General Chemistry for Medical Students 0815141

(2+1 credits)

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Textbook
Chemistry, 10th ed.
Author: Raymond Chang
Publisher: McGrawHill, NY

Office Hours:

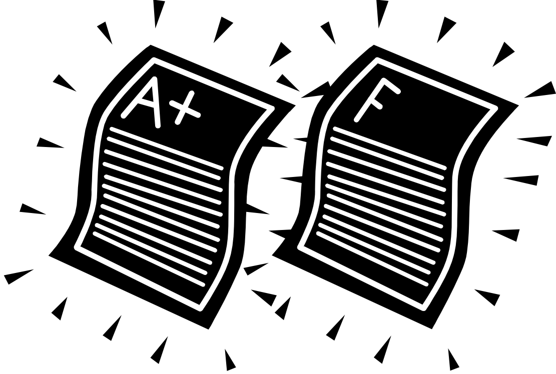
Saturdays 8:30 –10:30 a.m., Sundays 9:30 –10:30 a.m.

Mondays: 9:30 –10:30 a.m., Tuesdays: 9:30 –10:30 a.m.

Wednesdays: 9:30 –10:30 a.m.

and by appointment, Drop ins welcome

Grading Policy:



- Your grade in this course is based on your performance on the following items:

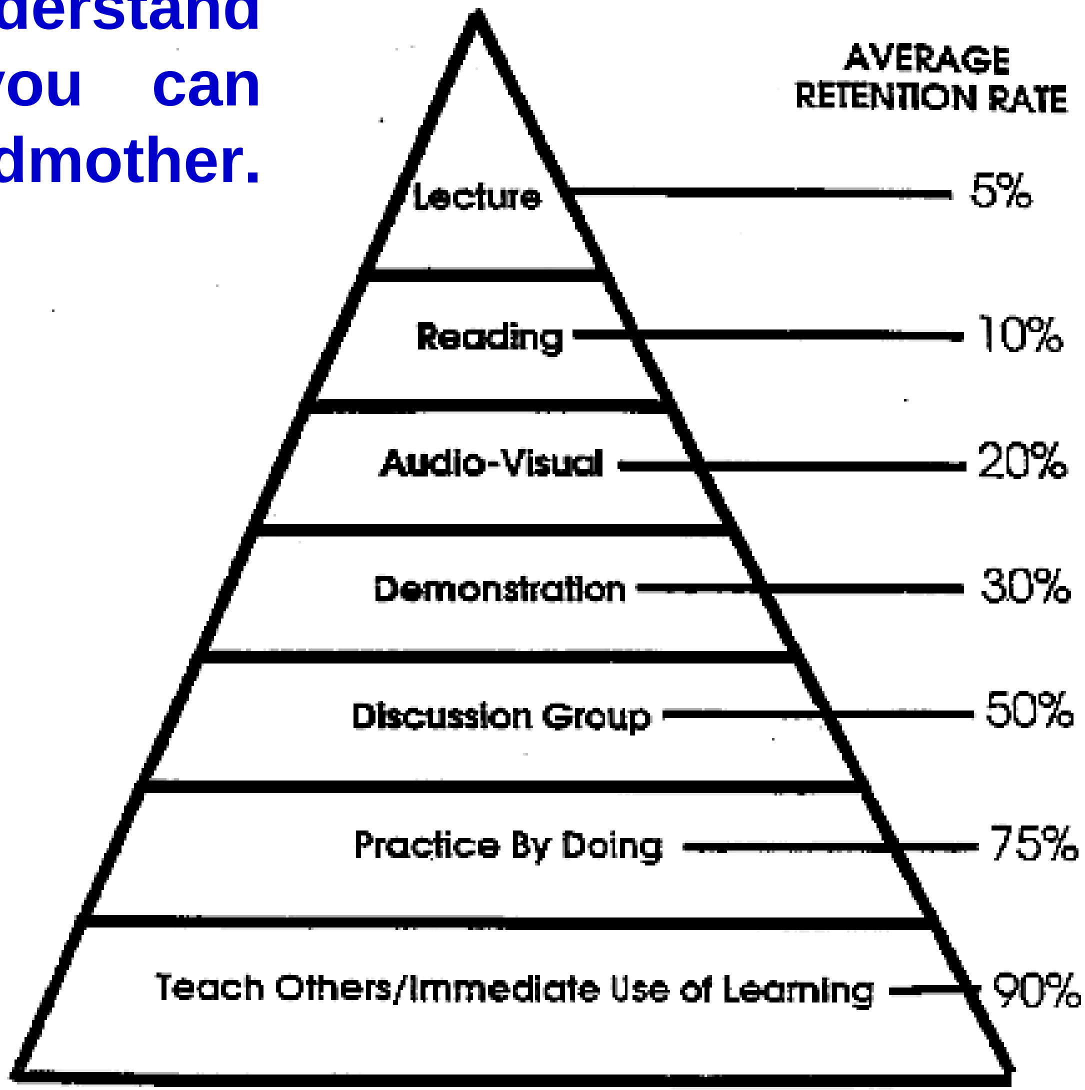
1. Attendance, active participation & home work
2. Mid term exam (10/50) 8th week
3. Final exam (20/50)
4. Laboratory work (10/50)
5. Two Quizzes (2x5 = 10 /50) 4th & 12th week

