Cairo University

Basic General Chemistry II

Chem 102 – Spring 2019

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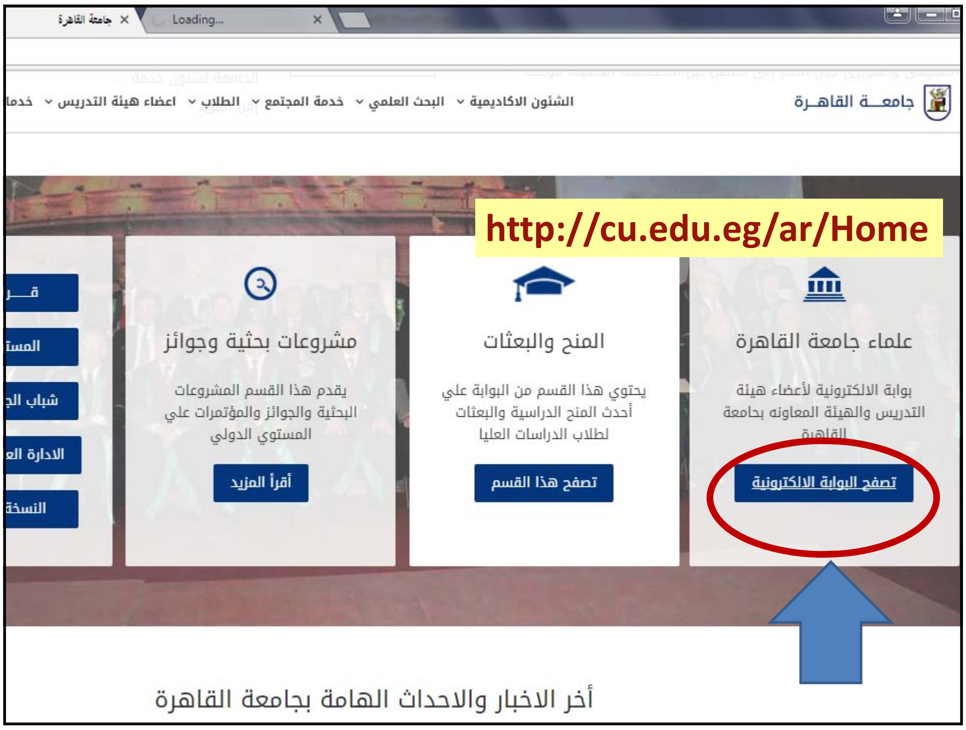

