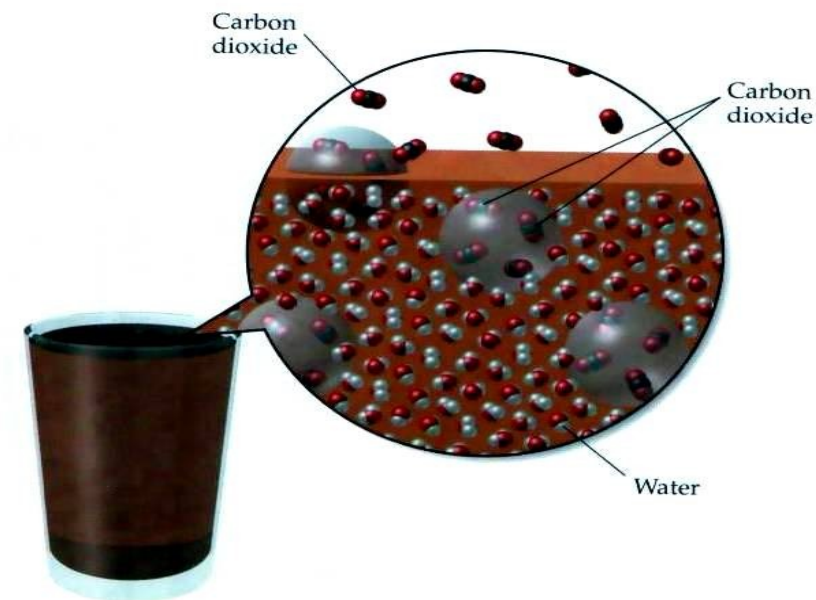
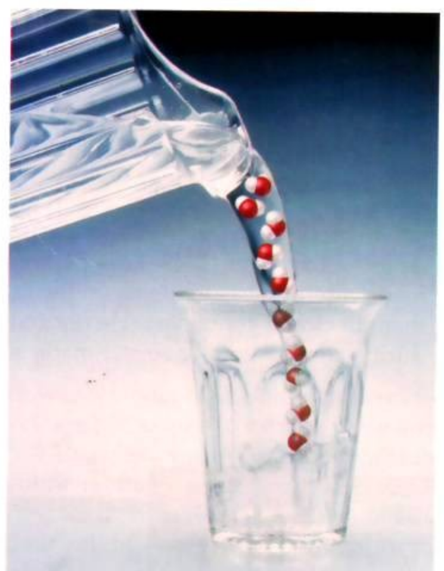
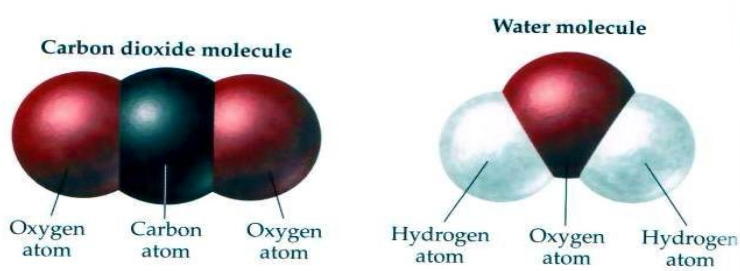
Introduction to Chemistry

Instructed by Taing You Nam, Ph.D





◀ Chemists are interested in knowing why ordinary things, such as water, are the way they are. When a chemist sees a pitcher of water, she thinks of the molecules that compose the liquid and how they determine its properties.



Background

In the early years of civilization, chemistry was unknown. A few men, then known as alchemists, who were often thought to be magicians rather than scientists, studied metals and some non-metal substances. The people knew about the metallurgy and glassmaking. They could melt metals in furnaces and they used gold and silver for jewelry and ornaments. Many excelled in the arts and crafts but little was known about chemical changes and reactions.

