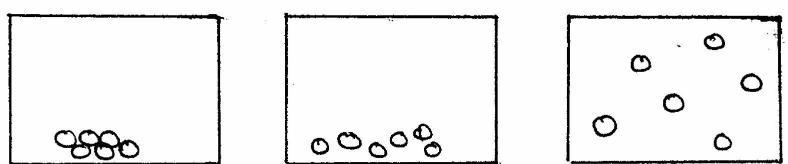
Chapter 1 : Matter and Measurement

Chemistry : The science that describes matter, its chemical and physical properties, the chemical and physical changes that matter undergoes, and the energy changes that accompany these processes.

Remember – chemistry is an experimental science based on observations. Theories have been developed by sceentists to
(1) explain phenomena
(2) predict phenomena that have not yet happened.

How do we classify matter?
state of matter: solid liquid gas

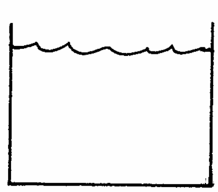


solid liquid gas

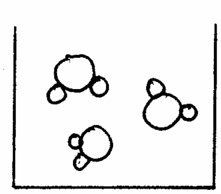
Kinetic molecular theory – all matter consists of tiny particles (atoms, molecules, ions) in constant motion (above 0 K)
solids: particles packed close together & vibrate in place
liquids: particles are close but can move past each other (fluid)
gases: particles are far apart, move quickly, occupy full volume of container (fluid)

Chemistry is thought to be a complex subject to learn because one has to think at several different levels at once:

consider water,



macroscopic
(observations, experimentations)



particulate view
molecules & atoms

H_2O

symbolic view

pure substance: has unique set of properties and cannot be separated by physical means. A compound or element are pure substances. (more later)

mixture: a conglomeration of pure substances.
- heterogeneous mixture: soil (rocks, clay, water, worms, bacteria, gases etc.)
  sugar & salt, sand & water
  you can easily see the uneven texture of material
- homogeneous mixture: solutions (salt in water, brass-alloy)
- can be separated by physical means into pure substances.
- for example: you can boil the water off a salt water solution, leaving the salt behind.

element: a pure substance composed of only 1 kind of small particle
- 116 elements are known - see periodic table.

Some Common Elements and Their Symbols

Symbol	Element	Symbol	Element	Symbol	Element
Ag	silver (*argentum*)	F	fluorine	Ni	nickel
Al	aluminum	Fe	iron (*ferrum*)	O	oxygen
Au	gold (*aurum*)	H	hydrogen	P	phosphorus
B	boron	He	helium	Pb	lead (*plumbum*)
Ba	barium	Hg	mercury (*hydrargyrum*)	Pt	platinum
Bi	bismuth	I	iodine	S	sulfur
Br	bromine	K	potassium (*kalium*)	Sb	antimony (*stibium*)
C	carbon	Kr	krypton	Si	silicon
Ca	calcium	Li	lithium	Sn	tin (*stannum*)
Cd	cadmium	Mg	magnesium	Sr	strontium
Cl	chlorine	Mn	manganese	Ti	titanium
Co	cobalt	N	nitrogen	U	uranium
Cr	chromium	Na	sodium (*natrium*)	W	tungsten (*Wolfram*)
Cu	copper (*cuprum*)	Ne	neon	Zn	zinc

know these ⟸

atom is smallest part of an element that contains the same characteristics as the element.

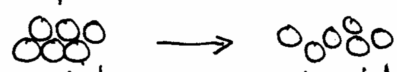
compound: substance containing 2 or more elements in fixed ratio. The smallest part of compound with same characteristics is the

molecule - if the compound is held together by covalent bonds
or formula unit - if the compound is held together by ionic bonds.
(more later)

properties: characteristics that describe samples of matter.
- physical property: observed in the absence of change in composition
  - examples: color, density, hardness, melting point, boiling pt. electrical conductivity, mass, volume.
  - divided into 2 kinds
    - extensive - depends on amount of substance present
      e.g. mass, volume
    - intensive - independent of amount of substance present.
      e.g. color, temperature, density, melting point
- chemical property: a property that matter exhibits as it changes composition, eg magnesium metal can react with molecular oxygen.

