

Chapter 1

Introduction: Matter and Measurements

Chapter 1 suggested problems

10th Ed. - 1.x: 9, 15, 17, 19, 25, 27, 33, 37, 39, 45, 49, 53

11th Ed. - 1.x: 11, 15, 17, 19, 24, 25, 27, 35, 37, 43, 45, 47, 51

links: **the world's smallest billboard** at <http://www.almaden.ibm.com/vis/stm/atomo.html>

Class Notes

I. The study of chemistry

A. What is chemistry? A working definition

1. Chemistry is the study of things made up of atoms and molecules
2. It is the study of why they behave the way they behave
3. It is the study of how to make them behave more usefully

B. Fundamental branches of chemistry

1. Analytical - the study of "what" and of "how much" is in substances
2. Biochemistry - the chemistry of life and living things, based on organic chemistry
3. Inorganic - the chemistry of all compounds that do not contain both carbon and hydrogen
4. Organic - the chemistry of all compounds that contain both carbon and hydrogen
5. Physical - the study of how the laws of physics affect atoms and molecules

C. Why is chemistry relevant?

1. A few examples: what problems do Ph.D. chemists I know study?

1. **Analysis of DNA and large biological molecules** at http://asweb.artsci.uc.edu/CollegeDepts/chemistry/fac_staff/facDetails.cfm?username=Limbach
2. Blood clotting mechanisms
3. The chemistry of cardiovascular disease
4. Breast cancer detection
5. Pharmaceuticals
6. Fertilizers and pesticides
7. Food chemistry
8. Personal hygiene & cosmetics
9. **Chemistry of surfaces** at <http://www.chem.ualberta.ca/%7Emcdermot/home.html>
10. Composite materials
11. Lubricants and additives
12. Nerve gas detection and detoxification
13. Detection of drugs and poisons

II. Classifications of matter based on state or composition

A. States of matter - solids, liquids, gases (vapors)

1. Changes of state (physical changes, phase changes) - melting, vaporization, sublimation, freezing, condensation, deposition

B. Composition

1. Element - all of the atoms are the same, can consist of more than one atom (O_2 , N_2 , S_8 , P_4)
2. Compound - made of two or more elements - H_2O , CO_2 , CH_3CH_2OH
3. Substance - cannot be separated into other kinds of matter by physical means (e.g., elements and compounds)
4. Mixture - can be separated into other kinds of matter by physical means; e.g. salt dissolved in water
 - a. Homogeneous mixture - uniform composition, properties, and appearance throughout the mixture - salt dissolved in water
 - b. Heterogeneous mixture - do not have uniform composition, properties, and appearance throughout the mixture - pizza, granite, salt and sand

III. Properties of matter, physical and chemical changes

A. The difference between changes of state (physical changes) and chemical reactions

1. Physical changes - changes involving physical properties, or, changes in which there is no change to the chemical composition of the material(s) under observation
 - a. Physical properties - properties which can be observed and measured without changing the composition of the material
 - i. Color, odor, state of matter, melting point, boiling point, heats of vaporization and fusion, density, solubility, metallic character, electrical and thermal conductivity, magnetic properties, crystal shape, malleability, ductility, viscosity, etc.
2. Most common physical changes either involve changes of state (phase transitions) or are the consequence of mechanical processing (e.g. grinding, crushing, slicing, pulverizing, gluing pieces together, etc.)

B. Chemical changes (chemical reactions) - starting materials are consumed and new materials are formed due to the breaking and making of chemical bonds

1. Chemical reactions *may* be indicated by color changes, the absorbing or release of energy (heat, light, sound electric), or by the formation of new materials such as gases, pure liquids, or solids (precipitates)

IV. Units of measurement

A. Base units

1. Mass - kilogram (kg)
2. Length - meter (m)
3. Time - second (s)
4. Temperature - Kelvin (K; note: *not* degrees Kelvin)

B. Prefixes

1.

tera	T	10^{12}
giga	G	10^9
mega	M	10^6
kilo	k	10^3
deci	d	10^{-1}
centi	c	10^{-2}

milli	m	10^{-3}
micro	u	10^{-6}
nano	n	10^{-9}
pico	p	10^{-12}
femto	f	10^{-15}
atto	a	10^{-18}

2. Note: much femto/atto scale work is at the cutting edge of chemistry

C. Measurements

1. Length: measured in m, cm, mm

2. Mass:

- a. Mass is measured in kilograms.
- b. There is a difference between mass and weight.
- c. Mass is the amount of matter present in a substance and does not change with respect to local gravitational fields.

