Ch 100: Fundamentals for Chemistry

Chapter 2: Measurements & Calculations
Lecture Notes

Types of Observations

- Qualitative
  - Descriptive/subjective in nature
  - Detail qualities such as color, taste, etc.

  Example: "It is really warm outside today"
- Quantitative
  - Described by a number and a unit (an accepted reference scale)
  - Also known as <u>measurements</u>
- Notes on Measurements:
    - Described with a value (number) & a unit (reference scale)
    - Both the value and unit are of equal importance!!
    - The value indicates a measurement's size (based on its unit)
    - The unit indicates a measurement's relationship to other physical quantities

  Example: "The temperature is 85^{o}F outside today"

Application of Scientific Notation

Writing numbers in Scientific Notation

1 Locate the Decimal Point
2 Move the decimal point to the **right** of the non-zero digit in the largest place
   – The new number is now between 1 and 10
3 Multiply the new number by 10^{n}
   – where ***n*** is the number of places you moved the decimal point
4 Determine the sign on the exponent, *n*
   – If the decimal point was moved left, *n* is **+**
   – If the decimal point was moved right, *n* is **–**
   – If the decimal point was not moved, *n* is **0**

Writing Scientific Notation numbers in Conventional form

1 Determine the sign of *n* of 10^{n}
   – If *n* is + the decimal point will move to the right
   – If *n* is – the decimal point will move to the left
2 Determine the value of the exponent of 10
   – Tells the number of places to move the decimal point
3 Move the decimal point and rewrite the number

Measurement Systems

There are 3 standard unit systems we will focus on:

1. United States Customary System (USCS)
   - formerly the British system of measurement
   - Used in US, Albania, and a couple other countries
   - Base units are defined but seem arbitrary (e.g. there are 12 inches in 1 foot)
2. Metric
   - Used by most countries
   - Developed in France during Napoleon's reign
   - Units are related by powers of 10 (e.g. there are 1000 meters in 1 kilometer)
3. SI (L'Systeme Internationale)
   - a sub-set set of metric units
   - Used by scientists and most science textbooks
   - Not always the most practical unit system for lab work

Measurements & the Metric System

- All units in the metric system are related to the fundamental unit by a power of 10
- The power of 10 is indicated by a prefix
- The prefixes are always the same, regardless of the fundamental unit
- When a measurement has a specific metric unit (i.e. 25 cm) it can be expressed using different metric units <u>without</u> changing its meaning

Example: 25 cm is the same as 0.25 m or even 250 mm

- The choice of measurement unit is somewhat arbitrary, what is important is the observation it represents

Measurement, Uncertainty & Significant Figures

- A measurement always has some amount of uncertainty
- Uncertainty comes from limitations of the techniques used for comparison
- To understand how reliable a measurement is, we need to understand the limitations of the measurement
- To indicate the uncertainty of a single measurement scientists use a system called **significant figures**
- The last digit written in a measurement is the number that is considered to be uncertain
- Unless stated otherwise, the uncertainty in the last digit is ±1

<u>**Examples:**</u>

1. The measurement: 25.2 cm uncertainty: 0.1 cm
2. The measurement: 25.20 cm uncertainty: 0.01 cm
3. The measurement: 25.200 cm uncertainty: 0.001 cm

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Rules for Counting Significant Figures

- Nonzero integers are always significant
- Zeros
  - Leading zeros never count as significant figures
  - Captive zeros are always significant
  - Trailing zeros are significant if the number has a decimal point
- Exact numbers have an unlimited number of significant figures

Rules for Rounding Off

- If the digit to be removed is
  1. less than 5, the preceding digit stays the same
  2. equal to or greater than 5, the preceding digit is increased by 1
- In a series of calculations, carry the extra digits to the final result and *then* round off
- **Don't forget to add place-holding zeros if necessary to keep value the same!!**

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Exact Numbers

Exact Numbers are numbers that are assumed to have unlimited number of significant figures are considered to be known with "absolute" certainty. *You do not need to consider or count significant figures for exact numbers.*

The following are considered **exact numbers** for CH100:

