Turn off cell phones

Dr. Stephen M. Holmes

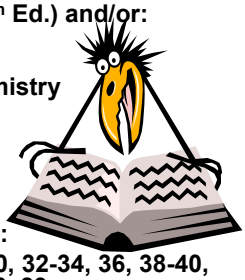
Office Hours: Monday: 1:00 – 3:00 (CP-15)
or by appointment

E-mail: smholmes2@gmail.com

Class Section: CHE105-001 MWF 11-10:50 AM

Homework: Due Jan. 9

Read CH. 1 from Chang text (9th Ed.) and/or:



View from Thinkwell:

- **1.1.1 An Introduction to Chemistry**
- **1.1.2 The Scientific Method**
- **1.2.1 States of Matter**
- **1.2.4 Properties of Matter**

and

Work End-of-Chapter problems:

Ch 1: 2, 3, 5-10, 12, 16, 21-26, 30, 32-34, 36, 38-40, 42-50 (even), 62, 65, 69, 70, 86, 89

Note: Keeping ahead of lecture will help you understand the material better!

Syllabus Topics

- **Textbook:**
  - **R. Chang *Chemistry*, 9th Ed.**
    - **Contains Book & Student Study Guide**
  - **Thinkwell Computer Disks**
- **Schedule:**
  - **Exam Dates and Times**
  - **Suggested Homework**
  - **Tips for Success**

Syllabus Topics

- **Homework Due Dates (page 2) – more later!**
- **Important Dates (page 3):**
  - **Jan. 15, last day to *add* course**
  - **Jan. 30, last day to *drop* course**
  - **Jan. 30, last day to *pass/fail* option**
  - **Mar. 7, last day to withdraw (Mar. 28, 1st-time Freshmen)**
- **Exams (page 4):**
  - **CALCULATORS: practice with the one you are allowed to use. *See syllabus for proper ones.***
  - **Don't miss the final!**
  - **See rules for "allowed conflicts"**
  - **Make-up exams? = NONE.**

Syllabus Topics

- **Grading (page 4):**
  - Assigned Homework (11 total), on-line (15%)
  - Final is comprehensive (25%)
  - Exams (3 x 20% = 60%)
- **Extra Credit Opportunities:**

  These add-on "bonus points" can add at most one (1) grade point to your final course average. *Earn them by attending General Chemistry Learning Center.*
- **Exam Help Sessions (page 5):**

Jan. 28	CP-139	4-5 PM	Holmes
Feb. 4	CP-139	4-5 PM	Ades
Feb. 25	CP-139	4-5 PM	Holler
Mar. 3	CP-139	4-5 PM	Holmes
Mar. 31	CP-139	4-5 PM	Ades
Apr. 7	CP-139	4-5 PM	Holler

Turn off cell phones

HW, Due Jan. 9

Read from Chang: 1.8-1.9

or

View from Thinkwell:

- 1.3.4 (Significant figures) and
- 1.3.5 (Dimensional Analysis)

and

Work End-of-Chapter problems:

Ch 1: 2, 3, 5-10, 12, 16, 21-26, 30, 32-34, 36, 38-40, 42-50 (even), 62, 65, 69, 70, 86, 89

Syllabus Topics

- **Make-up Exams** (page 5)

  Opportunity to Change to CHE 104 (Geri Gerke, CP-125)
- **By-Pass Exam** (see Geri Gerke, CP-125; Fall term only)
- **Cheating**

  Possible penalties = zero grade (w/o score replacement)
  Course failure (w/ no possibility to erase from record)
  Expulsion from University (see Senate Rules 6.3.0 etc.)
- **Help**

  Tutors (page 5), office hours (page 6)

  E-mail communication (page 6)

  Looking for readable with lots of examples & details?
  See any *Chemistry* textbook written by Steven S. Zumdahl

CHAPTER 1

Chemistry:
The Study of Change

1.1 Chemistry: A Science for the Twenty-First Century

- **Chemistry** - The study of matter and the changes it undergoes.

