g}}\right) = 2.5\underline{4} \times 10^4\text{ mg}$
- $(25.\underline{4}\text{ g})\left(\frac{1\text{ kg}}{10^3\text{g}}\right) = (2.5\underline{4} \times 10^1\text{ g})\left(\frac{10^{-3}\text{ kg}}{1\text{ g}}\right) = 2.5\underline{4} \times 10^{-2}\text{ kg}$

Ex 1.07 Conversion Factors derived by using Exponents

- The Earth possesses a total of $1.386 \times 10^9\text{ km}^3$ of water.
  Convert that value to L and kg using the density of water as 1.000 g/mL.
- $\left(\frac{10^3\text{ m}}{1\text{ km}}\right)^3 = \frac{10^{3\times3}\text{ m}^3}{1^3\text{ km}^3} = \frac{10^9\text{ m}^3}{1\text{ km}^3}$ $\left(\frac{10\text{ dm}}{1\text{ m}}\right)^3 = \frac{10^3\text{ dm}^3}{1\text{ m}^3}$
- $(1.38\underline{6} \times 10^9\text{ km}^3)\left(\frac{10^9\text{ m}^3}{1\text{ km}^3}\right) = 1.38\underline{6} \times 10^{18}\text{ m}^3$
- $(1.38\underline{6} \times 10^{18}\text{ m}^3)\left(\frac{10^3\text{ dm}^3}{1\text{ m}^3}\right)\left(\frac{1\text{ L}}{1\text{ dm}^3}\right) = 1.38\underline{6} \times 10^{21}\text{ L}$
- $(1.38\underline{6} \times 10^{21}\text{ L})\left(\frac{10^3\text{ ml}}{1\text{L}}\right)\left(\frac{1.000\text{ g}}{1\text{ mL}}\right)\left(\frac{1\text{kg}}{10^3\text{ g}}\right) = 1.38\underline{6} \times 10^{21}\text{ kg}$

Ex 1.08 Conversions between English (lbs and oz) and Metric (g)

- The first conversion factor is an exact number within the English system.
  So there is no change to number of significant digits.
- $(5.127\underline{5}\text{ lb})\left(\frac{16\text{ oz}}{1\text{ lb}}\right) = 82.04\underline{0}\text{ oz}$
- The second conversion factor converts between systems and is not an exact number.
  So it can and does reduce the number of significant digits.
- $(82.04\underline{0}\text{ oz})\left(\frac{28.35\text{ g}}{1\text{ oz}}\right) = 2.32\underline{6} \times 10^3\text{ g}$