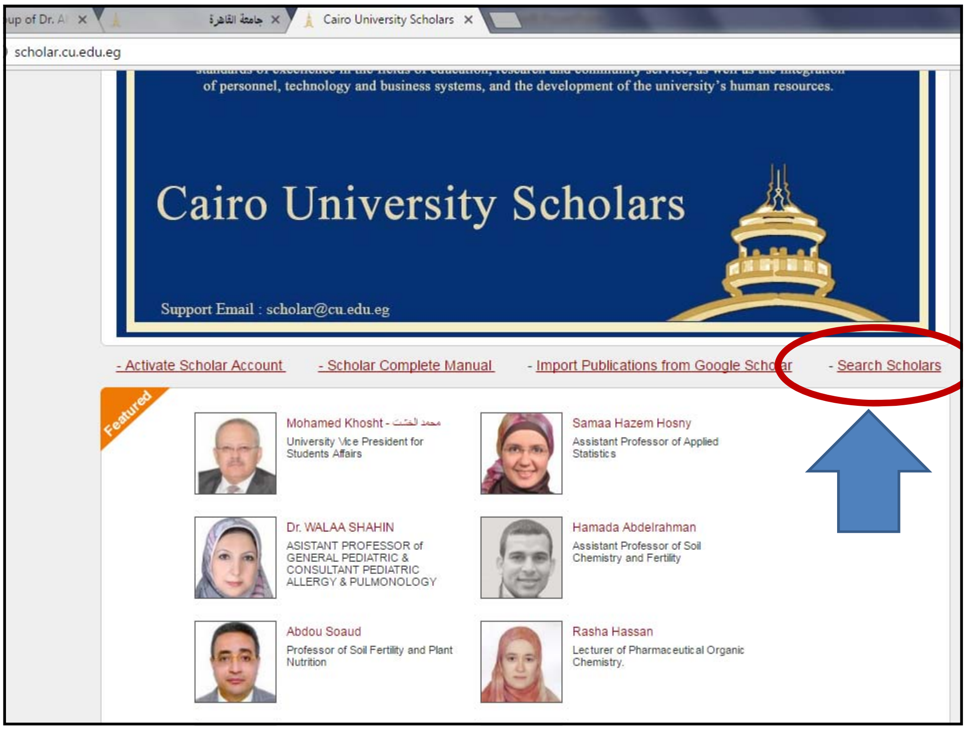
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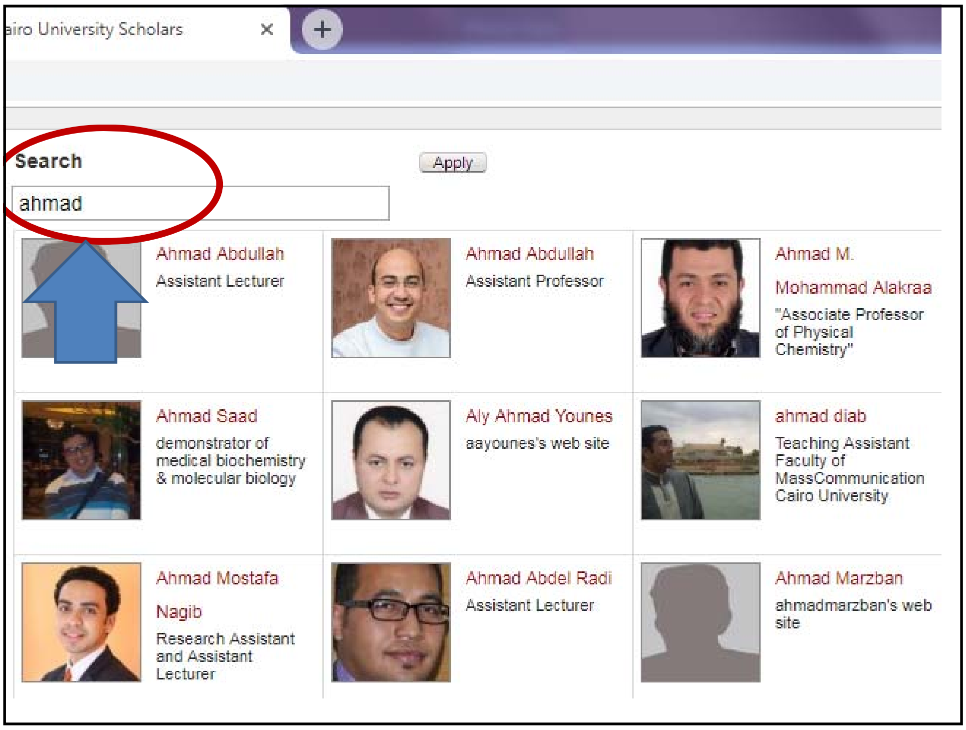
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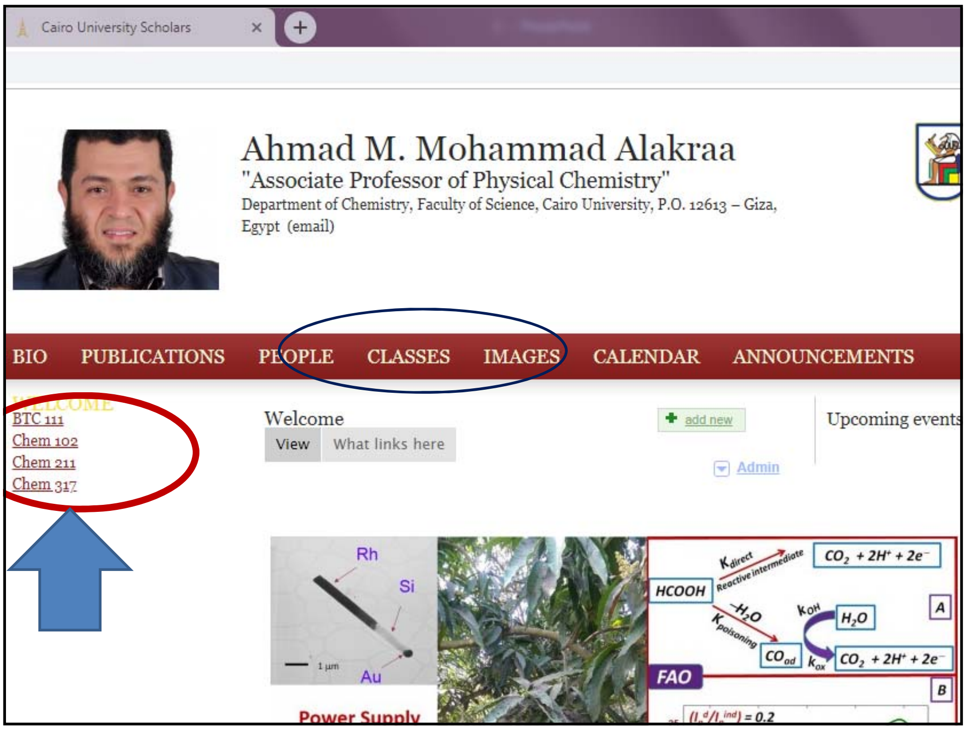
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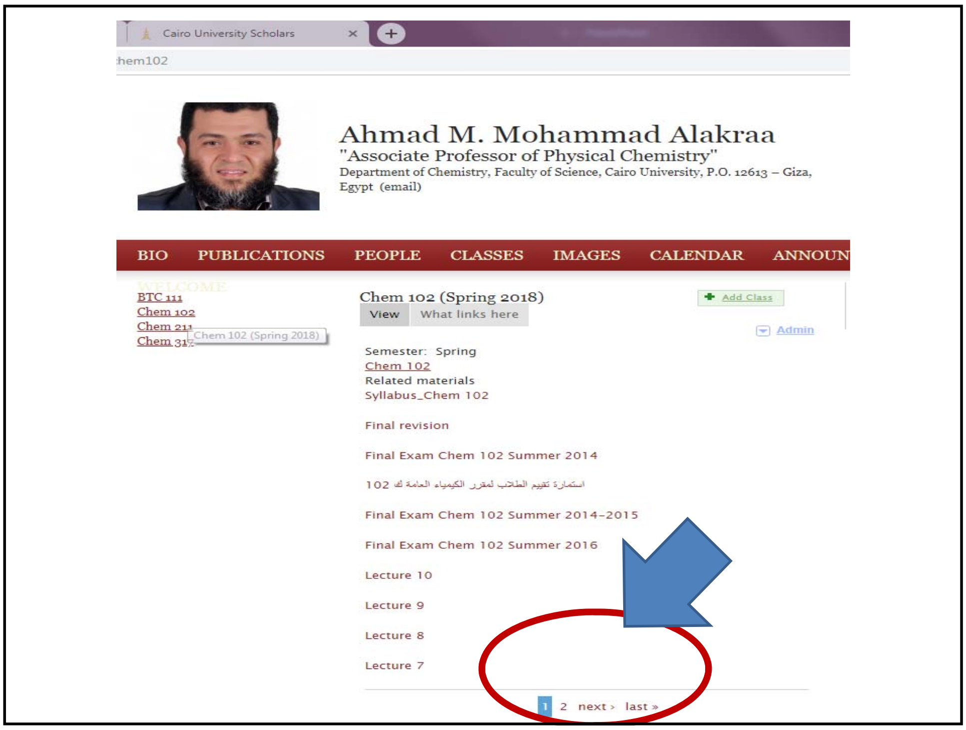
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Credit