3. Temperature

- a. Fahrenheit - based on the boiling point (BP) and the melting point (MP) of water at 212 °F and 32 °F respectively.
- b. Celsius (Centigrade) - based on based on the boiling point (BP) and the melting point (MP) of water at 100 °C and 0 °C respectively.
Note that 1 °C = 1.8 °F (i.e., the Celsius degree is bigger)
- c. Kelvin - based on 0 K at Absolute Zero - the point at which all molecular motion stops, MP water at 273 K, BP water at 373 K
Note that 1K = 1 °C (i.e., they are the same size)
- d. Relationships and conversions
 - i. F to C: $(F - 32) / 1.8$
 - ii. C to F: $(C \times 1.8) + 32$
 - iii. C to K: $C + 273$
 - iv. K to C: $K - 273$
 - v. K to F: convert K to C and then convert to F
 - vi. F to K: convert F to C and then convert to K

4. Volume - a derived quantity

- a. Volume is the product of length, width, and height.
- b. Volume is measured in cubic meters (m^3) but this is too large to be practical in many cases, so cubic centimeters (cm^3) is more common.
- c. $0.001 \text{ m}^3 = 1 \text{ dm}^3 = 1 \text{ liter (L)}$
- d. $1000 \text{ mL} = 1 \text{ L}$; $1 \text{ cm}^3 = 1 \text{ mL}$

5. Density - the ratio of mass and volume

- a. Density is an important physical property of substances
- b. Density is determined by the careful measurement of a substance's mass and volume
- c. Density gives information about a substance at both a macroscopic and a microscopic level
 - i. Given samples of two substances, the more dense material either has more particles per unit volume, heavier particles, or both
 - ii. Densities: Au - 19.32 g/cm^3 ; Ag - 10.5 g/cm^3 - what does this tell you about gold and silver at a microscopic level? (compare atomic masses)
- d. Examples
 - i. 10.0 g of liquid occupies 13.5 mL of space; density = $m/V = 0.741 \text{ g/mL}$
 - ii. Pure gold (24 K) has a density of 19.32 g/cm^3 . A wedding band weights 3.50 grams. How much water must it displace (i.e., what must its volume be?) if the band is in fact 24 K gold?
(answer: 0.181 cm^3)
 - iii. Benzene has a density of 0.880 g/mL . If you need exactly 78.12 grams of benzene, what volume would you measure out?
(answer: 88.77 mL)

V. Uncertainty in measurement

- A. Chemistry is empirical, based on observation and measurement
- B. There are always intrinsic limitations to all instruments, no matter how expensive
- C. Accuracy and precision
 1. Precision: the closeness of a series of measurements to each other

2. Accuracy - the closeness of one or of a series of measurements to the actual or correct value

D. Significant figures - used to convey instrumental limitations

1. all of the certain figures plus the first uncertain figure
2. Leading zeros
3. Captive zeros
4. Trailing zeros

E. Examples

1. 1234 - 4 sf
2. 10101 - captive zeros are significant - 5 sf
3. 0.000604999 - leading zeros - are not significant - 6 sf
4. 555,000 vs. 555,000. - trailing zeros are only significant if there is a decimal to indicate accuracy - 3 sf vs. 6 sf

- F. Rounding-off rules: 0-0.49, .51-.99; rounding reserved until all calculations are completed

VI. Dimensional analysis and problem solving

A. Conversion factors

1. Multiplicative identity: any number multiplied by 1 will result in that number i.e., $1 \times y = y$
2. A conversion factor permits the conversion from one set of units to another without changing the true value of the number associated with the units
3. Creation of conversion factors
 - a. $1 \text{ mile} = 5280 \text{ feet} \Rightarrow (1 \text{ mile} / 5280 \text{ feet}) = (5280 \text{ feet} / 5280 \text{ feet}) \Rightarrow (1 \text{ mile} / 5280 \text{ feet}) = 1$
 - b. $1 \text{ foot} = 12 \text{ inches} \Rightarrow (1 \text{ foot} / 12 \text{ inches}) = (12 \text{ inches} / 12 \text{ inches}) \Rightarrow (1 \text{ foot} / 12 \text{ inches}) = 1$
 - c. $1 \text{ hour} = 3600 \text{ seconds} \Rightarrow (1 \text{ hour} / 3600 \text{ seconds}) = (3600 \text{ seconds} / 3600 \text{ seconds}) \Rightarrow (1 \text{ hour} / 3600 \text{ seconds}) = 1$

B. Dimensional analysis - the process of problem solving using conversion factors (also called factor label method, factorial analysis, etc.)

