Chapter 1. Introduction: Matter and Measurement

Lecture Outline

1.1 The Study of Chemistry

- Chemistry:
  - is the 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, etc.).

The Atomic and Molecular Perspective of Chemistry

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

- **Matter**:
  - is the physical material of the universe.
  - has mass.
  - occupies space.
  - ~100 **elements** constitute all matter.
  - A **property** is any characteristic that allows us to recognize a particular type of matter and to distinguish it from other types of matter.
- **Elements**:
  - are made up of unique **atoms**, the building blocks of matter.
  - 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**:
    - are combinations of atoms held together in specific shapes.
    - Macroscopic (observable) properties of matter relate to submicroscopic 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

- Solids, liquids and gases are the three forms of matter called the **states of matter**.
- Properties described on the macroscopic level:
  - **gas** (vapor): no fixed volume or shape, conforms to shape of container, compressible.
  - **liquid**: volume independent of container, no fixed shape, incompressible.
  - **solid**: volume and shape independent of container, rigid, incompressible.
- Properties described on the molecular level:
  - **gas**: molecules far apart, move at high speeds, collide often.
  - **liquid**: molecules closer than gas, move rapidly but can slide over each other.
  - **solid**: molecules packed closely in definite arrangements.

Pure Substances

- **Pure substances**:
  - are matter with fixed compositions and distinct proportions.
  - are **elements** (cannot be decomposed into simpler substances, i.e. only one kind of atom) or **compounds** (consist of two or more elements).
- **Mixtures**:
  - are a combination of two or more pure substances.
  - Each substance retains its own identity.

Elements

- There are 116 known elements.
- They vary in abundance.
- Each is given a unique name and is abbreviated by a chemical *symbol*.
- they are organized in the periodic table.
- Each is given a one- or two-letter symbol derived from its name.

Compounds

- **Compounds** are combinations 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, but 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% H and 89% O).

Mixtures

- A **mixture** is a combination of two or more pure substances.
  - Each substance retains its own identity; each substance is a *component* of the mixture.
  - Mixtures have variable composition.
  - Heterogeneous mixtures do not have uniform composition, properties, and appearance, e.g., sand.
  - Homogeneous mixtures are uniform throughout, e.g., air; they are **solutions**.

1.3 Properties of Matter

- Each substance has a unique set of physical and chemical properties.
  - **Physical properties** are measured without changing the substance (e.g., color, density, odor, melting point, etc.).
  - **Chemical properties** describe how substances react or change to form different substances (e.g., hydrogen burns in oxygen).
  - Properties may be categorized as intensive or extensive.
  - **Intensive properties** do not depend on the amount of substance present (e.g., temperature, melting point etc.).
  - **Extensive properties** depend on the quantity of substance present (e.g., mass, volume etc.).
    - Intensive properties give an idea of the composition of a substance whereas extensive properties give an indication of the quantity of substance present.

Physical and Chemical Changes

- **Physical change**: substance changes physical appearance without altering its identity (e.g., **changes of state**).
- **Chemical change** (or **chemical reaction**): substance transforms into a chemically different substance (i.e. identity changes, e.g., decomposition of water, explosion of nitrogen triiodide).

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.

The Scientific Method

- **The scientific method** provides guidelines for the practice of science.
    - Collect data (observe, experiment, etc.).
    - Look for patterns, try to explain them, and develop a **hypothesis** *or tentative explanation*.
    - Test hypothesis, then 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, i.e., associated with numbers.
- 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**ystème **I**nternational d'Unités (**SI** Units).
- There are seven base units.
- Smaller and larger units are obtained by 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

- *Temperature* is the measure of the hotness or coldness of an object.
- Scientific studies use Celsius and Kelvin scales.
- **Celsius scale**: water freezes at 0°C and boils at 100°C (sea level).
- **Kelvin scale** (SI Unit):
    - Water freezes at 273.15 K and boils at 373.15 K (sea level).
    - is based on properties of gases.
    - Zero is the lowest possible temperature (absolute zero).
    - 0 K = –273.15°C.
- Fahrenheit (not used in science):
    - Water freezes at 32°F and boils at 212°F (sea level).
    - Conversions:

$$°\mathrm{F} = \frac{9}{5}\,°\mathrm{C} + 32$$

$$°\mathrm{C} = \frac{5}{9}(°\mathrm{F} - 32)$$

$$°\mathrm{C} = \mathrm{K} - 273.15$$

$$\mathrm{K} = °\mathrm{C} + 273.15$$

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 = m^3.
- This unit is unrealistically large, so we use more reasonable units:
  - cm^3 [also known as mL (milliliter) or cc (cubic centimeters)]
  - dm^3 (also known as liters, L).
- Important: the liter is not an SI unit.

Density

- Is used to characterize substances.
- **Density** is defined as mass divided by volume.
- Units: g/cm^3 or g/mL (for solids and liquids); g/L (often used for gases).
- Was originally based on mass (the density was defined as the mass of 1.00 g of pure water at 25□C).

1.5 Uncertainty in Measurement

- There are two types of numbers:
  - *exact numbers* (known as counting or defined).
  - *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 each other.
- **Accuracy**: how well measured quantities agree with the "true value."
- Figure 1.24 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 are:
  - 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 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 and zeros between nonzero numbers are always significant.
  2. Zeros before the first nonzero digit are not significant. (Example: 0.0003 has one significant figure.)
  3. Zeros at the end of the number after a decimal point are significant.
  4. Zeros at the end of a number before a decimal point are ambiguous (e.g., 10,300 g). Exponential notation eliminates this ambiguity.
- Method:
  1. Write the number in scientific notation.
  2. The number of digits remaining is the number of significant figures.
  3. Examples:
     2.50×10^2 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 g 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

- **Dimensional analysis** is a 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)
        - These are fractions whose numerator and denominator are the same quantity expressed in different units.
        - Multiplication by a conversion factor is equivalent to multiplying by a factor of one.

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.
- For example:
    - Suppose that we want to convert length in meters to length in inches. We could do this conversion with the following conversion factors:
        - 1 meter = 100 centimeters and 1 inch = 2.54 centimeters
- The calculation would involve both conversion factors; the units of the final answer will be inches:
        - (# meters) (100 centimeters / 1 meter) (1 inch / 2.54 centimeters) = # inches

Conversions Involving Volume

- We will often encounter conversions from one measure to a different measure.
- For example:
    - 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^3.
    - We could do this conversion with the following conversion factors:

$$2.54 \text{ cm} = 1 \text{ inch and } 1 \text{ cm}^3 = 19.3 \text{ g gold}$$

    - The calculation would involve both of these factors:

$$(2.00 \text{ in.}^3)\,(2.54 \text{ cm / in.})^3\,(19.3 \text{ g gold} / 1 \text{ cm}^3) = 633 \text{ g gold}$$

    - Note that the calculation will NOT be correct unless the centimeter to inch conversion factor 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?