Level One : First Semester						
Code	Subject	Pre-requisite	Practical (hr/wk)	Lecture (hr/wk)	Total cr	Contact hr
Chem102	Basic General Chemistry II		1 (3 × 1/3)	2 (2 × 1)	3	5

References

1- Chemistry: An Atoms First Approach, Steven S. Zumdahl and Susan A. Zumdahl, 2012, Brooks Cole, a part of Cengage Learning.

2- Chemistry: The Central Science, Theodore L. Brown et al., 2012, Pearson Prentice Hall, USA

3- BTC 111 Note from the Department of Chemistry

Assessment

- A 2h unseen written examination carries 60% of the total mark
- Mid term Exam. carries 10% of the total mark.
- Lab work carries 30 % of the total mark.

Passing Criteria

✓ 60% for the total course mark

Lectures attendance should exceed 70% in order to attend the final unseen examination

Chapter 1: Units and Dimensional Analysis

1.1 Units
- *The SI Units.*
- *Fundamental and derived units.*

1.2 Precision and Accuracy.
- *Systematic and random errors.*

1.3 Uncertainty in Measurements and Significant Figures

1.4 Rules for Counting Significant Figures
- *For multiplication and division.*
- *Rounding.*
- *For addition and subtraction.*
- *Scientific notation.*

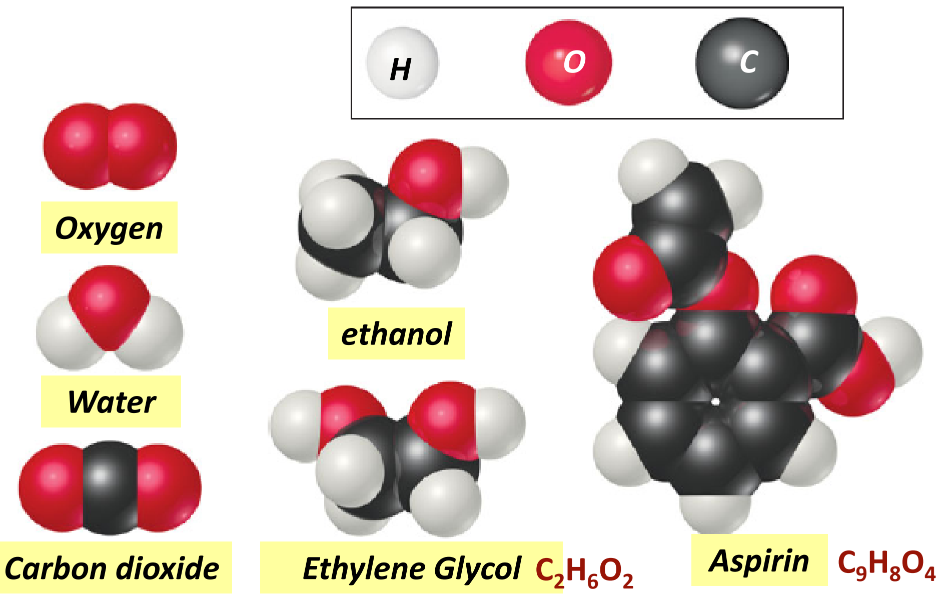
1.5 Dimensional Analysis

1.6 Temperature Scales (Celsius, Kelvin and Fahrenheit)

Matter and Chemistry

- *Matter is the physical material of the universe; it is anything that has a mass and occupies space.*
- *Chemistry deals with studying properties of matter and changes (behavior) that matter undergoes.*
- *A property is a characteristic recognizing a particular type of matter to distinguish it from other types.*
- *Properties of matter relate to both the kinds of atoms the matter contains (composition) and to the arrangements of these atoms (structure).*
- *The tremendous variety of matter in our world is due to combinations of only about 118 substances called elements.*