1. The process using an example: convert 150 lbs to kg
2. What is given? how should the problem wind up?

- a. Given: 150 lbs; the problem should wind up with units in kg
3. What conversion factors are needed?
 - a. $1 \text{ kg} = 2.2 \text{ lbs} \Rightarrow (1 \text{ kg} / 2.2 \text{ lbs}) = 1$
4. After using the conversion factor do the units cancel? is another conversion factor needed?
 - a. $150 \text{ lbs} \times (1 \text{ kg} / 2.2 \text{ lbs}) = \text{kg}$; the units are correct, no additional conversion factors are needed
5. Do the math
 - a. $150 \text{ lbs} \times (1 \text{ kg} / 2.2 \text{ lbs}) = 68.2 \text{ kg}$

C. Additional examples

1. Convert 62.5 cm to inches
 - a. $1 \text{ inch} = 2.54 \text{ cm}$
 - b. $62.5 \text{ cm} \times (1 \text{ inch} / 2.54 \text{ cm}) = 24.6 \text{ inches}$
2. Convert 3.62 ounces to mg
 - a. $1 \text{ lb} = 16 \text{ oz}$; $1 \text{ lb} = 453.6 \text{ g}$; $1 \text{ g} = 1000 \text{ mg}$
 - b. $(3.62 \text{ oz}) \times (1 \text{ lb} / 16 \text{ oz}) \times (453.6 \text{ g} / 1 \text{ lb}) \times (1000 \text{ mg} / 1 \text{ g}) = 102,627 \text{ mg} = 1.02 \times 10^5 \text{ mg}$
3. Convert 1100 m/s to mph
 - a. $60 \text{ sec} = 1 \text{ minute}$; $60 \text{ min} = 1 \text{ hr}$; $1 \text{ m} = 100 \text{ cm}$; $1 \text{ in} = 2.54 \text{ cm}$; $1 \text{ ft} = 12 \text{ in}$; $1 \text{ mile} = 5280 \text{ ft}$
 - b. $(1100 \text{ m/s}) \times (60 \text{ s} / 1 \text{ min}) \times (60 \text{ min} / 1 \text{ hr}) \times (100 \text{ cm} / \text{m}) \times (1 \text{ in} / 2.54 \text{ cm}) \times (1 \text{ ft} / 12 \text{ in}) \times (1 \text{ mile} / 5280 \text{ ft}) = 2460.6 \text{ mile/hr} = 2500 \text{ mph}$
4. Convert the area of an 8.5" by 11" piece of paper to cm^2
 - a. $1 \text{ in} = 2.54 \text{ cm}$
 - b. $8.5 \text{ in} \times 11 \text{ in} = 93.5 \text{ in}^2$
 - c. $(93.5 \text{ in}^2) \times (2.54 \text{ cm} / 1 \text{ in})^2 = 603.2 \text{ cm}^2 = 6.0 \times 10^2 \text{ cm}^2$
5. Convert 6 yds^3 of cement to cm^3
 - a. $1 \text{ yd} = 3 \text{ ft}$; $1 \text{ ft} = 12 \text{ in}$; $1 \text{ in} = 2.54 \text{ cm}$
 - b. $(6 \text{ yds}^3) \times (3 \text{ ft} / 1 \text{ yd})^3 \times (12 \text{ in} / 1 \text{ ft})^3 \times (2.54 \text{ cm} / 1 \text{ in})^3 =$

$$4,587,329.1 \text{ cm}^3 = 5 \times 10^6 \text{ cm}^3$$

6. Hexane is an organic liquid used in the manufacture of gasoline. If a railroad car contains 2.5×10^4 gallons of hexane, how much does the liquid weigh in pounds? (note: the density of hexane is 0.6594 g/mL)

a. $1 \text{ gal} = 4 \text{ qt}$; $1 \text{ L} = 1.057 \text{ qt}$; $1000 \text{ mL} = 1 \text{ L}$; $1 \text{ lb} = 453.6 \text{ g}$

b. $(2.5 \times 10^4 \text{ gallons}) \times (4 \text{ qt} / 1 \text{ gal}) \times (1 \text{ L} / 1.057 \text{ qt}) \times (1000 \text{ mL} / 1 \text{ L})$
 $= 94,607,379.376 \text{ mL}$

c. $(94,607,379.376 \text{ mL}) \times (0.6594 \text{ g} / \text{mL}) \times (1 \text{ lb} / 453.6 \text{ g}) = 137,531.1$
 $\text{lbs} = 1.4 \times 10^5 \text{ lbs}$

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