Fundamental Concepts

C020

Matter and Measurement *(M&H Ch. 1)*

Chemistry and Matter

Matter is anything that has a mass and occupies space, and chemistry is the study of matter at the **molecular level**, where chemists examine matter's qualitative and quantitative aspects

There are three phases of matter:

- **Solids:** have fixed volume and shape
- **Liquids:** have a fixed volume but take the shape of the container
- **Gases:** take the full volume and shape of the container

Matter can be classified as follows:

- **matter**
  - **mixtures**
    - **heterogeneous**
    - **homogeneous**
  - **pure substances**
    - **elements**
    - **compounds**

Elements

The simplest pure substance is an element, which cannot be further broken down into two or more pure substances

These are identified by symbols on the periodic table. Some are derived from English names, while others are not.

For example: English: C = carbon Al = aluminum

Latin: Fe = ferrum (iron) Hg = hydragyrum (mercury)

Oxygen gas (O_2) is a molecule that contains two oxygen atoms.
Since it can not be broken down into two or more pure *stable* substances, it is classified as an element.

Note that the symbol of oxygen on the periodic table is still "O"

Compounds

Compounds are pure substances that contain two or more elements.
For example, water (H_2O) is a compound

In any given compound, the proportions of each element (either by the # of atoms or by mass) are always the same

Methane (natural gas) is CH_4

Propane (BBQ tank fuel) is C_3H_8

*The properties of compounds are different from those of the elements that make the compound

Mixtures, homogeneous or heterogeneous, contain two or more substances mixed (but not reacted) together.
Each substance retains its own chemical identity, but the properties of the mixture may differ from those of the substances

Homogeneous mixtures

In a homogeneous mixture, the composition is the same throughout the mixture.

If the mixture is a liquid, it can be called a **solution**, which consists of one or more **solutes** (minimum concentration) dissolved in a **solvent** (maximum concentration)

- Salt (NaCl) dissolves in water to form a salt solution that has a melting point lower than water alone
- The two substances in solution can be isolated by evaporation, distillation and other techniques

Heterogeneous mixtures

In heterogeneous mixtures, the composition varies through the mixture. These are often found as solid-solid or solid-liquid mixtures

For example:

- Rocks such as granite
- Broken glass in water which can be separated by filtration

We can also describe the physical state of matter by appending the following notations after its chemical formula:

➢(s) = solid	e.g. $H_2O(s)$	ice
➢(ℓ) = liquid	e.g. $H_2O(\ell)$	water
➢(g) = gas	e.g. $H_2O(g)$	steam
➢(aq) = aqueous	e.g. NaCl(aq)	salt in water
➢(g,at) = gaseous atomic	e.g. O(g,at)	oxygen atom
	(Compare with $O_2(g)$	oxygen gas)

Qualitative Aspects

Qualitative aspects are descriptive.
For example:

- How do we make substance X in the lab?
- What happens when X reacts with Y?
- How can we convert X into Z?
- How can we separate X and Y?

Quantitative Aspects

Quantitative aspects involve numbers and measurements.
For example:

- What is the length of the bond in the molecule X-X?
- How much X reacts with one gram of Y?
- How fast does X react with Y?
- How much heat energy is evolved in the reaction?

In most sciences including chemistry, SI units are used.
(SI ≡ le système international d'unités = metric)

Examples of some SI units and their derivatives:

Mass

- kilogram kg
- gram g 10^{-3} kg
- milligram mg 10^{-6} kg = 10^{-3} g
- microgram µg 10^{-9} kg = 10^{-6} g

Length

- meter m
- kilometer km 1000 m
- centimeter cm 10^{-2} m
- millimeter mm 10^{-3} m
- micrometer µm 10^{-6} m
- nanometer nm 10^{-9} m
- picometer pm 10^{-12} m

Time

- second s
- use the same prefixes as above

Volume

- Although the official SI unit for volume is the cubic meter (m^3), the liter (L) is more commonly used

- One L = 10^{-3} m^3

- One mL = 10^{-3} L = 10^{-6} m^3 = $(1 \times 10^{-2}\ m)^3$ = 1 cm^3 (a common medical abbreviation for cm^3 is cc)

An important quantity is the mole (mol) is the exact number of units as there are carbon atoms in 12 grams of the isotope ^{12}C
= Avogadro's number = 6.022×10^{23} = N_A (more later)

Other units to be encountered include:

- Temperature (Celsius or Kelvin)
- Pressure (force per area: N/m^2 or Pascal)

Some important units are products or ratios of simple fundamental units:

- Density = mass/volume = kg/m^3 or g/cm^3
- Velocity (speed) = distance/time = m/s or cm/s
- Concentration is often expressed by comparing the mass or moles of solutes in a given amount of solvent

$$\frac{\text{mass solute}}{\text{volume solvent}} = \frac{\text{g}}{\text{L}}$$

$$\frac{\text{moles solute}}{\text{volume solvent}} = \frac{\text{mol}}{\text{L}} = \text{MOLARITY (M)}$$

Mole fraction:
if your basket contains 2 mol brown eggs and 3 mol white eggs,
then the respective mole fractions are 2/5 and 3/5 = 0.4 and 0.6 (no units)
(more later)