No make up exam is allowed

Keys to Success

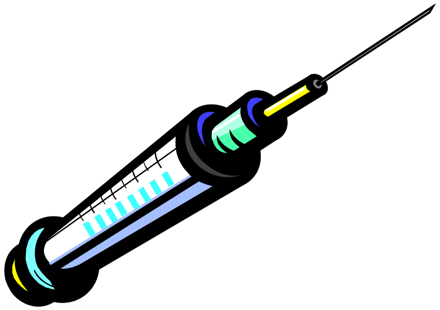
- Come to class
  - Mind as well as body
- Don't be bashful (shy)
- Actively solve problems. (keep a notebook)
- It's like a marathon – keep up a steady pace throughout
- Cooperation leads to graduation

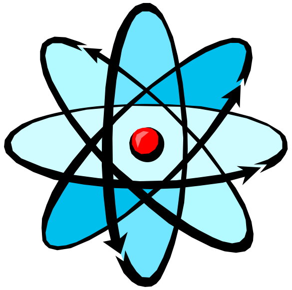
You don’t really understand something unless you can explain it to your grandmother.
A. Einstiein



Chemistry: A Science for the 21st Century

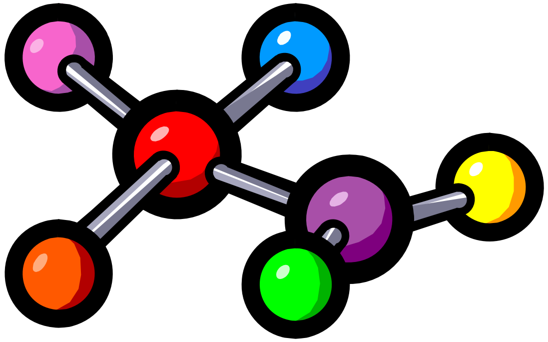
- Health and Medicine
  - Sanitation systems
  - Surgery with anesthesia
  - Vaccines and antibiotics

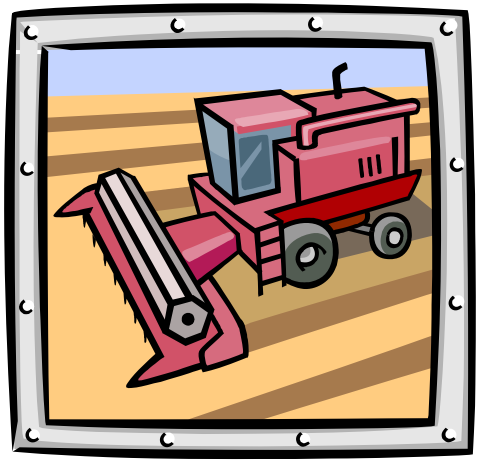




- Energy and the Environment
  - Fossil fuels
  - Solar energy
  - Nuclear energy

- Materials, Technology & Industry
    - Polymers, ceramics, liquid crystals
    - Room-temperature superconductors?
    - Molecular computing?





- Food and Agriculture
    - Genetically modified crops
    - “Natural” pesticides
    - Specialized fertilizers

The Scientific Method

- The **Scientific Method** consists of 4 parts:
  - Observation: observe, describe, and measure something to obtain data
  - Hypothesis: a guess which explains the data, and can be used to predict future experiments
  - Experiments: tests which question the validity of the hypothesis
  - Theory: an idea of how something works
  - Law: a theory that has been tested repeatedly

Chemistry is the study of matter and the changes it undergoes

1. ***Matter*** is anything that occupies space and has mass. (Weight = mass x acceleration)
2. A ***substance*** is a form of matter that has a definite composition and distinct properties.

water, ammonia, sucrose, gold, oxygen

A ***mixture*** is a combination of two or more substances in which the substances retain their distinct identities.

