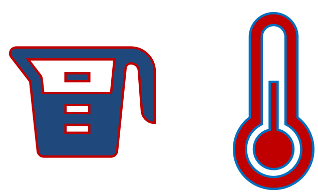
Chapter 1

Measurement and the Proper of Matter



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Chemistry: is the study of matter and the changes it undergoes

- The study of chemistry depends heavily on measurement. For instance, chemists use measurements to compare the properties of different substances

Chemistry Measurements scale

Macroscopic		Microscopic	
can be determined directly		Must be determined by an indirect method (calculation _	
Examples		Examples	
Property Measured	**Instrument(s)**	Number of atoms or	calculated from moles
Length	Meterstick, Scale	Atomic or ionic radius	determined by X-ray diffraction.
Volume	Volumetric Flask	Electron energy levels	measured from atomic spectra.
Mass	Balance	Molecular structure.	techniques like NMR or IR
Temperature	Thermometer	Bond strength	calculated from experimental data

A measured : quantity is usually written as a number with an appropriate unit.

SI Units

Quantity	Length	Mass	Time	Electric current	Temperature	Amount of substance	Luminous intensity
SI Unit	meter	kilogram	second	ampere	kelvin	mole	candela
Symbol	m	kg	s	A	K	mol	cd

SI Unit Prefixes

Tera	giga	mega	kilo	hecto	deca	SI	deci	centi	milli	micro	nano	pico	femto	atto
T	G	M	k	h	da		d	c	m	μ	n	p	f	a
10^{12}	10^{9}	10^{6}	10^{3}	10^{2}	10^{1}		10^{-1}	10^{-2}	10^{-3}	10^{-6}	10^{-9}	10^{-12}	10^{-15}	10^{-18}

Examples

Unit to prefix	prefix to Unit	Prefix to prefix
100m to Km	20 cm to m	250 mm to Km
÷	X	X ÷

Carry out the following conversions

a) 22.6 m to decimeters	b) 25.4 mg to kilograms
Solution: $22.6 \times 10 = 226$ dm	Solution: $25.4 \times 10^{-6} = 2.54 \times 10^{-5}$ kg
c) 556 mL to liters	d) 10.6 kg/m^3 to g/cm^3
Solution: $556 \times 10^{-3} = 0.556$ L	Solution: $10.6 \div 1000 = 0.0106$ g/cm^3
e) 242 lb to milligrams,	f) 68.3 cm^3 to cubic meters
Solution: $242 \times 453{,}600 = 1.0977 \times 10^{8}$ mg (≈ 109,771,200 mg)	Solution: $68.3 \times 10^{-6} = 6.83 \times 10^{-5}$ m^3
g) 7.2 m^3 to liters,	h) 28.3 µg to pounds
Solution: $7.2 \times 1000 = 7200$ L	$28.3 \div (4.536 \times 10^{8}) = 6.24 \times 10^{-8}$ lb

1- A 6.0-ft person weighs 168 lb. Express this person's height in meters and weight in kilograms. (1 lb = 453.6 g; 1 m = 3.28 ft.)

Height: $6.0 \div 3.28 = 1.83$ m
Weight: $168 \times 0.4536 = 76.2$ kg

3. 70 kg, the average weight of a male adult, to pounds

$70 \div 0.4536 = 154$ lb

Mass and Weight

mass	weight
a measure of the amount of matter in an object	is the force that gravity exerts on an object
SI unit Kg	SI unit N

Volume : is an example of a measured quantity with derived units

Unites : SI unit derived m^3	**Liter** : volume occupied by one cubic decimeter.
$1\ cm^3 = (1 \times 10^{-2}\ m)^3 = 1 \times 10^{-6}\ m^3$	$1\ L = 1000\ mL = 1000\ cm^3 = 1\ dm^3$
$1\ dm^3 = (1 \times 10^{-1}\ m)^3 = 1 \times 10^{-3}\ m^3$	$1\ mL = 1\ cm^3$

Density : Mass per unit volume

$$\text{Density} = \frac{mass}{Volume} = \frac{m}{V}$$

SI Unit : $\frac{Kg}{m^3}$

$1\ g/cm^3 = 1\ g/mL = 1000\ kg/m^3$

$1\ g/L = 0.001\ g/mL$

Note

Density is a property that does not depend on the quantity of mass present.

The ratio of mass to volume always remains the same. In other words, V increases as m does. Density usually decreases with temperature.

Examples

1- Calculate its density (in g/mL) if 586 g of the substance occupies 188 mL

$\rho = m / V = 586\ g / 188\ mL = 3.12\ g/ml$

2- The density of methanol is **0.7918** g/mL. Calculate the mass of 89.9 mL of the liquid

$m = \rho \times V$

$0.7918\ g/mL \times 89.9\ mL = 71.2\ g$

3- Aluminum is a lightweight metal (**density = 2.70 g/cm³**) used in aircraft construction, high-voltage transmission lines, beverage cans, and foils. What is its density in kg/m^3?