Properties vs. Structure and Composition



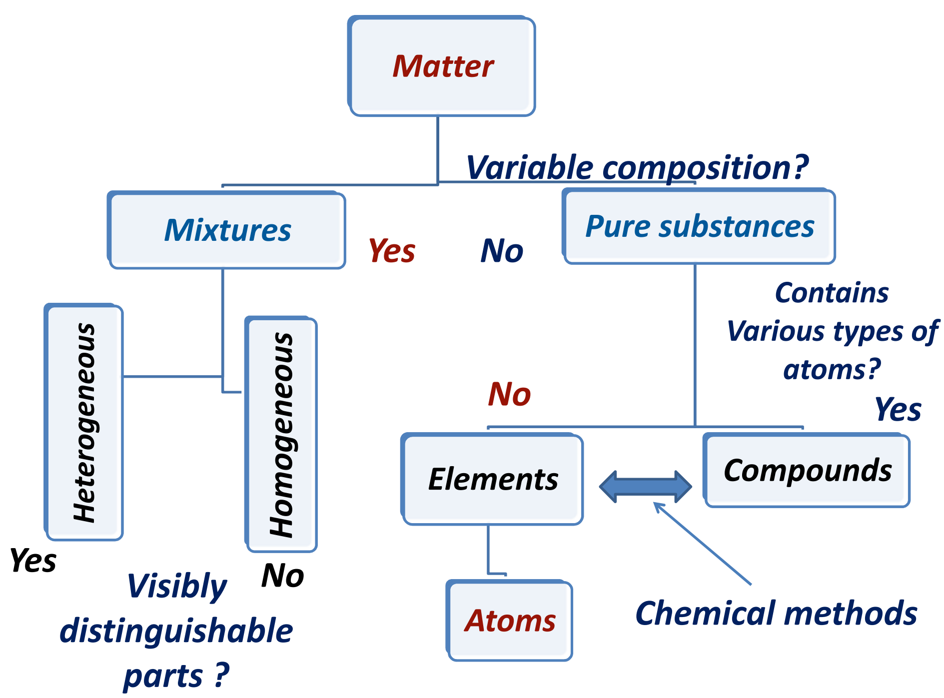


TABLE R.5 › Densities of Various Common Substances* at 20°C

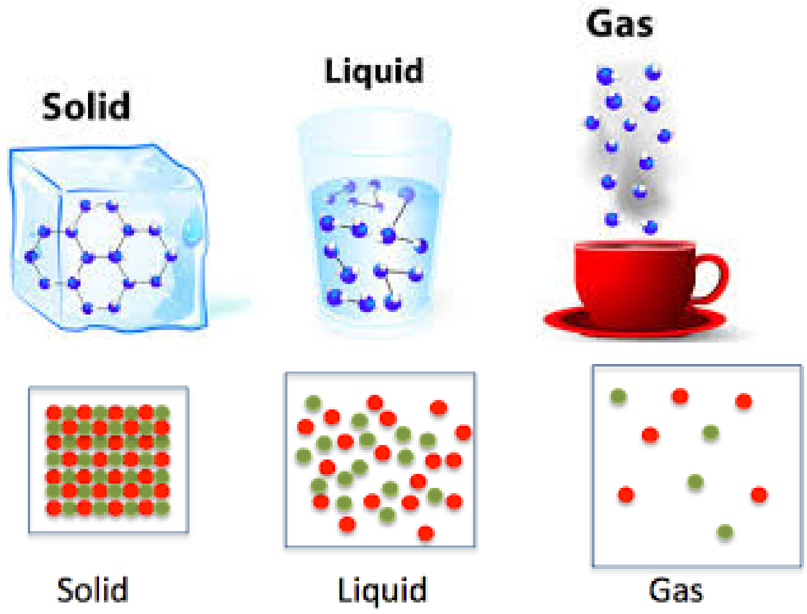
Substance	Physical State	Density (g/cm^3)
Oxygen	Gas	0.00133
Hydrogen	Gas	0.000084
Ethanol	Liquid	0.789
Benzene	Liquid	0.880
Water	Liquid	0.9982
Magnesium	Solid	1.74
Salt (sodium chloride)	Solid	2.16
Aluminum	Solid	2.70
Iron	Solid	7.87
Copper	Solid	8.96
Silver	Solid	10.5
Lead	Solid	11.34
Mercury	Liquid	13.6
Gold	Solid	19.32

*At 1 atmosphere pressure.

Properties "Characteristics" of Matter

Character	Solid	Liquid	Gas
Particle packing "arrangement"	Regular	Irregular	Irregular
Shape	Fixed	Not fixed	Not fixed
Volume	Fixed	Fixed	Not fixed
Motion	Only vibrating	Move around past each other	Freely - randomly
Compressibility	No	little	high

Characteristics of Matter



Elements that exist as *gases* at 25⁰C and 1 atmosphere

1A	2A	3B	4B	5B	6B	7B		8B		1B	2B	3A	4A	5A	6A	7A	8A
H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt									

Some Substances Found as Gases at 1 atm and 25°C

Elements	Compounds
H_2 (molecular hydrogen)	HF (hydrogen fluoride)
N_2 (molecular nitrogen)	HCl (hydrogen chloride)
O_2 (molecular oxygen)	HBr (hydrogen bromide)
O_3 (ozone)	HI (hydrogen iodide)
F_2 (molecular fluorine)	CO (carbon monoxide)
Cl_2 (molecular chlorine)	CO_2 (carbon dioxide)
He (helium)	NH_3 (ammonia)
Ne (neon)	NO (nitric oxide)
Ar (argon)	NO_2 (nitrogen dioxide)
Kr (krypton)	N_2O (nitrous oxide)
Xe (xenon)	SO_2 (sulfur dioxide)
Rn (radon)	H_2S (hydrogen sulfide)
	HCN (hydrogen cyanide)*

* The boiling point of HCN is 26°C, but it is close enough to qualify as a gas at ordinary atmospheric conditions.

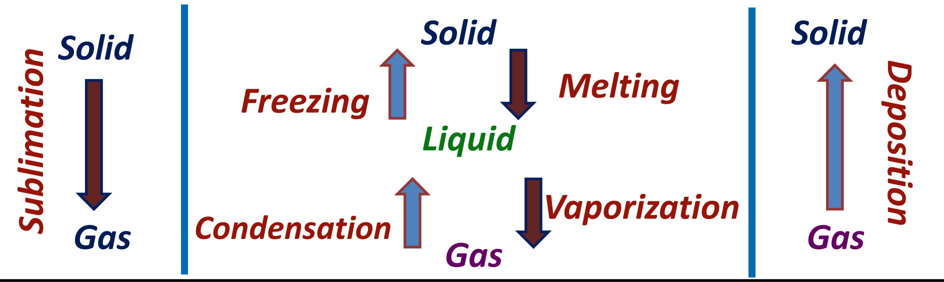
Gas	*Liquid*	*Solid*
Particles are far apart, run in rapid random motion (translation, rotational, vibrational)	*Particles lies in-between, intermediate motion (translation, rotational)*	*Particles are very close together, vibrate only in place*
High volumes and Low densities	*Intermediate volumes and densities*	*Small volumes and high densities*
Very weak attraction forces	*Intermediate forces*	*Strong forces*
assumes the shape and volume of its container	*assumes the shape of the part of the container which it occupies – has a fixed volume*	*retains a fixed volume and shape* *rigid - particles locked into place*
compressible *lots of free space between particles*	*not easily compressible* *little free space between particles*	*not easily compressible* *little free space between particles*
flows easily *particles can move past one another*	*flows easily* *particles can move/slide past one another*	*does not flow easily* *rigid - particles cannot move/slide past one another*