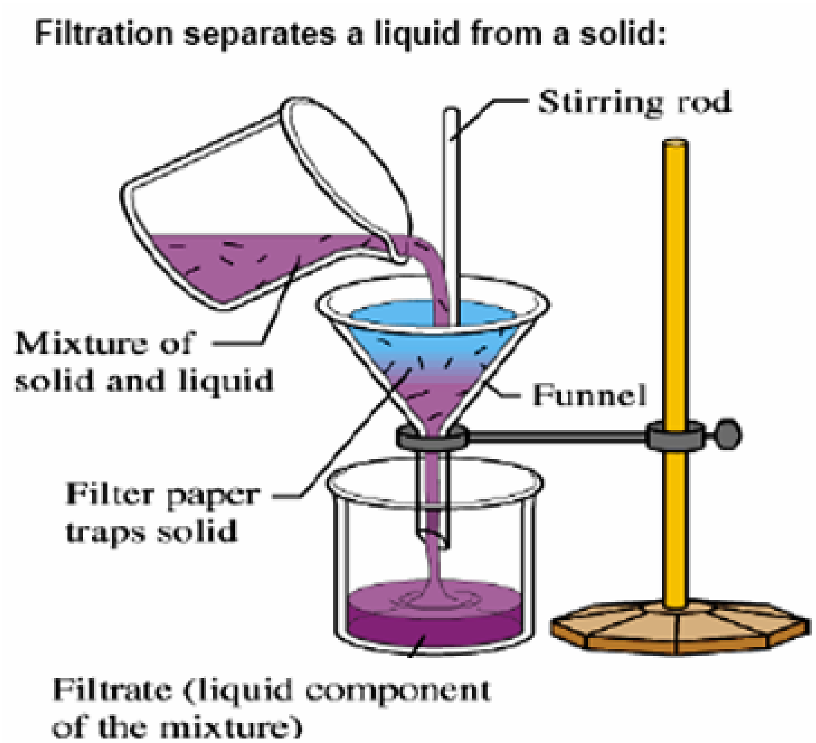
1. ***Homogenous mixture*** – composition of the mixture is the same throughout. (one phase)

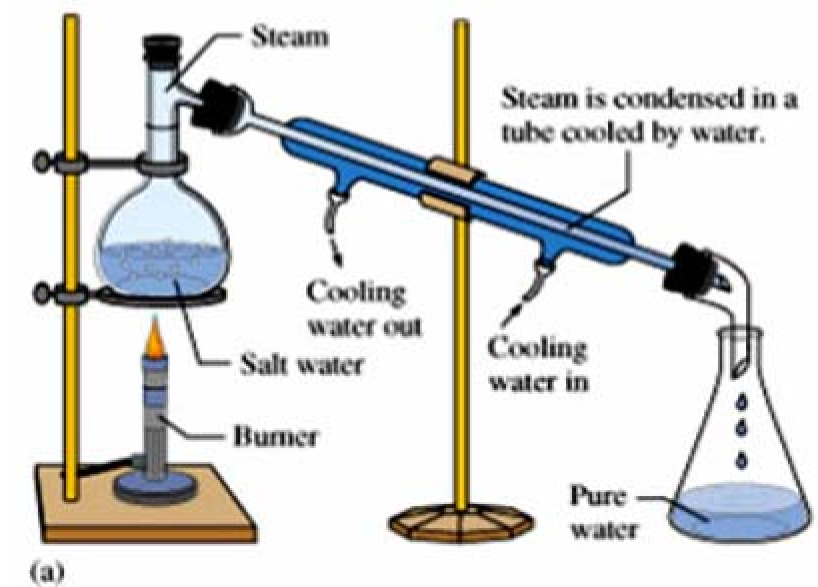
soft drink, milk, solder, fresh air, sea water, sugar in water

2. ***Heterogeneous mixture*** – composition is not uniform throughout. (Two phases, at least)

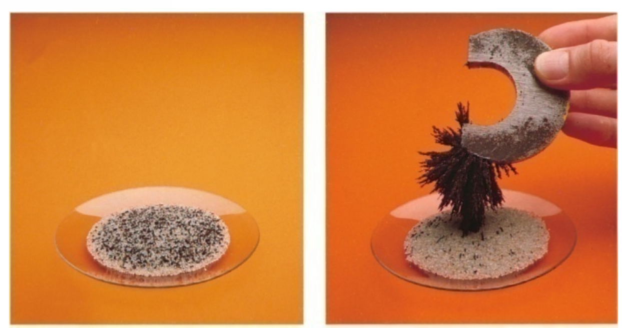
cement, coffee, smoke, iron filings in sand

A mixture can be separated into its pure components by ***Physical means*** (distillation, filtration, magnet, …).





Distillation of a solution consisting of salt dissolved in water



magnet

An ***element*** is a substance that cannot be separated into simpler substances by ***chemical means***. *There are* 114 elements known, of which 82 elements occur naturally on Earth (e.g., gold, aluminum, lead, oxygen, carbon) and 32 elements have been created by scientists (e.g., technetium, americium, seaborgium)

Some Common Elements and Their Symbols

Name	Symbol	Name	Symbol	Name	Symbol
Aluminum	Al	Fluorine	F	Oxygen	O
Arsenic	As	Gold	Au	Phosphorus	P
Barium	Ba	Hydrogen	H	Platinum	Pt
Bismuth	Bi	Iodine	I	Potassium	K
Bromine	Br	Iron	Fe	Silicon	Si
Calcium	Ca	Lead	Pb	Silver	Ag
Carbon	C	Magnesium	Mg	Sodium	Na
Chlorine	Cl	Manganese	Mn	Sulfur	S
Chromium	Cr	Mercury	Hg	Tin	Sn
Cobalt	Co	Nickel	Ni	Tungsten	W
Copper	Cu	Nitrogen	N	Zinc	Zn