1. Counting numbers, such as:
   - The number of sides on a square
   - The number of apples on a desktop
2. Defined numbers such as those used for conversion factors, such as:
   - 100 cm = 1 m, 12 in = 1 ft, 1 in = 2.54 cm
   - 1 kg = 1000 g, 1 LB = 16 oz
   - 1000 mL = 1 L; 1 gal = 4 qts.
   - 1 minute = 60 seconds
3. Numbers or constants defined in equations, such as:
   - **y = 3x + 15** (both the "3" and the "15" are exact numbers)

Converting between Unit Systems

- Converting units from one unit system to another (especially within the Metric system) can appear daunting at first glance. However, with a little guidance, and a lot of practice, you can develop the necessary skill set to master this process.
- To begin, here is a simple mnemonic to guide you through the unit conversion process:
  1. **Eliminate**
  2. **Replace**
  3. **Relate**
- All unit conversions, regardless of how complex they appear, involve these 3 simple steps. In the following sections, you will be stepped through the unit conversion process using these 3 words as a guide.

Example: Unit Conversion

1. Convert 25.0 m to cm

2. Convert 1.26 g to kg

Metric Prefixes

Table 2.2 The Commonly Used Prefixes in the Metric System

Prefix	Symbol	Meaning	Power of 10 for Scientific Notation
mega	M	1,000,000	10^{6}
kilo	k	1000	10^{3}
deci	d	0.1	10^{-1}
centi	c	0.01	10^{-2}
milli	m	0.001	10^{-3}
micro	μ	0.000001	10^{-6}
nano	n	0.000000001	10^{-9}

Temperature Scales

The 2 traditional temperature scales, Fahrenheit and Celsius, were originally defined in terms of the physical states of water at sea level:

1. Fahrenheit Scale, °F
   - For water: freezing point = 32°F, boiling point = 212°F
2. Celsius Scale, °C
   - For water: freezing point = 0°C, boiling point = 100°C
   - 1 Celsius temperature unit is larger than 1 Fahrenheit unit

The SI unit for temperature is a variant of the Celsius scale

3. Kelvin Scale, K
   - For water: freezing point = 273 K, boiling point = 373 K
   - The Kelvin temperature unit is the same size as the Celsius unit

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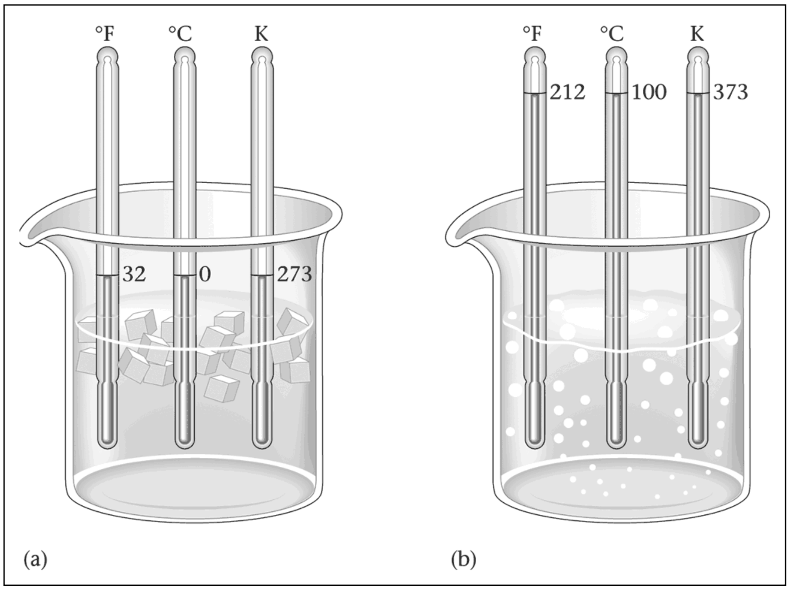
Temperature of ice water and boiling water.



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Unit Conversion & Temperature Scales

Unit conversion involving temperature is tricky since the "zero" value for each scale is different and thus requires accounting for this "offset" between the various scales. *At 0°C, the Kelvin scale has a 273.15 unit "head start" and the Fahrenheit scale has a 32 unit head start*

1. The temperature span between the freezing and boiling points of water reveal the relation between the temperature scale increments:

$$100^oC = 100K = 180^oF$$

2. However, the zero points are different as evident for the freezing point for water:

$$0^oC = 273.15K = 32^oF$$

3. The relations between the temperature scales:

   a. Celsius to Fahrenheit: $T_{^oF} = T_{^oC}\left(\frac{180^\circ F}{100^\circ C}\right) + 32^\circ F$

   b. Celsius to Kelvin: $T_K = T_{^oC}\left(\frac{100K}{100^\circ C}\right) + 273.15K$