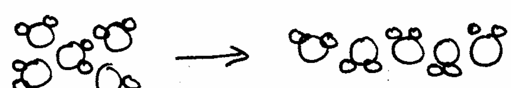
physical change - occurs with no change in chemical composition, but energy can be released or absorbed

(1) lead melting - process needs heat (endothermic)

heat + 

solid → liquid

(2) water freezing - process releases heat (exothermic)

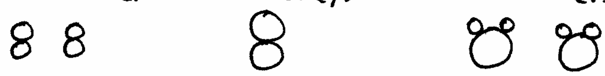


liquid → solid + heat

chemical change: occurs when (1) one or more substances are used up
(2) one are more substances are produced
(3) energy is released or absorbed.

reactants → product

$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$ balanced — Law of Conservation of Matter



Units of Measurement

International System of Units (SI) adopted by National Bureau of Standards includes:

mass	kilogram	kg	
length	meter	m	
time	second	s	absolute T scale
temperature	Kelvin	K	(K = 273 + °C)
amount of substance	mole	mol	
electric current	ampere	A	

Common prefixes:

			example
mega-	M	10^6	Mg = 1×10^6 grams
kilo-	k	10^3	kg = 1×10^3 grams
deci-	d	10^{-1}	dL = 0.1 liter = 100 mL
centi-	c	10^{-2}	cm = 0.01 meter
milli-	m	10^{-3}	mg = 1×10^{-3} grams
micro-	μ	10^{-6}	μm = 1×10^{-6} meter
nano-	n	10^{-9}	ng = 1×10^{-9} grams
pico-	p	10^{-12}	pg = 1×10^{-12} grams

Temperature: we will only be working with °C and K
absolute zero = 0 K = −273.15 °C

Volume: 1 cubic centimeter = 1cm × 1cm × 1cm = $\boxed{1\ cm^3 = 1\ mL}$

Mass vs Weight
mass: measure of quantity of matter a body contains (grams)
weight: measure of gravitational attraction of earth for the body
weight = force = mass × gravitational acceleration, g

length : measured in meters (or cm, mm, etc.)
volume : measured in liters (or mL)

Significant Figures

There are two kinds of numbers in the world

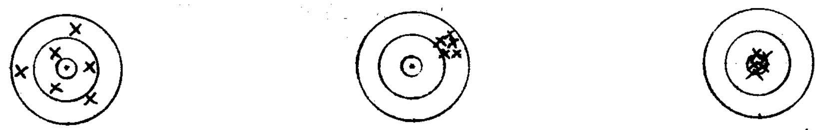
exact : there are exactly 12 eggs in a dozen
I have exactly 10 fingers and 10 toes – no more, no less

inexact : any measurement :

		# of sig. fig.
I quickly measure the width of this paper	: 220 mm	2
If I am more precise	: 216 mm	3
If I am even more precise	: 215.6 mm	4

Note : PRECISION vs ACCURACY

accuracy : refers to how closely a measured value agrees with the correct value
precision : refers to how closely individual measurements agree with each other.



accurate not precise — precise not accurate — accurate and precise

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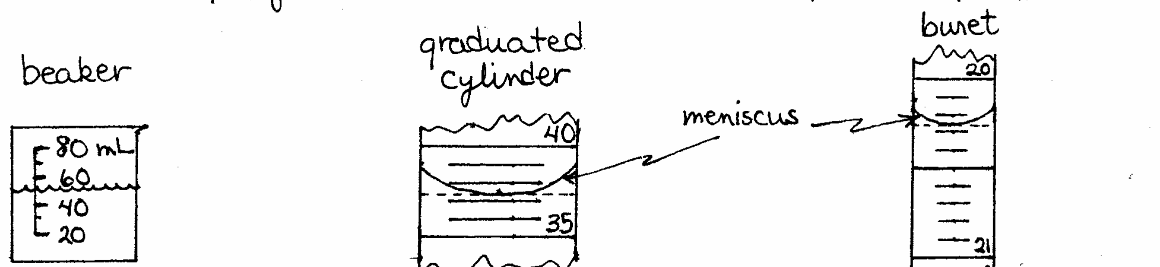
In any measurement, the number of significant figures is important: the number of significant figures is the number of digits believed to be correct by the person doing the measuring. It includes 1 estimated digit.