These abbreviations are derived from English or Latin names

Examples: Oxygen (**O**), Sulfur (S), Aluminum (Al)

Copper: (Cu) (*Cuprum*), Iron: Fe (*Ferrum*)
Lead: Pb (*Plumbum*), Potassium: K (*Kalium*)
Silver: Ag (*Argentum*), Sodium: Na (*Natrium*)

A ***compound*** is a substance composed of atoms of two or more elements chemically united in fixed proportions. Compounds can only be separated into their pure components (elements) by ***chemical*** means.

Water (H2O), Glucose (C6H12O6), Ammonia (NH3), … etc. For example, pure water is composed of 2 parts of (H) plus one part of (O)

Classification of Matter

Matter Anything having mass and volume.			
Substance Matter with constant composition		**Mixture** Matter with variable composition	
Element Made up of only one type of atom	**Compound** Made up of two or more types of atoms chemically unified together	**Homogeneous Mixtures** (solutions) Made up of only one phase	**Heterogeneous Mixture** Made up of more than one phase
Examples - gold, silver, carbon, oxygen and hydrogen	**Examples -** water, carbon dioxide, sodium bicarbonate, carbon monoxide	**Examples -** sea water, pure air, metal alloys, seltzer water.	**Examples -** sand, soil, chicken soup, pizza, chocolate chip cookies.

Three States of Matter

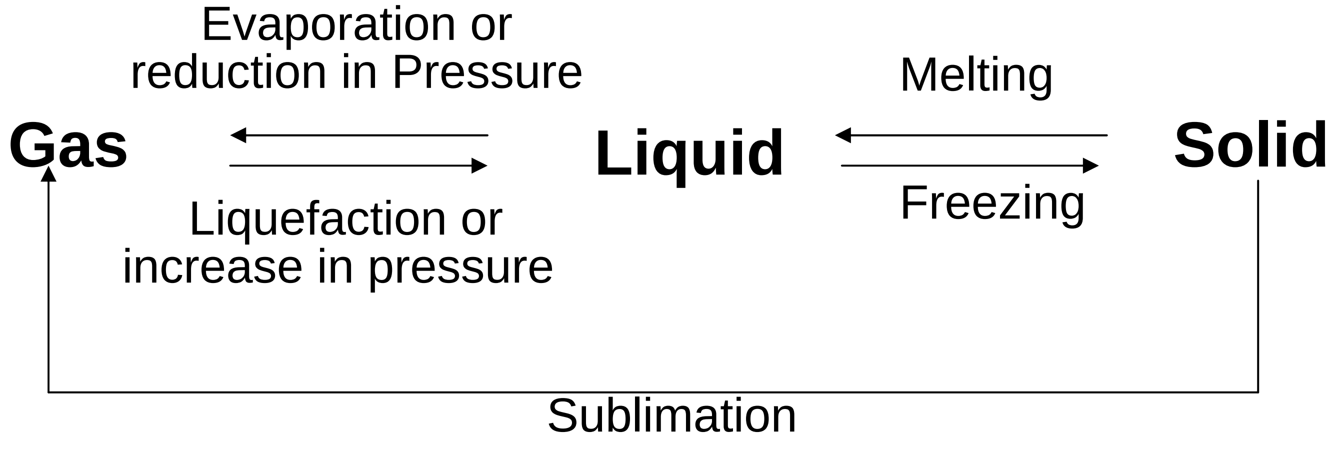
Gas Has no fixed volume or shape - it conforms to (assumes) the volume and shape of its container. Gases can be *compressed* or *expanded* to occupy different volumes.

Liquid has a distinct volume, independent of its container,

it has no specific *shape*. It assumes the shape of the container it is in. Liquids cannot be appreciably compressed.

A solid has a definite shape and volume; it is rigid.

Solids cannot be appreciably compressed



Physical or Chemical Change?

A ***physical change*** does not alter the composition or identity of a substance. e.g., ice melting, dissolution of salt in water

A ***chemical change*** alters the composition or identity of the involved substance(s); e.g., burning of H_2 in air to form water.