Mass

1. Mass is the quantity of matter in a substance
2. Mass is measured in units of grams
3. Mass does <u>not</u> reflect how much volume something has
4. The kilogram (kg) unit is the preferred unit of mass in the SI system.
   a. 1 kilogram is equal to the mass of a platinum-iridium cylinder kept in a vault at Sevres, France.
   b. 1 kg has the weight equivalent (on Earth) of 2.205 lb

Conservation of Mass: The total quantity of mass is never created nor destroyed during a chemical process

Distinguishing Mass vs. Weight

- The terms mass and weight are commonly used interchangeably but they are fundamentally different!
- The following are some important differences between mass and weight:

1. Mass is a fundamental property of matter, *it is the amount of "stuff" in an object*
2. Mass represents an object's inertia *(tendency to resist change in motion)*
3. Mass is the same everywhere in the universe
4. SI Units of mass are kilograms (kg)

1. Weight is the effect (or force) of gravity on an object's mass
2. Weight depends on location (& local gravity)
3. Weight is <u>not</u> a fundamental property of matter
4. SI units of weight are newtons (N)
5. USCS units are pounds (lb)

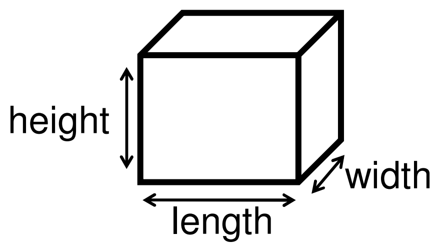
Volume

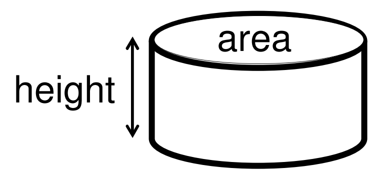
- Volume is the 3-dimensional space that an object occupies
- Volume Units:
  - The SI unit for volume is the cubic meter, or m^3 (meters x meters x meters)
  - The more common metric unit of volume is the Liter (L)

$$1\ m^3 = 10^3\ L$$

  - In the laboratory, the milliliter (mL) is often more convenient

$$1\ mL = 10^{-3}\ L$$





Note: mass and volume are not the same thing *(try not to confuse them…).* Two objects with the same volume *(e.g. a pillow & a sack of potatoes can have different masses and vice versa)*

Density

Density is a property of matter representing the mass per unit volume

- For equal volumes, a denser object has greater mass
- For equal masses, a denser object has smaller volume

Commonly used units:

1. Solids = g/cm^3 (Note: 1 cm^3 = 1 mL)
2. Liquids = g/mL
3. Gases = g/L

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

Useful Notes on Density:

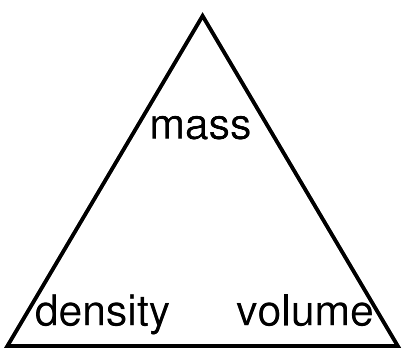
- Volume of a solid can be determined by water displacement
- Density of matter in various states: solids > liquids >>> gases *(exception: water)*
  - In a heterogeneous mixture, the denser matter will tend to sink to the bottom

Table 2.8 Densities of Various Common Substances at 20 °C

Substance	Physical State	Density (g/cm^3)
oxygen	gas	0.00133*
hydrogen	gas	0.000084*
ethanol	liquid	0.785
benzene	liquid	0.880
water	liquid	1.000
magnesium	solid	1.74
salt (sodium chloride)	solid	2.16
aluminum	solid	2.70
iron	solid	7.87
copper	solid	8.96
silver	solid	10.5
lead	solid	11.34
mercury	liquid	13.6
gold	solid	19.32

*At 1 atmosphere pressure

Manipulating the Density Equation



$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

$$\text{Volume} = \frac{\text{Mass}}{\text{Density}}$$

$$\text{Mass} = \text{Density} \times \text{Volume}$$