Let's look at another example – Volume measurement
Volumes of liquids are measured in the laboratory in many different ways :



50 ± 2 mL ($\frac{1}{5}$ of 10 mL)
digit containing error

37.3 ± 0.2 mL
($\frac{1}{5}$ of smallest division = 1 mL)

20.27 ± 0.02 mL
(smallest division = .1 mL)

Rule of Thumb: read measurement to $\frac{1}{5}$ or 0.2 of the smallest division

Volume obtained from the buret has 4 significant figures – the last digit is a best estimate. whereas the volume measured from a beaker only has 2 significant figures. If a person needed only a rough estimate of volume the beaker volume is satisfactory, otherwise one would use the graduated cylinder or better yet the buret.

Rules for significant figures:

(1) leading zeros are never significant
imbedded zeros are always significant
trailing zeros are significant only if the decimal point is specified.

(2) addition or subtraction : last digit retained is set by the first doubtful digit

(3) multiplication or division : answer contains no more sig. fig. than the least accurately known number.

Examples:

	Number of sig. fig.	Scientific Notation
0.00682	3	6.82×10^{-3}
1.072	4	$1.072 \ (\times 10^{0})$
300	1	3×10^{2}
300.	3	3.00×10^{2}
300.0	4	3.000×10^{2}

Addition : (same for subtraction)

$$\begin{array}{r} 4.7832 \\ 1.234 \\ +\ 2.02 \\ \hline 8.0372 \\ \Downarrow \\ 8.04 \end{array}$$

Multiplication : (same for division)

$$\begin{array}{r} 2.8723 \\ \times\ 1.6 \\ \hline 4.59568 \\ \Downarrow \\ 4.6 \end{array}$$

Dimensional Analysis (factor-label or unit factor method)

This is a very powerful mathematical tool for solving problems using unit factors (a factor in which numerator + denominator have different units but represent equivalent amounts ∴ it is equal to unity)

Rounding Off:

When rounding off numbers to a certain number of significant figures – do so to the nearest value.

2.3467×10^4	to 3 significant figures	2.35×10^4
1.612×10^3	to 2 significant figures	1.6×10^3

What happens if there is only a 5. For example

2.35×10^2 to 2 significant figures

2.4×10^2
2.35×10^2
2.3×10^2

How do we round this number.

There is an arbitrary rule: if the number before the 5 is odd, round up
if the number before the 5 is even, ~~round dow~~ leave alone

$\therefore$ 2.35×10^2	to 2 significant figures	2.4×10^2
2.45×10^2	to 2 significant figures	2.4×10^2

Ofcourse,
2.451×10^2 to 2 significant figures 2.5×10^2
since 2.451×10^2 is closer to 2.5×10^2 than 2.4×10^2

2.5×10
2.451×10
2.4×10

I suggest using dimensional analysis for converting units – but using it only as a last resort for problem solving – unless you are very skillful at it. It is useful as a check for correct units.

Examples:

(1) How many centimeters are in 20.0 inches (1 inch = 2.54 cm)

$$? \text{ cm} = 20.0 \text{ inches} \times \underbrace{\left(\frac{2.54 \text{ cm}}{1 \text{ inch}}\right)}_{\text{unit factor}} = 50.8 \text{ cm}$$

(2) Express 6.2 gallons in milliliters
(1 gallon = 4 qt ; 1 qt = 0.946 L ; 1 L = 1000 mL)

$$? \text{ mL} = 6.2 \text{ gallons} \times \frac{4 \text{ quarts}}{1 \text{ gal}} \times \frac{0.946 \text{ L}}{1 \text{ qt}} \times \frac{1000 \text{ mL}}{1 \text{ L}}$$

$$= 2.3 \times 10^4 \text{ mL} \quad \left(\begin{matrix} 4 \text{ quarts} \\ 1000 \text{ mL} \end{matrix} \text{ are exact numbers}\right)$$