Extensive and Intensive Properties

An ***extensive (Quantitative) property*** of a material depends upon how much matter is being considered. e.g., length, volume, mass and weight

An ***intensive (Qualitative) property*** of a material does not depend upon how much matter is being considered. e.g., density, freezing point, melting point, colour, and conductivity

Matter - anything that occupies space and has ***mass***.

mass – measure of the quantity of matter

SI unit of mass is the ***kilogram*** (kg)

1 kg = 1000 g = 1 x 10^3 g

weight – force that gravity exerts on an object

weight = *c* x mass

on earth, *c* = 1.0

on moon, *c* ~ 0.1

A 1 kg bar will weigh

1 kg on earth

0.1 kg on moon

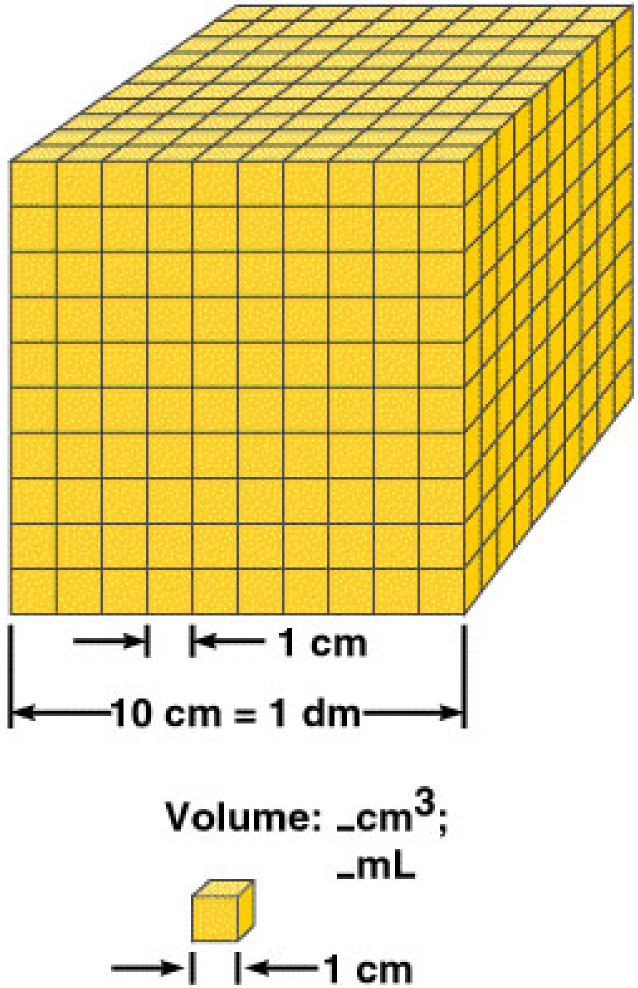
SI Base Units

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

Prefixes Used with SI Units

Prefix	Symbol	Meaning
Tera-	T	1,000,000,000,000, or 10^{12}
Giga-	G	1,000,000,000, or 10^{9}
Mega-	M	1,000,000, or 10^{6}
Kilo-	k	1,000, or 10^{3}
Deci-	d	1/10, or 10^{-1}
Centi-	c	1/100, or 10^{-2}
Milli-	m	1/1,000, or 10^{-3}
Micro-	μ	1/1,000,000, or 10^{-6}
Nano-	n	1/1,000,000,000, or 10^{-9}
Pico-	p	1/1,000,000,000,000, or 10^{-12}

Volume – SI derived unit for volume is cubic meter (m^3)