Health and Medicine

Food and Agriculture

Energy and the Environment

Materials and Technology

1.2 The Study of Chemistry

- Transformation of elements, molecules, or compounds into other species/materials
- Often absorbs or releases energy
- Rates of reactions often important (e.g. kinetics)
- Biological Processes involve chemistry
- Chemistry is a "work in progress"

  many unresolved an unexplained chemical issues still a developing field of study
- How do you study chemistry?

  using the scientific method!

1.3 The Scientific Method: Steps

1. Make observations

   QUALITATIVE: simple and general observation

   (often easy, but w/ little information)

   the sky is blue ; CP-139 is hot
   mercury is a liquid at 298 K

   QUANTITATIVE: requires an experiment; also called a measurement

   the temperature of water is 54 ºC

2. Look for patterns in data (observations)
3. After several measurements develop an assumption

   HYPOTHESIS: a tentative (temporary or preliminary) explanation for a set of observations
4. Design/perform more experiments to check validity of your assumptions and understanding of the problem.

   alter hypothesis if data suggests your understanding is wrong and/or too simple to account for observations made

   A short description of the valid hypothesis is called a….

   LAW: a concise verbal or mathematical statement describing relationship between phenomena that is always the same under the same conditions (e.g. F = ma)
5. Develop a theory

   hypotheses that survive several cycles of this are termed **THEORIES**

Theory Development: Solar System Example

Theory - unifying principle that explains lots of facts and laws based on them ; often called models

additional experiments test these theories and ideally correct errors

Egyptians (2000 B.C.): Sun is a boat inhabited by Ra (a god) that daily sails the sky

Exodus (400 B.C.): earth-centered nested spheres of planets/stars

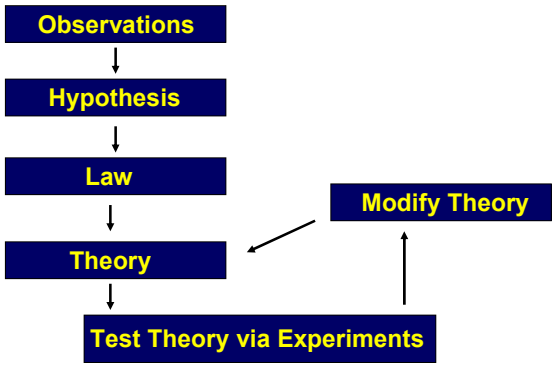
Ptolemy (100 A.D.): planets roll around Exodus spheres - explains planetary motion

Copernicus (1543): sun-centered circular planetary orbits
Kepler: elliptical planetary orbits around sun
Newton: gravity of sun keeps planets orbiting
Einstein: Newton explanation (mechanics) too simple - special case (simple) of quantum mechanics

1.4 Classification of Matter

- Remember: Chemistry – The study of matter and the changes it undergoes.
- **Matter** - Anything which has mass and takes up space (volume).
- If you can touch it, it is certainly matter.
- What about air? Does it take up space?

Fundamental Steps in Scientific Method



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

1. ***Homogeneous mixture*** – composition of the mixture is the same throughout.

   soft drink, fresh air, solder

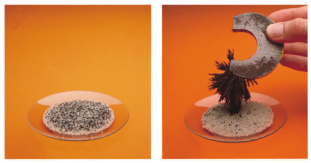
2. ***Heterogeneous mixture*** – composition is not uniform throughout.

cement,
iron filings in sand

Physical means **can be used to separate a mixture into its pure components.**

distillation

e.g. ethanol from water "moonshine"



magnetic separation

Note: Components remain unchanged

A ***compound*** **is a substance composed of atoms of two or more elements chemically united in fixed proportions.**

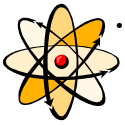
Compounds can only be separated by ***chemical*** **means into their pure components (elements).**

Water (H_2O) **Sucrose ($C_6H_{12}O_6$)**

Ammonia (NH_3)

An ***element*** **is a substance that cannot be separated into simpler substances by** ***chemical means. Building blocks for all materials.***

- **115 elements have been identified**
  - **83 elements occur naturally on Earth**
    gold, aluminum, lead, oxygen, carbon
  - **32 elements have been created by scientists**
    technetium, americium, seaborgium



and counting…..