(3) Express 6.12×10^3 cm^2 in ft^2.

$$? \text{ ft}^2 = 6.12 \times 10^3 \text{ cm}^2 \times \frac{(1 \text{ inch})^2}{\underbrace{(2.54 \text{ cm})^2}_{= 6.45}} \times \frac{(1 \text{ ft})^2}{\underbrace{(12 \text{ inch})^2}_{= 144}}$$

$$= 6.59 \text{ ft}^2$$

Density and Specific Gravity

density is defined as mass per unit volume

$1 \text{ mL} \equiv 1 \text{ cm}^3$

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

liquids and solids : units : $\frac{g}{mL}$ or $\frac{g}{cm^3}$
gases : units : $\frac{g}{L}$ $(g \cdot L^{-1})$

Example: Given a substance with mass of 472 g with a volume of 97.3 cm^3, what is its density?

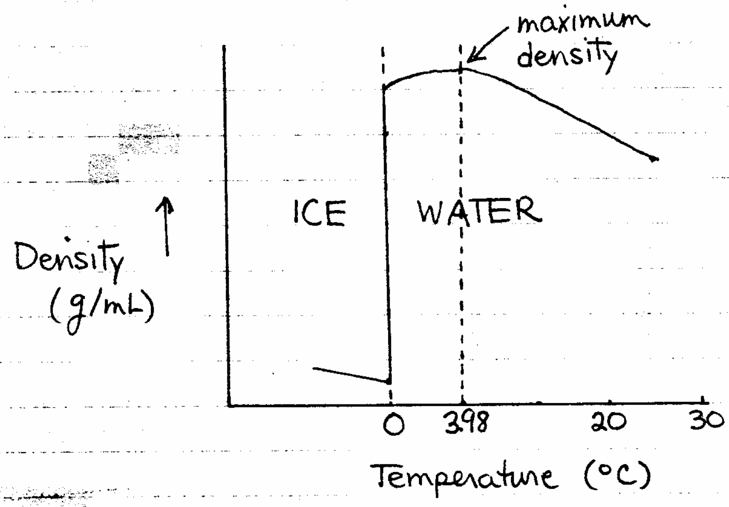
$$D \left(\frac{g}{cm^3}\right) = \frac{472 \text{ g}}{97.3 \text{ cm}^3} = 4.85 \text{ g/cm}^3$$

specific gravity has no units. It is a ratio : $\dfrac{\text{density of substance}}{\text{density of water}}$ ~1.00 g/mL

∴ numerically, density ≡ specific gravity

Example: The density of water is dependent both on temperature and phase.

phase	ice →	water			
T (°C)	0°	0°	3.98	20	30
density (g/mL)	0.917	0.999841	1.00000	0.998203	0.995646



NOT TO SCALE

* Notes on Plasma : another form of the gaseous state.

Plasma

At several thousand degrees gas molecules are traveling with such high energies that on collision they drive off electrons. This process may continue to strip off each of the electrons of an atom until at very high temperatures all of the electrons may be removed. This process produces a high concentration of high-energy charged particles that have become known as *plasma*. This is believed to be the condition of much of the gaseous material in the sun, stars, and interstellar space. Electrons may also be driven off in other processes involving the absorption of a great deal of energy. Today the term plasma is used for a high concentration of these high-energy particles that have lost one or more electrons.

At very high temperatures, 10,000,000°C and higher, the plasma particles have gained so much translational energy that on collision their nuclei approach within 10^{-12} cm of each other where the powerful nuclear attractive forces take effect and cause nuclear fusion processes to occur. These are some of the main reactions that produce the energy and the light elements in the sun and stars. The fusion process will be discussed in Chap. 26.

The study of plasma is one of the new and very important fields of research today.