$1\ cm^3 = (1 \times 10^{-2}\ m)^3 = 1 \times 10^{-6}\ m^3$

$1\ dm^3 = (1 \times 10^{-1}\ m)^3 = 1 \times 10^{-3}\ m^3$

$1\ L = 1000\ mL = 1000\ cm^3 = 1\ dm^3$

$1\ mL = 1\ cm^3$

Volumetric flask

Density – SI derived unit for density is kg/m^3

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

$$\text{density} = \frac{\text{mass}}{\text{volume}} \qquad d = \frac{m}{V}$$

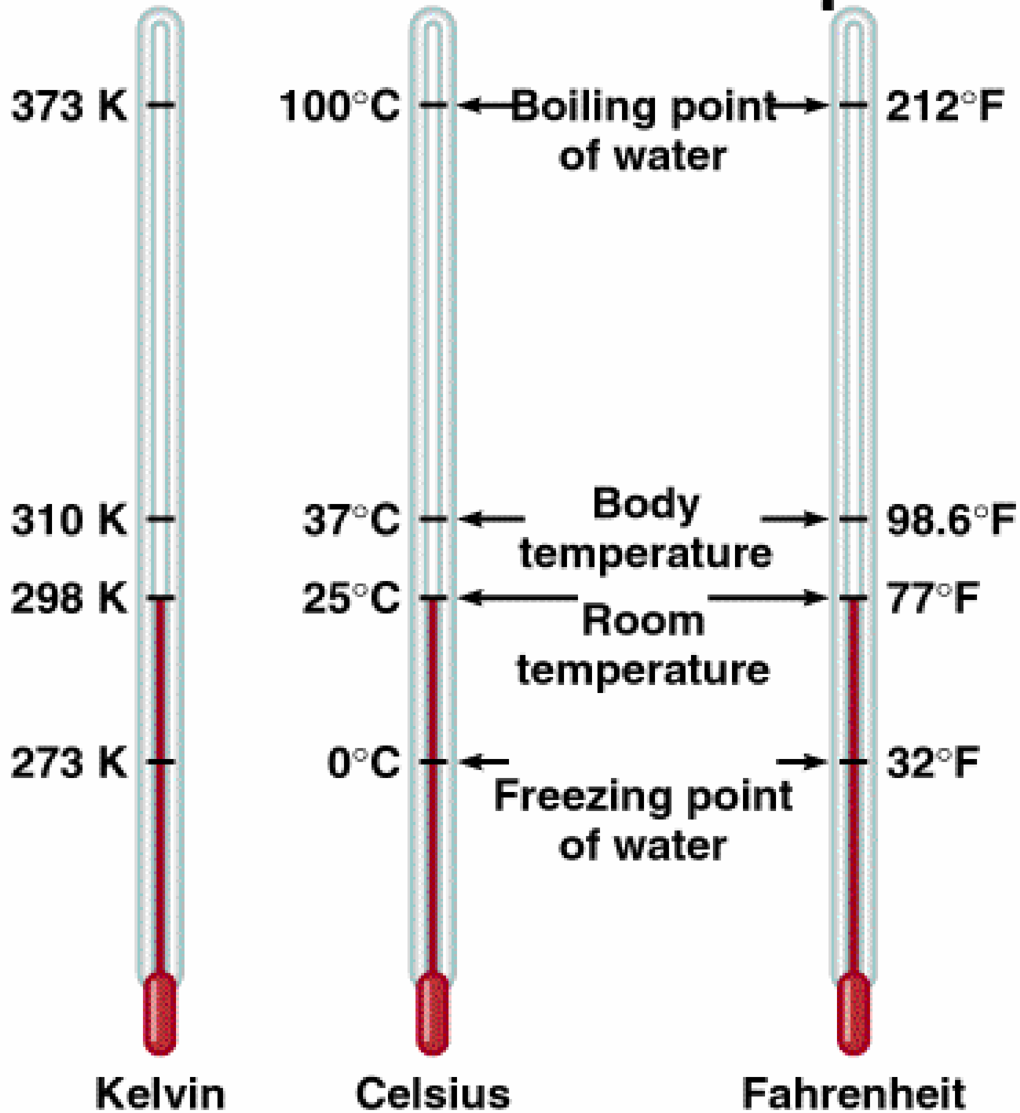
$$d = \frac{m}{V}$$

$$m = d \times V = 21.5\ g/\cancel{cm^3} \times 4.49\ \cancel{cm^3} = 96.5\ g$$

Derived Units

Base Quantity	Common Units
Volume	dm^3
Density	kg/dm^3
Acceleration	m/s^2
Force	$kg \times m/s^2$

Comparison of the Three Temperature Scales



$$K = {}^{0}C + 273.15$$

$273\ K = 0\ {}^{0}C$

$373\ K = 100\ {}^{0}C$

$${}^{0}F = \frac{9}{5} \times {}^{0}C + 32$$

$32\ {}^{0}F = 0\ {}^{0}C$

$212\ {}^{0}F = 100\ {}^{0}C$

Convert 172.9 ^{0}F to degrees Celsius.