$1\ g/cm^3 = 1000\ kg/m^3$

$2.70\ g/cm^3 \times 1000 = 2700\ kg/m^3$

4- Ammonia gas is used as a refrigerant in large-scale cooling systems. The density of ammonia gas under certain conditions is 0.625 g/L. Calculate its density in g/cm^3

$1\ L = 1000\ cm^3$

$0.625\ g / 1000\ cm^3 = 0.000625\ g/cm^3$

Temperature Scales

SI unit Kelvin

°F (degrees Fahrenheit),	°C (degrees Celsius)	K (kelvin)
$F = (\frac{9}{5} \times C) + 32$	$C = (F - 32) \times \frac{5}{9}$	$K = C + 273.15$

Examples

Convert the following temperatures to degrees Celsius or Fahrenheit:

(a) **95°F**, the temperature on a hot summer day

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Solution: (95 − 32) × 5/9 = 35.0 °C

(b) **−273.15**°C (theoretically the lowest attainable temperature

Solution: −273.15 × 9/5 + 32 = −459.67 °F

(c) Normally the human body can endure a temperature of **105**°F for only short periods of time without permanent damage to the brain and other vital organs. What is this temperature in degrees Celsius?

Solution: (105 − 32) × 5/9 = 40.6 °C

Scientific Notation

Scientific notation is a way of expressing very large or very small numbers in the form:

$$N = a \times 10^n$$

a	n
is a number greater than or equal to 1 and less than 10	is an integer (positive or negative exponent)

Examples

1. 4,500 = 4.5×10^3 (Here, a = 4.5 and n = 3)
2. 0.0062 = 6.2×10^{-3} (Here, a = 6.2 and n = -3)
3. 72,000,000 = 7.2×10^7

Express the following numbers in scientific notation:

(a) 0.000000027

(b) 356

(c)47,764

(d) 0.096

Express the following numbers as decimals:

(a) 1.52×10^{-2}

(b) 7.78×10^{-8}

Significant Figures

Rule	Explanation	Example	Significant Figures
1. Non-zero digits are significant	Every non-zero-digit count	456	3
2. Zeros between non-zero digits are significant	Middle zeros always count	1002	4
3. Leading zeros are not significant	They only locate the decimal point	0.0045	2
4. Trailing zeros with a decimal are significant	Decimal makes zeros meaningful	45.00	4
		0.0900	3
5. Trailing zeros without a decimal are ambiguous	Usually not significant unless in scientific notation	1200	2 (or 3 if 1.20×10^3)

Examples

1- What is the number of significant figures in each of the following measurements?

(a) 4867 mi
(b) 56 mL
(c) 60,104 tons
(d) 2900 g
(e) 40.2 g/cm^3
(f) 0.0000003 cm
(g) 0.7 min
(h) 4.6×10^{19} atoms

2- How many significant figures are there in each of the following?

(a) 0.006 L
(b) 0.0605dm
(c) 60.5 mg
(d) 605.5 cm^2
(e) 960×10^{-3} g
(f) 6 kg
(g) 60 m.

Accuracy and Precision

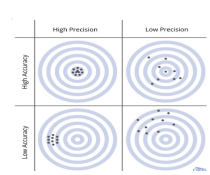
Accuracy	Precision
Tells us how close a measurement is to the true value	refers to how closely two or more measurements of the same quantity agree with one another

Three students are asked to determine the mass of a piece of copper wire. The results of three successive weighing by each student are

(The true mass of the wire is 2.000 g)

Student A	Student B	Student C
1.964 g	1.970 g	2.000
1.971 g	1.972 g	2.002
1.978 g	1.968 g	2.001

Student **B's** results are more precise than those of Student A
Student **C's** results are not only the most precise.



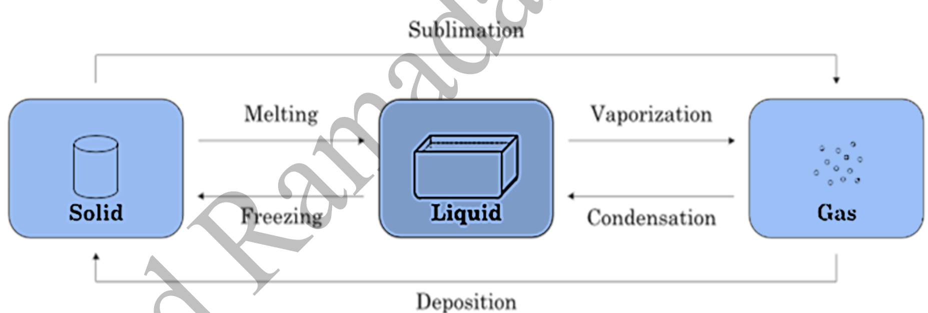
Classifications of Matter

A **substance** is a form of matter that has a definite (constant) composition and distinct properties. Examples are water, ammonia, table sugar (sucrose), gold, and oxygen.
Matter : anything that has mass and occupies space

Matter	Pure substance	Element	a substance that cannot be separated further into simpler substances by chemical methods.	
		Compound	substance composed of atoms of two or more elements chemically united in fixed proportions	
	Mixture	Homogenous	uniform composition is.	
	combination of two or more substances in which the substances retain their distinct identities	Heterogeneous	not uniform. Composition	

States of Matter

term	Solid	Liquid	Gas
Shape	Definite	No definite	No definite
Volume	Definite	Definite	No definite
Compressibility	no	no	Yes



Physical and Chemical Properties of Matter

physical Properties

can be measured and observed without changing the composition or identity of a substance

Extensive properties	intensive properties
depends on how much matter is being	does not depend on how much matter
EX: mass , volume , length	**Ex**: Density , temperature , color , boiling point

Chemical Properties :

can only be observed when a substance is changed into another substance.
(Flammability, corrosiveness, or reactivity with acid.)

Physical Changes

changes in matter that do *not* change the composition of a substance
(Changes of state, temperature, and volume. ice melting or salt dissolving in water)

Chemical Changes

result in new substances
(Combustion, oxidation, and decomposition, rusting of iron.).