Classification of Matter

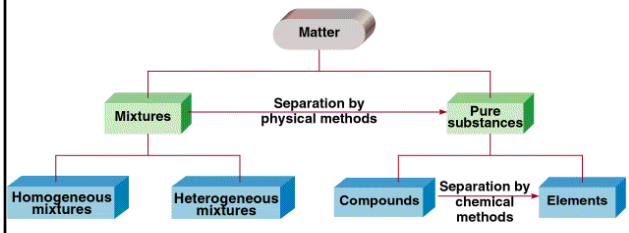
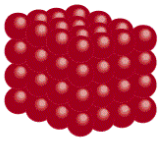


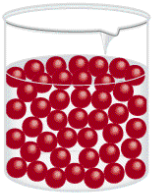
Figure 1.5 in Chang, 9th Ed.

1.5 The Three States of Matter

Figures 1.6 and 1.7



Solid



Liquid

Gas

1. **Extensive property** – depends on how much matter is being considered.

   Examples:

   - **Mass** – the quantity of matter in a given sample of a substance
   - **Volume** – length cubed

2. **Intensive Properties** – does not depend on how much matter is being considered.

   Example:

   - **Density** – mass of an object divided by its volume

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

1.6 Physical or Chemical Properties of Matter

A *physical change* does not alter the composition or identity of a substance.

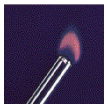
ice melting

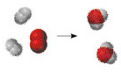
sugar dissolving in water

A *chemical change* alters the composition or identity of the substance(s) involved.

hydrogen gas burns in oxygen gas to form water

Figure on pg. 15





1.7 Measurement

SI Units – *International System of Units*

Based on the metric system

Metric units (multiples of 10)

Table 1.2: 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

All other units of measurement can be derived from these base units
Think of these as root words

mass – resistance of an object to a change in its state of motion (formal def.)

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

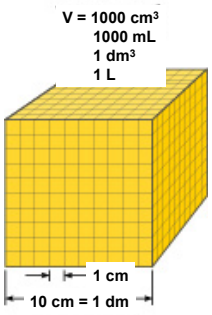
Table 1.3: Prefixes Used with SI Units

Prefix + root = unit of measure

Prefix	Symbol	Meaning	Example
Tera-	T	1,000,000,000,000 or 10^{12}	1 Tm = 1x10^{12} m
Giga-	G	1,000,000,000 or 10^9	1 Gm = 1 x 10^9 m
Mega-	M	1,000,000 or 10^6	1 Mm = 1 x 10^6 m
Kilo-	k	1,000 or 10^3	1 km = 1 x 10^3 m
Deci-	d	1/10 or 10^{-1}	1 dm = 0.1 m
Centi-	c	1/100 or 10^{-2}	1 cm = 0.01 m
Milli-	m	1/1,000 or 10^{-3}	1 mm = 0.001 m
Micro-	μ	1/1,000,000 or 10^{-6}	1 μm = 1 x 10^{-6} m
Nano-	n	1/1,000,000,000 or 10^{-9}	1 nm = 1 x 10^{-9} m
Pico-	p	1/1,000,000,000,000 or 10^{-12}	1 pm = 1 x 10^{-12} m

Note: Memorize these!

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



1 cm^3 = (1 x 10^{-2} m)^3 = 1 x 10^{-6} m^3

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

1 mL = 1 cm^3

Figures 1.8 and 1.10

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

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

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

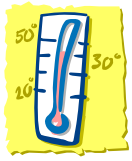
A piece of platinum metal with a density of 21.5 g/cm^3 has a volume of 4.49 cm^3. What is its mass?

Answer = 96.5 g

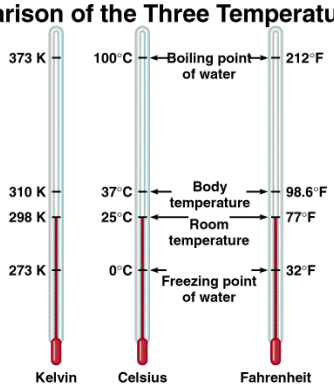
Convert 172.9 ^{0}F to degrees Celsius.

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

YOU DO ON YOUR OWN



Comparison of the Three Temperature Scales



$K = {}^0C + 273.15$

273 K = 0 ^{0}C
373 K = 100 ^{0}C

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

32 ^{0}F = 0 ^{0}C
212 ^{0}F = 100 ^{0}C

Convert 172.9 ^{0}F to degrees Celsius.