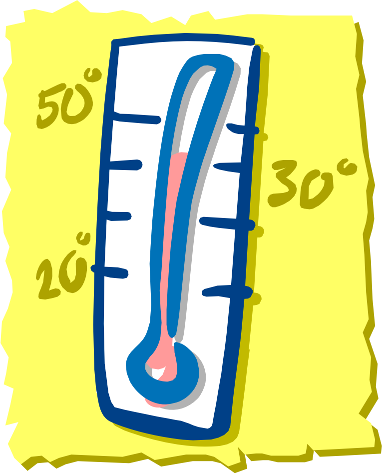
$$^0F = \frac{9}{5} \times {^0C} + 32$$

$$^0F - 32 = \frac{9}{5} \times {^0C}$$

$$\frac{5}{9} \times ({^0F} - 32) = {^0C}$$

$$^0C = \frac{5}{9} \times ({^0F} - 32)$$

$$^0C = \frac{5}{9} \times (172.9 - 32) = 78.3$$



Comparison of Temperature Scales

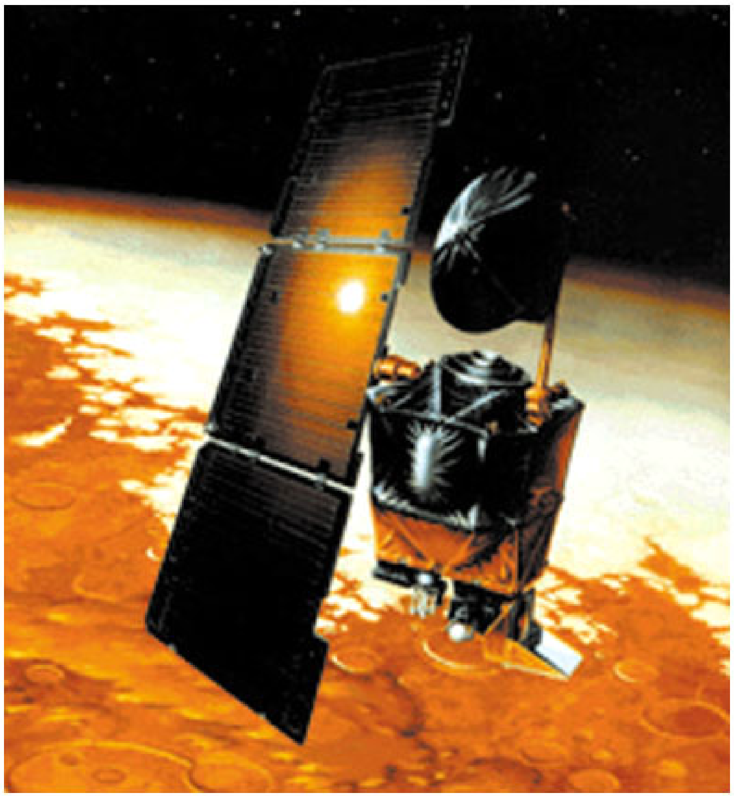
Set Points	Fahrenheit	Celsius	Kelvin
water boils	212	100	373
body temperature	98.6	37	310
water freezes	32	0	273
absolute zero	-460	-273	0

Temperature Conversion Formulas

Conversion	Formula	Example
Celsius to Kelvin	**K = C + 273**	**$21^{o}C = 294\ K$**
Kelvin to Celsius	**C = K - 273**	**$313\ K = 40\ ^{o}C$**
Fahrenheit to Celsius	**C = (F - 32) x 5/9**	**$89\ ^{o}F = 31.7\ ^{o}C$**
Celsius to Fahrenheit	**F = (C x 9/5) + 32**	**$50\ ^{o}C = 122\ ^{o}F$**

Chemistry In Action

On 9/23/99, $125,000,000 Mars Climate Orbiter entered Mar's atmosphere 100 km lower than planned and was destroyed by heat.



1 lb ~~=~~ 1 N

1 lb = 4.45 N

"This is going to be the cautionary tale that will be embedded into introduction to the metric system in elementary school, high school, and college science courses till the end of time."

Scientific Notation

The number of atoms in 12 g of carbon:

602,200,000,000,000,000,000,000

6.022×10^{23}

The mass of a single carbon atom in grams:

0.0000000000000000000000199

1.99×10^{-23}

$$N \times 10^{n}$$

N is a number between 1 and 10

n is a positive or negative integer

Scientific Notation

568.762

← move decimal left

$n > 0$

$568.762 = 5.68762 \times 10^{2}$

0.00000772

→ move decimal right

$n < 0$

$0.00000772 = 7.72 \times 10^{-6}$

Addition or Subtraction

1. Write each quantity with the same exponent n
2. Combine N_1 and N_2
3. The exponent, n, remains the same

$4.31 \times 10^{4} + 3.9 \times 10^{3} =$

$4.31 \times 10^{4} + 0.39 \times 10^{4} =$

4.70×10^{4}

Scientific Notation

Multiplication

1. Multiply N_1 and N_2
2. Add exponents n_1 and n_2

$$(4.0 \times 10^{-5}) \times (7.0 \times 10^{3}) =$$
$$(4.0 \times 7.0) \times (10^{-5+3}) =$$
$$28 \times 10^{-2} =$$
$$2.8 \times 10^{-1}$$

Division

1. Divide N_1 and N_2
2. Subtract exponents n_1 and n_2

$$8.5 \times 10^{4} \div 5.0 \times 10^{9} =$$
$$(8.5 \div 5.0) \times 10^{4-9} =$$
$$1.7 \times 10^{-5}$$

Significant Figures

• Any digit that is not zero is significant

1.234 kg   4 significant figures

• Zeros between nonzero digits are significant

606 m   3 significant figures

• Zeros to the left of the first nonzero digit are **not** significant

0.08 L   1 significant figure

• If a number is greater than 1, then all zeros to the right of the decimal point are significant

2.0 mg   2 significant figures

• If a number is less than 1, then only the zeros that are at the end and in the middle of the number are significant

0.00420 g   3 significant figures

0.00502 g   3 significant figures

How many significant figures are in each of the following measurements?

24 mL	2 significant figures
3001 g	4 significant figures
0.0320 m^3	3 significant figures
6.4×10^4 molecules	2 significant figures
560 kg	2 significant figures

Significant Figures

Addition or Subtraction

The answer cannot have more digits to the right of the decimal point than any of the original numbers.