Liquids and Solids: condensed phases

Liquids and Gases: Fluids

Conversion of States

- ***Sublimation*** ***is the process of changing from the solid phase right to a gas phase, without passing by the liquid state.***
- ***Deposition*** ***is a process in which a gas will form a solid, again without any apparent liquid phase in between.***



Boiling Point

- ***The temperature at which a liquid boils and at which the vapor pressure of the liquid equal the external atmospheric pressure.***

Normal Boiling Point

- ***The temperature at which a liquid boils and at which the vapor pressure of the liquid equal 1 atm.***

Chemists are concerned principally with

- ***Evaluating* the properties of materials.**
- ***Inspecting* the feasibility of occurring reactions.**
- ***Estimating* the kinetics (how fast) of reactions.**
- ***Recommending* a convenient pathway for a given reaction (catalysis, inhibition).**

Chemical Reactions/Balancing

- ***Reorganization of atoms in one or more substances.***
- ***Mass is conserved in a chemical reaction.***

$$CH_4 + O_2 \rightarrow CO_2 + H_2O$$

Unbalanced

- ***Atoms are neither created nor destroyed***

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

Balanced

Meaning of a chemical equation

- The chemical equation for a reaction gives important information:
  - ✓ the nature of the reactants and products
  - ✓ the relative numbers of each.
  - ✓ the physical states of the reactants and products:

State	Symbol
Solid	(s)
Liquid	(l)
Gas	(g)
Dissolved in Water (aqueous)	(aq)

$$HCl\ (aq) + NaHCO_3\ (s) \rightarrow CO_2\ (g) + H_2O\ (l) + NaCl\ (aq)$$

Reading Chemical Equations

	Reactants				Products		
	$CH_4(g)$	+	$2O_2(g)$	$\rightarrow$	$CO_2(g)$	+	$2H_2O(g)$
Molecules	1		2		1		2
Moles	1		2		1		2
Molecules /mole	6.022×10^{23}		$2\ (6.022 \times 10^{23})$		6.022×10^{23}		$2\ (6.022 \times 10^{23})$
g	16		2 (32)		44		2 (18)
Total mass	80				80		

Mass is conserved in a chemical reaction

Stoichiometric Calculations

Reaction of propane with oxygen: (Combustion)

"What mass of oxygen will react with 96.1 g of propane?"

- *Balance equation*

$$C_3H_8\ (g) + 5O_2\ (g) \rightarrow 3CO_2\ (g) + 4H_2O\ (g)$$

- *Convert masses to moles:*

$$96.1\ \cancel{g\ C_3H_8} \times \frac{1\ mol\ C_3H_8}{44.1\ \cancel{g\ C_3H_8}} = 2.18\ mol\ C_3H_8$$

- *Number of moles of O_2 necessary to react with 2.18 mole C_3H_8*

$$2.18\ mol\ C_3H_8 \times \frac{5\ mol\ O_2}{1\ mol\ C_3H_8} = 10.9\ mol\ O_2$$

- *Convert from moles to grams O_2*

$$10.9\ mol\ O_2 \times \frac{32\ g\ O_2}{1\ mol\ O_2} = 349\ g\ O_2$$

- *Therefore, 349 g O_2 is required to burn 96.1 g propane.*

- *What mass of carbon dioxide is produced when 96.1 grams of propane is combusted with oxygen? (Homework)*

Units of measurements

- Many properties of matter are quantitative, that is, associated with numbers.
- A pencil's length of 17.5 is meaningless. 17.5 cm is acceptable.
- The units used for scientific measurements are those of the metric system which was developed in France during the late eighteenth century.

- The United States has traditionally used the English system, although use of the metric system has become more common

English units (fluid ounces, FL OZ) and metric units (milliliters, mL).

- In 1960 an international agreement was reached specifying a particular choice of metric units for use in scientific measurements.
- These preferred units are called SI units, after the French Système International d'Unités.

7 SI Base UNITS *from which all other units are derived*

Dimension	Unit	Unit Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Temperature	kelvin	K
Electrical Current	ampere	A
Amount of light	candela	cd
Amount of matter	mole	mol

Recognize the capital and small letters

Standard prefixes in SI Base units

Multiple	Prefix
10^{24}	yotta, Y
10^{21}	zetta, Z
10^{18}	exa, E
10^{15}	peta, P
10^{12}	tera, T
10^{9}	giga, G
10^{6}	mega, M
10^{3}	kilo, k
10^{2}	hecto, h
10^{1}	deka, da
10^{-1}	deci, d
10^{-2}	centi, c
10^{-3}	milli, m
10^{-6}	micro, μ
10^{-9}	nano, n
10^{-12}	pico, p
10^{-15}	femto, f
10^{-18}	atto, a
10^{-21}	zepto, z
10^{-24}	yocto, y

