1.1 The Study of Chemistry

- Chemistry – study of properties of materials and changes that they undergo.
  - Can be applied to all aspects of life (e.g. development of pharmaceuticals, leaf color change in fall).

The Molecular Perspective of Chemistry

Chemistry involves the study of the properties and behavior of matter.

- **Matter**
  - Physical material of the universe
  - Has mass
  - Occupies space
  - ~100 **Elements** constitute all matter
  - **Elements:**
    - Made up of unique **atoms.**
    - Names of the elements are derived from a wide variety of sources (e.g. Latin or Greek, mythological characters, names of people or places).
  - **Molecules**:
    - Combinations of atoms held together in specific *shapes*.
    - Macroscopic (observable) properties of matter relate to microscopic realms of atoms.
    - Properties relate to composition (types of atoms present) and structure (arrangement of atoms) present.

Why Study Chemistry?

We study chemistry because:

- It has a considerable impact on society (health care, food, clothing, conservation of natural resources, environmental issues, etc.).
- It is part of your curriculum! Chemistry serves biology, engineering, agriculture, geology, physics, etc. **Chemistry is the *central science.***

1.2 Classifications of Matter

Matter is classified by **state** (solid, liquid, or gas) or by **composition** (element, compound, or mixture).

States of Matter

On the macroscopic level:

- **Gas** No fixed volume or shape, conforms to volume and shape of container, compressible.
- **Liquid** Volume independent of container, no fixed shape, incompressible.
- **Solid** Volume and shape independent of container, rigid, incompressible.

On the molecular level

- **Gas** Molecules far apart, move at high speeds, collide often.
- **Liquid** Molecules closer than those in gas, move rapidly but can slide over one another.
- **Solid** Molecules packed closely in definite arrangements.

Pure Substances and Mixtures

- **Pure Substance**
  - Matter with fixed composition and distinct proportions.
  - **Elements** (cannot be decomposed into simpler substances, i.e. only one kind of atom) or **compounds** (consist of two or more elements).
- **Mixtures**
  - Combination of two or more pure substances.
  - Variable composition.
  - Heterogeneous (do not have uniform composition, properties and appearance, e.g. sand).
  - Homogeneous (uniform throughout, e.g. air). Homogeneous mixtures are **solutions**.

Separation of Mixtures

Key: Separation techniques exploit differences in properties of the *components*.

- Filtration: Remove solid from liquid
- Distillation: Boil off one or more components of the mixture
- Chromatography: Exploit solubility of components

Elements

- 112 known
- Vary in abundance
- Each is given a unique name
- Organized in periodic table
- Each given a one- or two-letter symbol derived from its name.

Compounds

- Combination of elements
  Example: The compound H_2O is a combination of the elements H and O.
- The opposite of compound formation is decomposition
- Compounds have different properties than their component elements (e.g. water is liquid, hydrogen and oxygen are both gases at the same temperature and pressure).
- **Law of constant (definite) proportions** (Proust): A compound always consists of the same combination of elements (e.g. water is always 11 percent H and 89 percent O).

1.3 Properties of Matter

Physical vs. Chemical Properties

- **Physical properties**: Measured without changing the substance (e.g. color, density, odor, melting point).
- **Chemical properties**: Describe how substances react or change to form different substances (e.g. hydrogen burns in oxygen).
- **Intensive properties**: Do not depend on the amount of substance present (e.g. temperature, melting point); give an idea of the composition of a substance.
- **Extensive properties**: Depend on quantity of substance present (e.g. mass, volume).

Physical and Chemical Change

- **Physical change**: Substance changes physical appearance without altering its identity (e.g. **changes of state**).
- **Chemical changes** (or **chemical reactions**): Substances transform into chemically different substances (i.e. identity changes, e.g. decomposition of water, explosion of nitrogen triiodide).

The Scientific Method

The scientific method: guidelines for the practice of science.

- Collect data (observe, experiment, etc.)
- Look for patterns, try to explain them and develop a **hypothesis**.
- Test hypothesis; refine it.
- Bring all information together into a **scientific law** (concise statement or equation that summarizes tested hypotheses).
- Bring hypotheses and laws together into a theory. a **theory** should explain general principles.

1.4 Units of Measurement

- Many properties of matter are quantitative.
- A measured quantity must have BOTH a number and a unit.
- The units most often used for scientific measurement are those of the **metric system**.

SI Units

- 1960: all scientific units use **S**ysteme **I**nternational d'Unites (**SI** Units)
- There are seven base units.
- Smaller and larger units are decimal fractions or multiples of the base units.

Length and Mass

- SI base unit of length = meter (1 m = 1.0936 yards).
- SI base unit of mass (not weight) = kilogram (1 kg = 2.2 pounds).
  - **Mass** is a measure of the amount of material in an object.

Temperature

- Scientific studies use Celsius and Kelvin scales.
- **Celsius scale**: Water freezes at 0 ^{o}C and boils at 100 ^{o}C (sea level).
- **Kelvin scale** (SI unit)
  - Water freezes at 273.15 K and boils at 373.15 K (sea level).
  - Based on properties of gases.
  - Zero is lowest possible temperature (absolute zero).
  - 0 K = -273.15 ^{o}C.

