

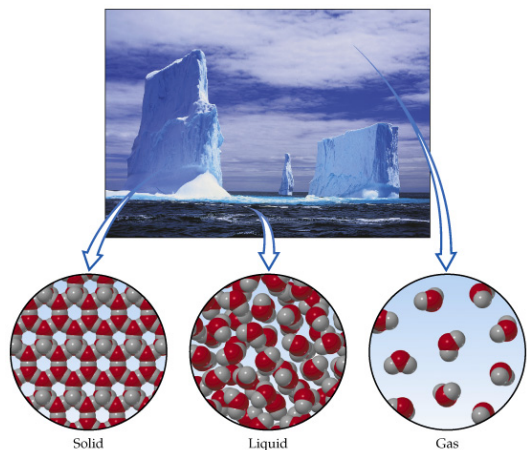
Chemistry 1314 – Lecture Outline and Notes

Text: Chemistry – The Central Science, 10th edition, Brown, LeMay and Bursten

Chapter 1. Introduction: Matter and Measurement

- Overview of Chemistry
- Why is it called the central science?
- Chemistry is the study of materials and changes that they undergo. All sciences work with materials. Thus central to all sciences.
- Examples of chemistry – All things made of chemicals
- The chemical Industry:
 - Oil
 - Plastics
 - Drug development
 - Computers
 - Cooking
- Matter
- Material is another name for matter.
 - Matter is formed from chemicals – either pure or a mixture
- Three states of matter: Solid, Liquid and Gas
- GAS
 - No fixed volume or shape – Conforms to container
 - Can be compressed or expanded
- LIQUID
 - Fixed volume – but shape conforms to the container it is in
 - Essentially no expansion or compression capability
- SOLID
 - Fixed volume and shape
 - Essentially no expansion or compression capability

The case for ice:



The solid form (ice) is notable in that it is lighter (less dense) than the liquid form (water).

Matter can be a pure substance or a mixture

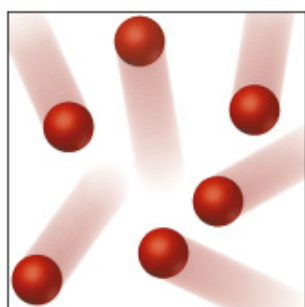
Pure substance:

- Elements – cannot be broken down into simpler substances
- Compounds – composed of 2 or more elements combined in a specific combination. The composition is always the same – Law of Constant Composition (or Law of Definite Proportions).
 - Examples: water, carbon dioxide, ethylene glycol, ethane, methane (natural gas)

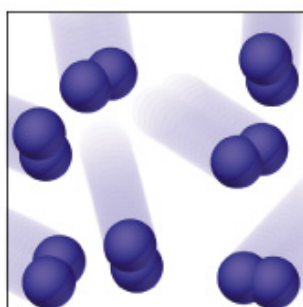
Mixtures – combination of 2 or more substances, but each substance maintains its own chemical identity.