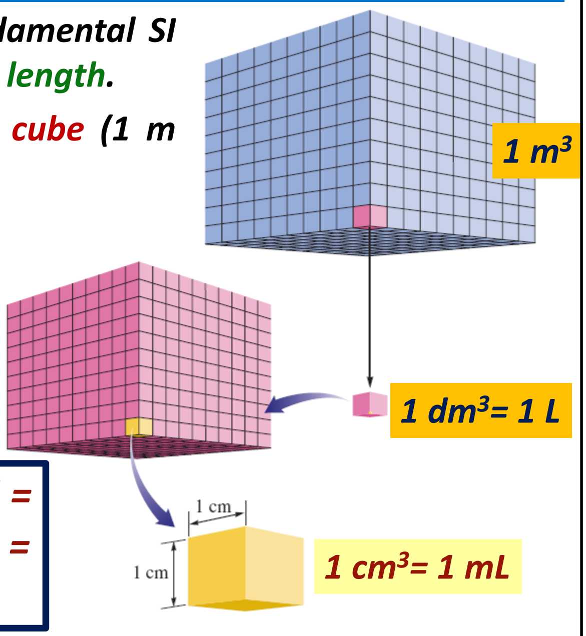
Exercises

- Which is the smallest: 1 mg, 1 μg, or 1 pg? 1 pg
- What is the name of the unit that equals to (a) 10^{-9} g , (b) 10^{-6} s, (c) 10^{-3} m? ng, μs, mm.
- How many picometers are there in 1 m? 10^{12} pm
- Express 6.0×10^3 m using a prefix to replace the power of ten? 6.0 km.
- Use exponential notation to express 4.22 mg in grams? 4.22×10^{-3} g.
- Use decimal notation to express 4.22 mg in grams? 0.00422 g

Derived SI Units (Volume)

- Volume is not a fundamental SI unit but derived from length.
- The volume (V) of a cube (1 m edge) = $(1\ m)^3$ =1 m^3.
- As 1 m = 10 dm, V= $(1\ m)^3$ = $(10\ dm)^3$ = 1000 dm^3.
- 1 dm^3 is commonly called a liter (L)

1L = $(1\ dm)^3$ = 1 dm^3 = $(10\ cm)^3$ = 1000 cm^3 = 1000 mL = 1L



Other Derived SI Units

Physical quantity	Symbol (s)	Name of SI unit	Derived Unit	Definition
Frequency	v, f	Hertz	Hz	s^{-1}
Force	F	Newton	N	$kg\ m\ s^{-2} = J\ m^{-1}$
Energy	E, H,V, etc	Joule	J	$N\ m = kg\ m^2\ s^{-2}$
Pressure	P	Pascal	Pa	$N\ m^{-2} = kg\ m^{-1} s^{-2}$
Power	p	Watt	W	$J\ s^{-1} = kg\ m^2\ s^{-3}$
Charge	Q	Coulomb	C	A s
Potential	E,…etc	Volt	V	$JA\ s^{-1}$
Resistance	R	Ohm	Ω	$V\ A^{-1}$
Conductance	G	Siemens	S	Ω^{-1}
Capacitance	C	Farad	F	$C\ V^{-1}$

Other Units

Physical quantity	Symbol	SI unit
Area	A	m^2
Volume	V	m^3
Velocity	U, V, c	$m\ s^{-1}$
Acceleration	a, g	$m\ s^{-2}$
Weight	G,W	N
Density	p	$kg\ m^{-3}$
Volume	liter (l)	dm^3
Force	dyne (dyn)	10^{-5} N
Concentration	Molar (M)	$mol\ dm^{-3}$
Energy	Calorie (Cal)	4.18 J
Energy	Erg (erg)	10^{-7} J
Pressure	Atmosphere (atm)	1.013×10^5 Pa
Pressure	(mm Hg)	133.322 Pa
Pressure	Torr (torr)	133.322 Pa
Pressure	Bar	10^5 Pa
Pressure	Atmosphere	760 mm Hg = 76 cm Hg

Errors

- ***Systematic "determinate" error:***
  - *Error existing in every reading of a series of repeated measurements. (like a speck of dust on a pan).*
  - *Error occurring in the same direction each time; It is either always high or always low*
- ***Random "indeterminate" errors:***
  - *errors varying randomly in a series of repeated measurements and can average to zero over a series of observations.*
  - *It means that a measurement has an equal probability of being high or low. This type of error occurs in estimating the value of the last digit of a measurement*

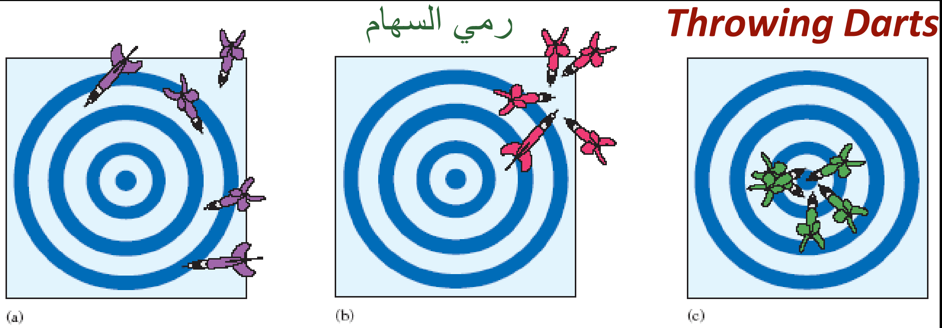
Precision and Accuracy

- ***Accuracy** refers to the agreement of a particular value with the true value.*
- ***Precision** refers to the degree of agreement among several measurements of the same quantity.*

***Precision** reflects the reproducibility of a given type of measurement.*

رمي السهام

Throwing Darts



Neither accurate nor precise (large random errors).