- Fahrenheit (not used in science)
  - Water freezes at 32 ^{o}F and boils at 212 ^{o}F (sea level)
  - Conversions:

$$^{o}F - 32 = 1.8\ ^{o}C \qquad\qquad K = {^{o}C} + 273.15$$

Derived SI Units

- These are formed from the seven base units.
- Example: velocity is distance traveled per unit time, so units of velocity are units of distance (m) divided by units of time (s): m / s

Volume

- Units of volume = (units of length)^3 e.g. m^3
- This unit is unrealistically large, so we use more reasonable units:
  - cm^3 (also known as mL or cc (cubic centimeters))
  - dm^3 (also known as liters, L)
  - Important: the liter is *not* an SI unit.

Density

- Used to characterize substances.
- **Density** is defined as mass divided by volume.
- Units are usually g / cm ^{3}.
- Originally based on mass (the density was defined as the mass of 1.00 g of pure water).

1.5 Uncertainty in Measurement

- Two types of numbers:
  - Exact numbers (known by counting or definition)
  - Inexact numbers (derived from measurement)
- All measurements have some degree of uncertainty or *error* associated with them.

Precision and Accuracy

- **Precision**: how well measured quantities agree with one another.
- **Accuracy**: how well measured quantities agree with the "true value."
- Figure 1.25 is very helpful in making this distinction.

Significant Figures

- In a measurement it is useful to indicate the exactness of the measurement. This exactness is reflected in the number of significant figures.
- Guidelines for determining the number of significant figures in a measured quantity:
  - The number of significant figures is the number of digits known with certainty plus one uncertain digit. (Example 2.2405 g means we are sure that the mass is 2.240 g, but we are uncertain about the nearest 0.0001 g.)
- Final calculations are only as significant as the least significant measurement.
- Rules:
  1. Nonzero numbers are always significant.
  2. Zeros between nonzero numbers are always significant.
  3. Zeros before the first nonzero digit are **not significant**. (Example: 0.0003 has one significant figure.)
  4. Zeros at the end of the number after a decimal place are significant.
  5. Zeros at the end of the number after a decimal place are ambiguous (e.g. 10,300 g).

- Method:
  1) Write the number in scientific notation.
  2) The number of digits remaining is the number of significant figures.
  3) Examples:
     - 2.50×10^4 cm has 3 significant figures as written
     - 1.03×10^4 g has 3 significant figures
     - 1.030×10^4 g has 4 significant figures
     - 1.0300×10^4 has 5 significant figures

Significant Figures in Calculations

- Multiplication and Division
  - Report to the least number of significant figures
    - (e.g. 6.221 cm x 5.2 cm = 32 cm^2)

- Addition and Subtraction
  - Report to the least number of decimal places
    - (e.g. 20.4 g - 1.322 g = 19.1 g).

- In multiple-step calculations always retain an extra significant figure until the end to prevent rounding errors.

1.6 Dimensional Analysis

- Method of calculation utilizing a knowledge of units.
- Given units can be multiplied and divided to give the desired units.
- Conversion factors are used to manipulate units:
  - Desired unit = given unit x (conversion factor)
- The **conversion factors** are simple ratios:
  - **Conversion factor =** (desired unit) / (given unit)

Using Two or More Conversion Factors

- We often need to use more than one conversion factor in order to complete a problem.
- When identical units are found in the numerator and denominator of a conversion. they will cancel. The final answer **MUST** have the correct units.
  - Example:
    - Suppose that we want to convert length in meters to length in inches. We can do this conversion with the following conversion factors:

      1 meter = 100 centimeters and 1 inch = 2.54 centimeters

    - The calculation will involve both conversion factors; the units of the final answer will be inches:

$$\#\ inches = (\#\ meters)\left(\frac{100\ centimeters}{1\ meter}\right)\left(\frac{1\ inch}{2.54\ centimeters}\right) = \#\ inches$$

Conversions Involving Volume

- We often will encounter conversions from one measure to a different measure.
  - Example:
    1. Suppose that we wish to know the mass in grams of 2.00 cubic inches of gold given that the density of the gold is 19.3 g/cm .

    2. We can do this conversion with the following conversion factors:

       2.54 cm = 1 inch and 1 cm^3 = 19.3 g gold

    3. The calculation will involve both of these factors:

$$\text{x g gold} = (2.00\ in^3)\left(\frac{2.54\ cm}{1\ inch}\right)^3\left(\frac{19.3\ g\ Au}{1\ cm^3}\right) = 633\ g\ Au$$

    4. Note that the calculation will NOT be correct unless the centimeter to inch conversion is cubed! Both the units AND the number must be cubed.

Summary of Dimensional Analysis

- In dimensional analysis always ask three questions:
  1. What data are we given?
  2. What quantity do we need?
  3. What conversion factors are available to take us from what we are given to what we need?