89.332		
+1.1	←	one significant figure after decimal point
90.432	←	round off to 90.4

3.70	←	two significant figures after decimal point
-2.9133		
0.7867	←	round off to 0.79

Significant Figures

Multiplication or Division

The number of significant figures in the result is set by the original number that has the ***smallest*** number of significant figures

$$4.51 \times 3.6666 = 16.536366 = 16.5$$

- 4.51 → 3 sig figs
- 16.536366 → round to 3 sig figs

$$6.8 \div 112.04 = 0.0606926 = 0.061$$

- 6.8 → 2 sig figs
- 0.0606926 → round to 2 sig figs

Significant Figures

Exact Numbers

Numbers from definitions or numbers of objects are considered to have an infinite number of significant figures

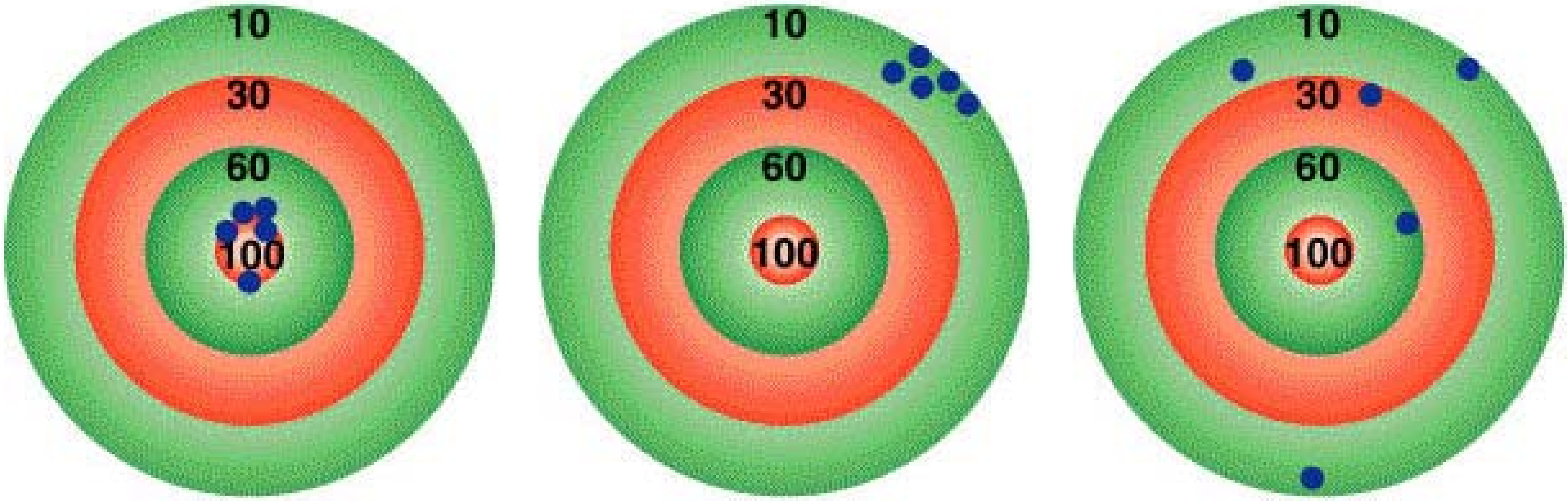
The average of three measured lengths; 6.64, 6.68 and 6.70?

$$\frac{6.64 + 6.68 + 6.70}{3} = 6.67333 = 6.67 = \cancel{7}$$

Because 3 is an exact number

Accuracy – how close a measurement is to the *true* value (closeness to the true value)

Precision – how close a set of measurements are to each other (concordance between repeated measurements)



accurate & precise	precise but not accurate	not accurate & not precise

Dimensional Analysis Method of Solving Problems

1. Determine which unit conversion factor(s) are needed
2. Carry units through calculation
3. If all units cancel except for the desired unit(s), then the problem was solved correctly.

How many mL are in 1.63 L?

1 L = 1000 mL

$$1.63 \cancel{L} \times \frac{1000 \text{ mL}}{1 \cancel{L}} = 1630 \text{ mL}$$

$$1.63 \text{ L} \times \frac{1 \text{L}}{1000 \text{ mL}} = 0.001630 \frac{\text{L}^2}{\text{mL}}$$

The speed of sound in air is about 343 m/s. What is this speed in miles per hour?

meters to miles

seconds to hours

1 mi = 1609 m   1 min = 60 s   1 hour = 60 min

$$343 \frac{\cancel{m}}{\cancel{s}} \times \frac{1\ mi}{1609\ \cancel{m}} \times \frac{60\ \cancel{s}}{1\ \cancel{min}} \times \frac{60\ \cancel{min}}{1\ hour} = 767 \frac{mi}{hour}$$

Homework

- Solve Problems 1.6, 1.7 & 1.8
- It’s highly recommended to solve the odd or even numbered questions after each chapter!
- Just reading the questions would make students familiar with various question styles.