- Homogeneous – same throughout
- Heterogeneous – different throughout

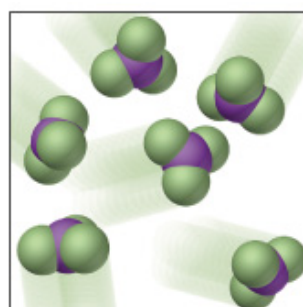
Forms of matter



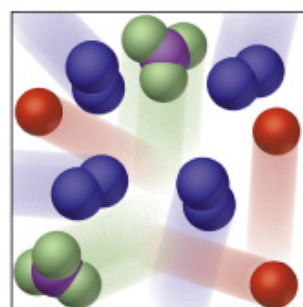
(a) Atoms of an element



(b) Molecules of an element

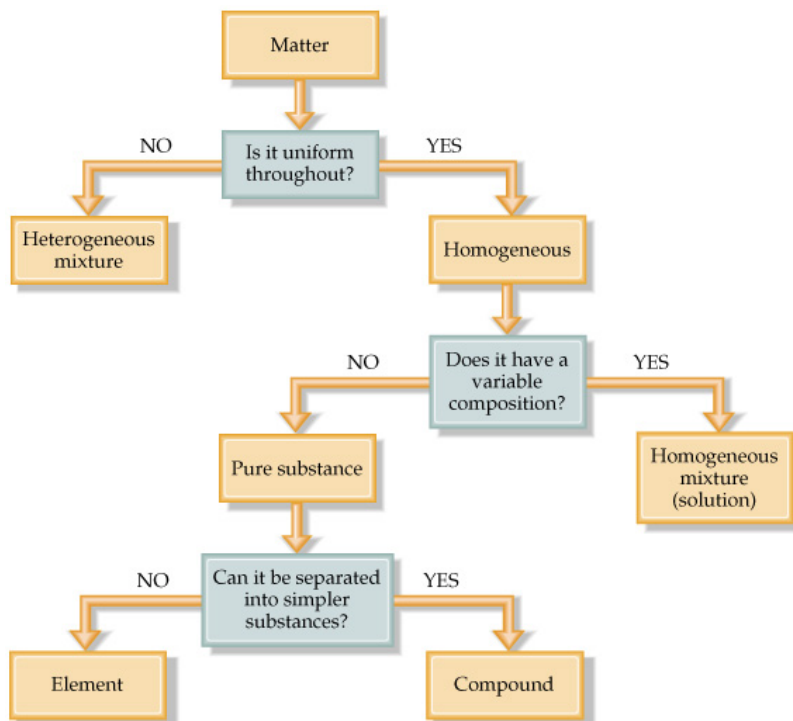


(c) Molecules of a compound



(d) Mixture of elements and a compound

Flow chart to determine nature of the matter:



Elements:

- Introduction to the Periodic Table of Elements
 - Need to become familiar with many of the elements, especially the more common ones.
 - Expect symbol recognition and proper spelling

TABLE 1.2 Some Common Elements and Their Symbols

Carbon	C	Aluminum	Al	Copper	Cu (from <i>cuprum</i>)
Fluorine	F	Barium	Ba	Iron	Fe (from <i>ferrum</i>)
Hydrogen	H	Calcium	Ca	Lead	Pb (from <i>plumbum</i>)
Iodine	I	Chlorine	Cl	Mercury	Hg (from <i>hydrargyrum</i>)
Nitrogen	N	Helium	He	Potassium	K (from <i>kalium</i>)
Oxygen	O	Magnesium	Mg	Silver	Ag (from <i>argentum</i>)
Phosphorus	P	Platinum	Pt	Sodium	Na (from <i>natrium</i>)
Sulfur	S	Silicon	Si	Tin	Sn (from <i>stannum</i>)

Properties of Matter: (unique characteristics that distinguish that substance).

- Physical Properties – measured without changing the identity and composition of the substance (examples: color, odor, density, melting point, hardness)
- Chemical Properties – how a substance may change or react to form other substances (example: flammability)

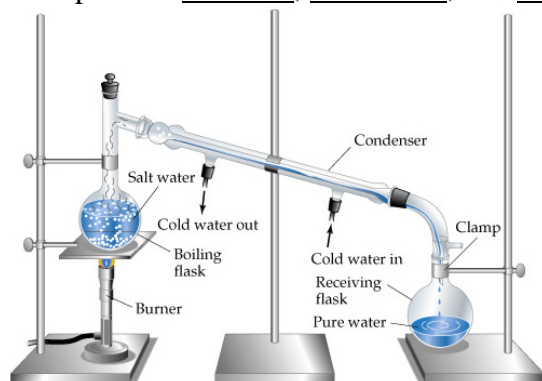
Properties can be further categorized:

- Intensive Properties – do not depend on amount of sample (“internal” or “intrinsic” characteristic of the substance (examples: temperature, melting point, density)
- Extensive Properties – depend on quantity of the sample (examples: volume and mass).

Changes that matter can go through:

- Physical Changes – change in physical appearance, but not its composition. Changes of State
- Chemical Changes – chemical reactions, the substance is transformed into a chemically different substance.

Separation of Mixtures – in a mixture, the chemical components retain their individual properties, and those properties can be used to separate the components. Some common examples are filtration, distillation, and chromatography.



Example of distillation.

Introduction of the Scientific Method

- Data Collection – observations and experiments
- Identify patterns and trends
- Formulate a hypothesis and test
- If an explanation is able to explain most phenomena then it will become a law

(For further reading if interested: Thomas S. Kuhn *The Structure of Scientific Revolution*)

Units of Measurement

Units for scientific measurement are based on the metric system. The internationally agreed upon units for the measure of physical quantity are called the SI units.

Physical Quantity	Name of Unit	Abbreviation
Mass	Kilogram	kg
Length	Meter	m
Time	Second	s ^a
Electric current	Ampere	A
Temperature	Kelvin	K
Luminous intensity	Candela	cd
Amount of substance	Mole	mol

^aThe abbreviation sec is frequently used.