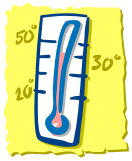
$$^0F = \frac{9}{5} \times {}^0C + 32$$

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

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

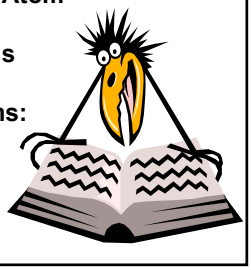
$$^0C = \frac{5}{9} \times ({}^0F - 32)$$

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



HW, Due Jan. 16

Read from Chapter 2 Chang:
or
View from Thinkwell:
2.1.1 Early Discoveries and the Atom
2.1.2 Understanding Electrons
2.1.3 Understanding the Nucleus
and
Work End-of-Chapter 2 problems:
5, 7, 8, 14-18 (even), 22, 24, 26, 27, 28, 30-36, 40, 44-48 (even), 57-60, 64, 66, 68, 69, 74, 90, 99



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}

1.8 Handling Numbers: Scientific Notation

The number of atoms in 12 g of carbon:

602,200,000,000,000,000,000,000 or 6.022×10^{23}

The mass of a single carbon atom in grams:

0.0000000000000000000000199 or 1.99×10^{-23}

N is a number between 1 and 10 | $N \times 10^{n}$ | *n* is a positive or negative integer

Scientific Notation

Multiplication

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

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

Division

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

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

Significant Figures

Addition or Subtraction

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

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 x 10^4 molecules	2 significant figures
560 kg	2 significant figures

Note: zero is a place holder in 560
If it were 560. then *three* sig. figs.

Significant Figures

Multiplication or Division

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

4.51 x 3.6666 = 16.536366 = 16.5
3 sig figs round to 3 sig figs

6.8 ÷ 112.04 = 0.0606926 = 0.061
2 sig figs round to 2 sig figs

Significant Figures

Exact Numbers

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

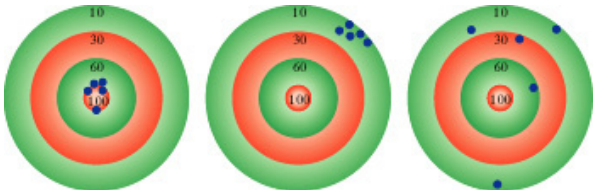
The average of three measured lengths; 6.64, 6.68 and 6.70?

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

Why? Because 3 is an exact number

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

Precision – how close measurements set are to each other



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

Accuracy –

Precision –

YOU READ AND BE RESPONSIBLE FOR THIS

1.9 Dimensional Analysis in Solving Problems

- !!!!!!!!!!!!VERY IMPORTANT!!!!!!!!!!!!!!
- Another name: Factor Label Method
  - Used to convert between desired units
- Let your units do the work for you by simply memorizing connections between units.
  - For example: How many donuts are in one dozen?

Dimensional Analysis, Cont.

- We say: "Twelve donuts are in a dozen."
- Or: 12 donuts = 1 dozen donuts

- **What does any number divided by itself equal?**
- **ONE!**

$$\frac{12\ \text{donuts}}{1\ \text{dozen}} = 1$$

Called a unit factor

Dimensional Analysis, Cont.

- **We use these two mathematical facts to do the factor label method**
  - a number divided by itself = 1
  - any number times one gives that number back
- **Let us see some examples of using factor label method.**

Dimensional Analysis, Cont.

- **What does any number times one equal?**
- **That number.**

How many donuts are in 3 ½ dozen?

How many meters are in 3 km?

If a box is 3.0 cm by 5.0 m by 4.0 in. What is the volume of the box in cm^3? (2.54 cm = 1 in)

The speed of sound in air is about 343 m/s.

What is this speed in miles per hour?

1 mi = 1609 m 1 min = 60 s 1 h = 60 min.

How many cm^3 are in 15 m^3?

Note: Also know English-Metric Conversions.

2.54 cm = 1 inch (in.) 1 pound (lb) = 453.6 g

1 L = 1.0567 quarts (qt)