It was not until 1774, when Joseph Priestley discovered oxygen that chemistry was recognized as a science. Then Antoine Lavoisier showed that oxygen, a part of air, was responsible for combustion. The words of John Dalton of England, Amadeo Avogadro of Italy, Friedrich Kekule of Germany, and Dmitri Mendeleyev of Russia and many others contributed to the progress of chemistry. All these and growth of industries advanced chemistry further.

Important of Chemistry

There are three reasons to study chemistry. First, chemistry has important practical application, for example, the development of lifesaving drugs.

Second, chemistry is an intellectual enterprise, a way of explaining our material world. Example, when Rosenberg and his coworkers saw that cell division in the bacteria had ceased, they systematically looked for the chemical substance that caused it to cease. They sought a chemical explanation for the occurrence.

Finally, chemistry figures prominently in other fields. For example, Rosenberg's experiment began as a problem in biology; through the application of chemistry it led to an advance in medicine.

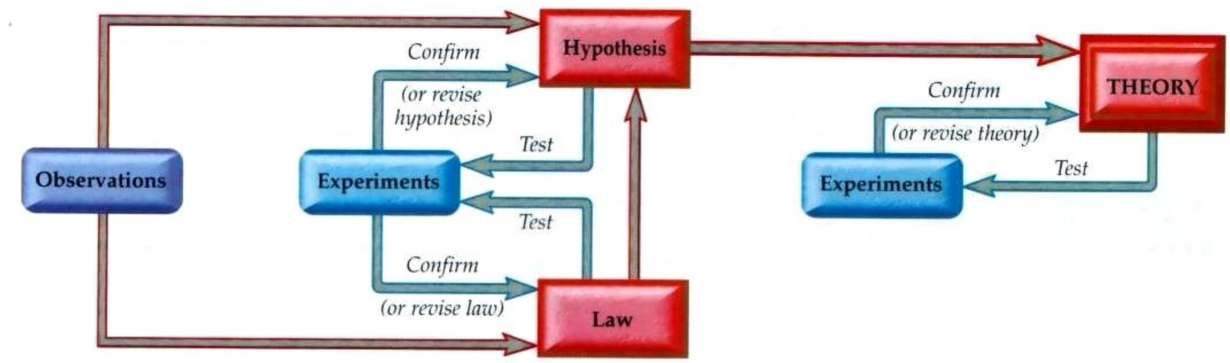
All of the objects around us – book, pen or pencil, and the things of nature such as rocks, water, and plants and animal substances – constitute the matter of the universe. Each of the particular kind of matter, such as certain kind of paper or plastic or metal, is referred to as a material.

Chemistry is needed in practices and industries – metallurgy, manufacture of drugs, dyes, textiles, ceramics, fertilizers and others; in agriculture and in medicine. The changing of one material to another, separation of materials, identification of substances and combining of materials to form new ones – all involve chemistry. Chemistry is concerned with nature of fire, water, air and the earth, the original four substances of ancient times. At home, the care of metal objects and furniture use chemistry.

Chemistry can be defined as the science that deal with the composition and properties of materials, the changes in the composition of materials and the energy transformations that accompany such changes. When chemistry applies to life processes is called **bio-chemistry** and the combination of the principles and methods of physics and chemistry is known as **physical chemistry**.

Scientific Method

Chemists are scientists and use Scientific method – a way of learning that emphasizes observation and experimentation – to understand the world. The scientific method has keys characteristic: observation of nature, formulation of hypotheses, testing of hypotheses by experiment, and formulation of laws and theories.



The scientific method

Observation. Some observations are simple, requiring nothing more than naked eye. Other observations emerge from experiments that rely on the use of increasingly sensitive instrumentation.

Hypothesis. Observation often lead scientists to formulate a hypothesis, a tentative interpretation or explanation of the observation. Hypotheses are tested by experiments, highly controlled observation designed to validate or invalidate hypotheses.

Scientific law. Scientific law is a brief statement that synthesizes past observations and predicts future ones. Sometimes a number of similar observations can lead the development of a scientific law. Example, law of conservation of mass.