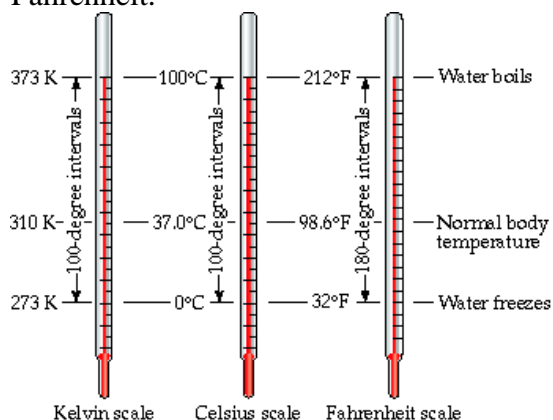
Select prefixes used in the metric system:

giga	G	10 ⁹
mega	M	10 ⁶
kilo	k	10 ³
hecto	h	10 ²
deka	da	10
BASE UNIT		
deci	d	10 ⁻¹
centi	c	10 ⁻²
milli	m	10 ⁻³
micro	μ	10 ⁻⁶
nano	n	10 ⁻⁹
pico	p	10 ⁻¹²
femto	f	10 ⁻¹⁵

NOTE: Mass is the amount of material in an object. Weight is a force (the mass acted upon by gravity).

Temperature:

The SI unit is Kelvin – show how related to Celsius. Introduce formula for converting Celsius to Fahrenheit.



$$K = C + 273.15$$

$$C = 5/9(F - 32)$$

$$F = 9/5C + 32$$

Other quantities can be derived from the SI Units

- Basic unit of volume is m^3 . More useful measurements are $cm^3 = 1 \text{ ml}$ and $dm^3 = 1 \text{ l}$.
- Density (the amount of mass in a unit volume) is $\text{Density} = \text{mass}/\text{volume}$ and usually expressed as g/cm^3 (or g/ml).

Uncertainty in Measurement

Two kinds of numbers:

- Exact numbers – those whose values are known exactly
- Inexact numbers – those whose values contain some uncertainty.
 - Measured quantities are always inexact.

Precision and Accuracy:

- Precision – how closely individual measurements agree with one another
- Accuracy – how closely individual measurements agree with the correct value
 - We take several measurements and must make sure the instruments used for measurement are in good order.

Significant figures:

- Measured quantities are usually reported in such a way that the last digit is uncertain (if it is ± 1 , otherwise state the uncertainty).
- All digits of a measured quantity are called significant figures. The greater the number of significant figures, the greater the certainty implied for the measurement.

Examples:

- 1) 25.0 m (3) and 25.00 m (4) – numbers after the decimal are significant
- 2) 1.0005 (5) – zeros between digits significant
- 3) 6.023×10^{23} (4) – exponential format signifies the number of significant figures
- 4) 10000 (1, 2, 3, 4, or 5) – these zeros could or could not be significant
- 5) 0.0004100 (4) – zeros that fall at the end of a decimal number are significant, but the ones before the digits but after the decimal are just place holders and not significant.

In calculations – the least certain measurement limits the certainty of the calculation. The answer should contain only one uncertain digit.

Addition and subtraction – the result can have no more decimal places than the measurement with the fewest number of decimal places

- 1) $20.33 + 5.1 = 25.43 = 20.4$
- 2) $4.6 + 38 + 2.555 = 45.155 = 45$

Multiplication and division – result must be reported with the same number of significant figures as the measurement with the fewest significant numbers

- 1) $5.92 \times 4.6 = 27.232 = 27$
- 2) $21.3/4 = 5.325 = 5$

Round the following numbers to a whole number

- 1) $20.40 = 20$
- 2) $20.60 = 21$
- 3) $20.5000000001 = 21$
- 4) $20.50 = 20$
- 5) $21.50 = 22$

When the digit to be removed is 5, my method is to round up if odd and down if even

In calculation that require more than one step, carry additional digits with the calculation and round at the end.

Dimensional Analysis – in calculation, carry the units along to make sure that the calculation is correct.

For example: To convert 2.6 in to cm. We know that 1 in = 2.54 cm

$$2.6 \text{ in} \times 2.54 \text{ cm/in} = 6.604 \text{ cm} = 6.6 \text{ cm}$$

The relation between inches and centimeters is a conversion factor.

We can use more than one conversion factor – the goal is to get from the units the number is currently in to units that we desire.

Convert 3.4 miles to km

$$3.4 \text{ miles} \times 5280 \text{ ft/mile} \times 12 \text{ in/ft} \times 2.54 \text{ cm/in} \times 1 \text{ km}/100,000 \text{ cm} = 5.4718 = 5.5 \text{ km}$$

Conversion of volumes:

Suppose that we wish to know the mass in grams of 2.00 cubic inches of gold given that the density of gold is 19.3 g/cm^3

Conversion factors that we should know:

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

Calculation:

$$(2.00 \text{ in}^3)(2.54 \text{ cm/in})^3(19.3 \text{ g gold/1 cm}^3) = (2.00 \text{ in}^3)(16.387 \text{ cm}^3/\text{in}^3)(19.3 \text{ g gold/cm}^3) \\ = 632.54067 = 633.$$

To keep in mind:

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?