Precise but not accurate (small random errors, large systematic error)

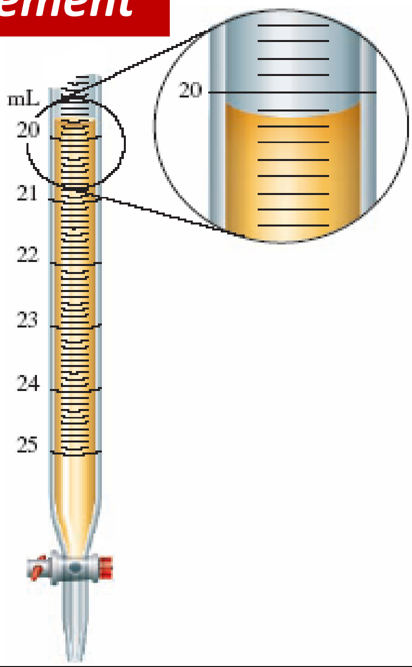
Both precise and accurate (small random errors, no systematic error)

Uncertainty in Measurement

- ***If five different people read the same volume, the results in mL might be***

  20.16, 20.14, 20.16, 20.17, 20.15

- ***20.1 are certain digits but the last digit is estimated (uncertain) digit***



Significant Figures, SF

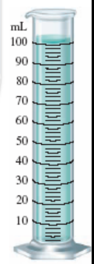
- Any measurement is reported by recording the significant figures, which represent all the certain digits plus the first uncertain digit.

- When using the burette, it would not make any sense to try to record the volume of thousandths of a milliliter (0.004) (insignificant) because the value for hundredths of a milliliter must be estimated.

- In analyzing a sample of polluted water, a chemist measured out a 25.00-mL water sample with a pipette. Another chemist used a graduated cylinder to measure 25 mL of a solution. What is the difference between the measurements?

Solution

- The quantity 25 mL means that the volume is between 24 mL and 26 mL, whereas the quantity 25.00 mL means that the volume is between 24.99 mL and 25.01 mL.



A pipette measures volume with much greater precision

Rules for counting SF

- ***Nonzero integers*. always count as SF**
- ***Zeros*: three classes**

A- Leading zeros (zeros precede nonzero digits) do not count as SF

0.0025 → has only two SF

B- Captive zeros (zeros between nonzero digits) always count as SF

1.008 → has four SF

C- Trailing zeros (zeros at the right end of the number) are significant only if the number contains a decimal point.

100 → has only one SF

1.00×10^2 → has three SF

- ***Exact numbers*. numbers that were not obtained using measuring devices but were determined by counting.**

Examples:

- ***10 experiments, 3 apples, 8 molecules. Such numbers are assumed to have an infinite number of SF.***
- ***The 2 in $2\pi r$ (circumference of a circle) and the 4/3 in the 4/3 r^3 (volume of a sphere).***
- ***Exact numbers also can arise from definitions. For example, one inch is defined as exactly 2.54 centimeters. Thus, in the statement 1 in = 2.54 cm, neither the 2.54 nor the 1 limits the number of SF when used in a calculation.***

Exercise

- ***Count the number of SF in the following:***
  - ***A student's extraction procedure on tea yields 0.0105 g of caffeine. 3 SF. The zeros to the left of the 1 are leading zeros (not SF), but remaining zero (a captive zero) is SF.***
  - ***A chemist records a mass of 0.050080 g in an analysis. 5 SF. (2 captive zeros between 5 and 8 + 1 trailing zero because of the decimal point all are SF plus 5 and 8).***
  - ***In an experiment a span of time is determined to be 8.050 x 10^{-3} s. 4 SF.***

Rules of multiplication and division

- The number of significant figures in the result is the same as the number in the least precise measurement used in the calculation.

$$4.56 \times 1.4 = 6.38 \xrightarrow{\text{Corrected}} 6.4$$

Limiting term has 2 SF

2 SF

Rules of addition and subtraction

- For addition or subtraction, the result has the same number of decimal places as the least precise measurement used in the calculation.

$$\begin{array}{l} 12.11 + \\ 18.0 + \quad \text{Limiting term has 1 DP} \\ \;\;1.013 \\ \hline 31.123 \xrightarrow{\text{Corrected}} 31.1 \end{array}$$

Rules for rounding

- *In a series of calculations, carry the extra digits through to the final result, then round.*
- *If the digit to be removed*
  - *is less than 5, the preceding digit stays the same. For example, 1.33 rounds to 1.3.*
  - *is equal to or greater than 5, the preceding digit is increased by 1. For example, 1.36 rounds to 1.4.*
- *When rounding, use only the first number to the right of the last significant figure.*
  - *Do not round sequentially. The number 6.8347 rounded to three significant figures is 6.83, not 6.84.*

Calculate and count SF

$$1.05\times10^{-3} \div 6.135 = 1.711491\times10^{-4}$$

$$\xrightarrow{\text{Corrected to 3 SF}} 1.71\times10^{-4}$$

$$21 - 13.8 = 7 \quad \textit{no DP } 1\text{ SF}$$

Determine the value of the gas constant (R) for a gas with a pressure (P = 2.560 atm), molar volume (V =8.8 L/mol), and temperature (T = 275.15 K)?

$$R = \frac{PV}{T} = \frac{(2.560)(8.8)}{275.15} = \frac{22.528}{275.15} =$$

$$0.081875321 = 0.082 = 8.2\times10^{-2} \quad 2\text{ SF}$$

Dimensional Analysis (unit factor method)

- It is a method used to convert a given result from one system of units to another.