Theory. Theories try to provide a broader and deeper explanation for what we observe in terms of underlying causes. One or more well-established hypotheses may form the basis for a scientific theory. Example, atomic theory explains the law of conservation of mass.

MEASUREMENT in CHEMISTRY

Instructed by Taing Nam You, Ph.D.

I. Scientific Notation

Scientists often use scientific notation to write the very large and the very small numbers compactly.

0.000 000 000 1 becomes 1×10^{-10}

14,000,000,000 becomes 1.4×10^{10}

A number written in scientific notation consists of a decimal part, a number that is usually between *1* and *10*, and an exponential part, *10* raised to an exponent, *n*.

1.2×10^{-10} ← exponent *n*

↑ decimal part ↑ exponential part

A positive exponent (*n*) means 1 multiplied by 10 *n* times.

$10^0 = 1$

$10^1 = 1 \times 10$

$10^2 = 1 \times 10 \times 10 = 100$

$10^3 = 1 \times 10 \times 10 \times 10 = 1{,}000$

A negative exponent (-n) means divided by 10 *n* times.

$10^{-1} = \frac{1}{10} = 0.1$

$10^{-2} = \frac{1}{10 \times 10} = 0.01$

$10^{-3} = \frac{1}{10 \times 10 \times 10} = 0.001$

To express a number in scientific notation:

1. Move the decimal point to obtain a number between 1 and 10.
2. Write the result from step 1 multiplied by 10 raised to the number of places you moved the decimal point.
   - The exponent is positive if you moved the decimal point to the left.
   - The exponent is negative if you moved the decimal point to the right.

     293,168,000 people = 2.93168×10^8 people

     $0.000038 = 3.8 \times 10^{-5}$

II. Significant Figures: Writing Number to Reflect the Precision

Significant figures are those digits in a measured number (or result of a calculation with measured numbers) that include all certain digits plus a final one having some uncertainty.

A reported number is 45.87*2* ← estimated

↑

certain

Precision refers to the closeness of a set of values obtained from identical measurements of a quantity.

Accuracy refers to the closeness of a single measurement to its true value.

To determine the **number of significant figures** (digits) in a number follow these rules:

1. All non-zero digits are significant.

   1.05.1 0.0110

2. Interior zeros (zeros between two numbers) are significant.

   4.0208 50.1

3. Trailing zeros (zeros after a decimal point) are significant.

   5.10 3.00

4. Leading zeros (zeros to the left of the first non-zero number) are not significant. They only serve to locate the decimal point.

   0.0005 has only one significant digit.

5. Zeros at the end of a number but before a decimal point are ambiguous and should be avoided by using scientific notation.

   Does 350 have two or three significant figures?

   3.5×10^2 has 2 significant figures.

   3.50×10^2 has 3 significant figures.

Practice: How many significant figures are there in each of the following numbers?

1. 0.0035
2. 1.080
3. 2371
4. 2.97×10^5
5. 100,000

Exact numbers have an *unlimited number of significant figures*. Exact numbers originated from three sources:

- From the accurate counting of discrete objects.
  For example, 3 atoms means 3.00000… atoms
- From *defined quantities*, such as the number of centimeters in 1 m. because 100 cm is defined as 1 m, 100 cm = 1 m means 100.00000… cm = 1.0000000… m
  Note that some conversion factors are defined quantities while others are not.
- From integral numbers that are part of an equation.
  For example, in the equation, radius $= diameter/2$, the number 2 is exact and therefore has an unlimited number of significant figures.

III. Significant Figures in Calculation

Multiplication and Division. In multiplication or division, the result carries the same number of significant figures as the factor with the fewest significant figures.

5.02	x	89.665	x	0.10	=	45.0118	=	45
(3 sig. fig.)		(5 sig. fig.)		(2 sig. fig.)				(2 sig. fig.)

5.892	÷	6.10	=	0.96590	=	0.966
(4 sig. fig.)		(3 sig. fig.)				(3 sig. fig)

Addition and Subtraction. In addition or subtraction, the result carries the same number of decimal places as the quantity carrying the fewest decimal places.

$$\begin{array}{r} 5.74 \\ 0.823 \\ +\ 2.651 \\ \hline 9.214 \end{array} = 9.21 \qquad \begin{array}{r} 4.8 \\ -\ 3.965 \\ \hline 0.835 \end{array} = 0.8$$

Note: For multiplication and division, the quantity the quantity with the fewest significant figures determines the number of significant figures in the answer.

For addition and subtraction, the quantity with fewest decimal places determines the number of decimal places in the answer.

Practice a. 0.987 + 125.1 – 1.22 b. 0.765 – 3.449 – 5.98

Rounding. When rounding to the correct number of significant figures:

Round down if the last (or leftmost) digit dropped is *4 or less*;

Round up if the last (or leftmost) digit dropped is *4 or more*.

2.33 rounds to 2.3	Be certain to use the last (or leftmost) digit being dropped
2.34 rounds to 2.3	to decide in which direction to round – ignore all digits to
2.35 rounds to 2.4	the right of it.
2.37 rounds to 2.4	2.349 rounds to 2.3

Calculation involving both Multiplication/Division an Addition/Subtraction do the steps in parentheses first; determine the number of significant figures in the intermediate answer; then do the remaining steps.

3.489 x (5.67 – 2.3)

first 5.67 – 2.3 = 3.37, next 3.489 x 3.<u>3</u>7 = 11.758 = 12

Practice

a. 6.78 x 5.903 x (5.489 – 5.01) b. 19.667 – (5.4 x 0.916) c. (4.58 ÷ 1.239) – 0.578

IV. The Basic Units of Measurement

Important SI Standard Units

Quantity	Unit	Symbol
length	meter	m
mass	kilogram	kg
time	second	s
temperature*	kelvin	K

SI Prefix Multipliers

Prefix	Symbol	Multiplier	
tera-	T	1,000,000,000,000	(10^{12})
giga-	G	1,000,000,000	(10^{9})
mega-	M	1,000,000	(10^{6})
kilo-	k	1,000	(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})
pico-	p	0.000000000001	(10^{-12})
femto-	f	0.000000000000001	(10^{-15})

Some Common Units and Their Equivalents

Length

1 kilometer (km) = 0.6214 mile (mi)

1 meter (m) = 39.37 inches (in.)
= 1.094 yards (yd)

1 foot (ft) = 30.48 centimeters (cm)

1 inch (in.) = 2.54 centimeters (cm) (exact)

Mass

1 kilogram (kg) = 2.205 pounds (lb)

1 pound (lb) = 453.59 grams (g)

1 ounce (oz) = 28.35 grams (g)

Volume

1 liter (L) = 1000 milliliters (mL)
= 1000 cubic centimeters (cm^3)

1 liter (L) = 1.057 quarts (qt)

1 U.S. gallon (gal) = 3.785 liters (L)

V. Factor-Label Method

The factor-label or dimensional analysis method is simple but powerful technique that can be applied to a great variety of calculation problems. In this method the units (labels) of the quantities involved are used as guide in setting up a calculation. The method is based on the use of conversion factors and on the idea that units or dimensions of various quantities can be handled algebraically in the same way as numbers are handled. Units are multiplied, divided, and canceled like any other algebraic quantity.

Remember:

1. Always write every number with its associated unit. Never ignore units; they are critical.
2. Always include units in your calculations, dividing them and multiplying them as if they were algebraic quantities. Do not let units magically appear or disappear in calculations. Units must flow logically from beginning to end.