  Apples and oranges do not add

- Calculate in inches the length of a 2.85 cm pen knowing that 2.54 cm = 1 in?

$$1 = \frac{1\,in}{2.54\,cm}$$

is called a unit factor which does not affect SF

$$2.85\,cm \times \frac{1\,in}{2.54\,cm} = \frac{2.85}{2.54}\,in = 1.12\,in \quad 3\,SF$$

Temperature

- Temperature, is a physical property measuring the hotness or coldness of an object and determining the direction of heat flow.
- Heat always flows spontaneously from a substance of a higher temperature to another of a lower temperature. We feel the influx of heat when we touch a hot object.
- The temperature scales commonly employed in science are the Celsius and Kelvin scales.
- The Celsius scale was originally based on the assignment of 0 ℃ to the freezing point of water and 100 ℃ to its boiling point at sea level
- The SI unit of temperature is the kelvin (K).

Temperature

- Zero on the Kelvin scale is the lowest attainable temperature, –273.15 ℃, referred to as absolute zero
- The Celsius and Kelvin scales have equal-sized units.

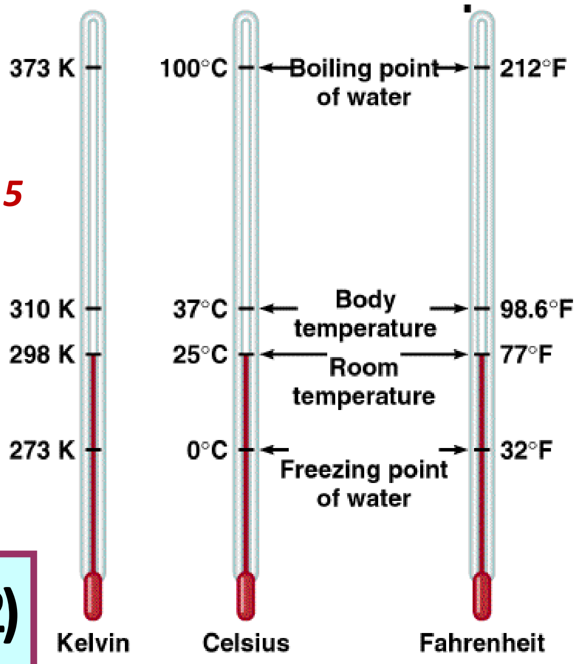
$$K = {}^{\circ}C + 273.15$$

- The common temperature scale in the United States is the Fahrenheit scale.
- Water freezes at 32 ℉ and boils at 212 ℉.

Scales comparison

$${}^{O}F = \frac{9}{5}({}^{O}C) + 32$$

$$K = {}^{\circ}C + 273.15$$



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

Exercises

Normal body temperature is 98.6°F. Convert this temperature to the Celsius and Kelvin scales?

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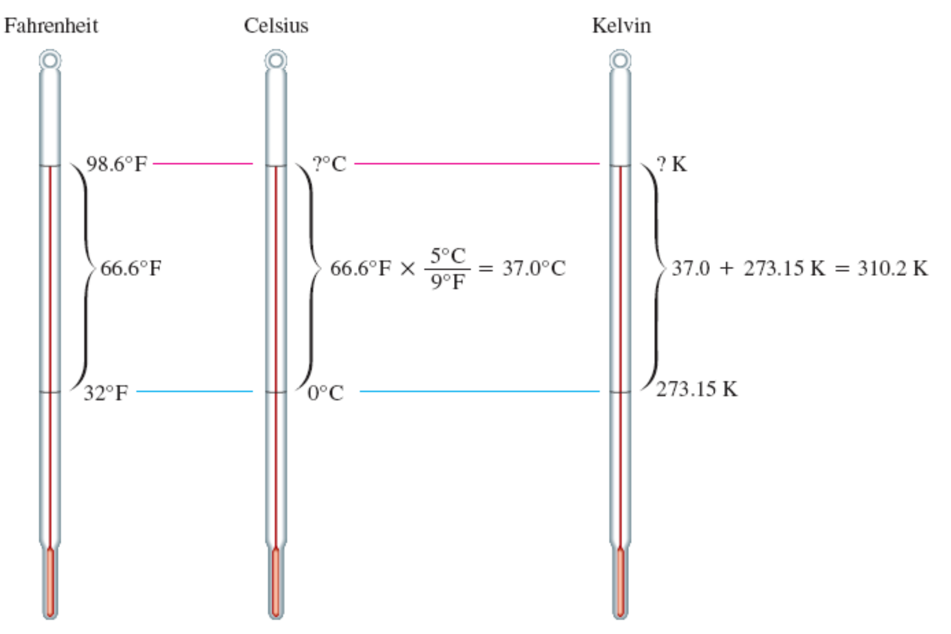
$$^{o}C = \frac{5}{9}(98.6 - 32.0) = 37.0\ ^{o}C$$

- *Convert to the Kelvin scale:*

$$T = 37.0\ ^{o}C + 273.15 = 310.2\ K$$

- *Note that the final answer has only one decimal place (37.0 is limiting)*



Exercise

One interesting feature of the Celsius and Fahrenheit scales is that –40 ℃ and –40 ℉ represent the same temperature. Verify this information?

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- *The difference between 32 ℉ and – 40 ℉ is 72 ℉. The difference between 0 ℃ and – 40 ℃ is 40 ℃. The ratio of these is*

$$^{o}F = \frac{9}{5}(^{o}C) + 32$$

$$\frac{72\ ^{o}F}{40\ ^{o}C} = \frac{8 \times 9\ ^{o}F}{8 \times 5\ ^{o}C} = \frac{9\ ^{o}F}{5\ ^{o}C}$$

as required. Thus – 40 ℃ is equivalent to – 40 ℉.

Exercises

- *How many mL are in 1.63 L?*

$$1.63\ L \times \frac{1000\ mL}{1\ L} = 1630 mL = 1.63 \times 10^{3}\ mL$$

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

$$343\ \frac{m}{s} \times \frac{3600\ s}{1\ h} \times \frac{1\ mi}{1609\ m} = 767\ \frac{mi}{h}$$