Conversion factors are constructed from equalities such as the following:

2.54 cm = 1 in   1000 g = 1 kg   24 h = 1 day

We can write conversion factors as $\frac{2.54\ cm}{1in}$ $\frac{1000\ g}{1kg}$ $\frac{24\ h}{1\ day}$

Consider converting 17.6 in. to centimeters. 17.6 in x $\frac{2.54\ cm}{1in}$ = 44.7 cm

Information given x **conversion factor(s) = information sought**

Convert 194 cm to ft. Solution map: cm → in. → ft

$\frac{1\ in.}{2.54\ cm}$ $\frac{1\ ft}{12\ in.}$

$$194\ cm\ x\ \frac{1\ in.}{2.54\ cm}\ x\ \frac{1\ ft}{12\ in.} = 6.3648\ ft = 6.36\ ft$$

Convert 0.75 L to cups, 4 cups = 1 quart (qt), 1 L = 1.057 qt

Convert 1055 m to mile (mi), 1 km = 0.6214 mi

Convert 15, 615 dm^3 to $in.^3$

VI. Calculation for Temperature and Density

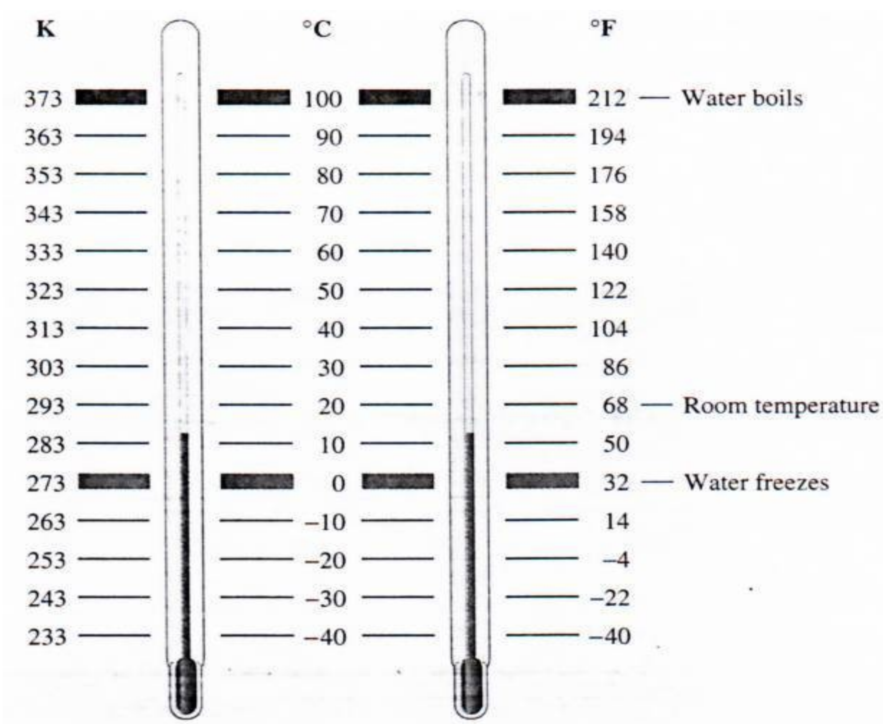
Temperature. Temperature is difficult to define precisely. It is a measure of “hotness”. A hot object place nest to a cold one becomes cooler, while the cold object becomes hotter. A thermometer is a device for measuring temperature.

The Celsius scale (formerly the centigrade scale) is the temperature scale in general scientific use. On this scale, the freezing point of water is 0˚C and the boiling point is 100˚C. The Kelvin (K) is a unit on an absolute temperature scale. The Celsius and the Kelvin scales have equal-size units, but 0˚C is equivalent to 273.15K. Thus, it is easy to convert from one scale to the other, using the formula $K = {}^\circ C + 273.15$

The Fahrenheit scale is at present the common temperature scale in the United States. The equation for converting from degrees Fahrenheit to degrees Celsius is

$^\circ C = (^\circ F - 32\ ^\circ F) \times \frac{5^\circ C}{9^\circ F}$ or $^\circ F = (1.8 \times {}^\circ C) + 32$ $^\circ C = \frac{^\circ F - 32}{1.8}$

Convert 134 ˚F to ˚C and K.



Density. Density (d) of an object is its mass per unit volume.

$d = \frac{m}{V}$ where m = mass, V = volume

A piece of titanium has mass of 9.0 g and volume of 2 mL. What is the density of titanium?