General Chemistry (101 Chem) for Natural Sciences Students (3+4 contact hours)

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Textbook
Chemistry, 10th, by: Raymond Chang
McGraw-Hill, 2010

Office Hours:
Saturday 8:00 –12:00 a.m., Tuesday: 8:00 –12:00 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. Final exam ((3x30=90)/125)
  3. Practical (28/125)
  4. Oral exam (7/125)

No make up exam is allowed
Keys to Success

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

You don’t really understand something unless you can explain it to your grandmother.
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
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* (magnet, chromatography, filtration, distillation, crystallization, extraction, ...).
Separating Mixtures

**Filtration** : Separates components of a mixture based upon *differences in particle size*. Normally separating a precipitate from a solution, or particles from an air stream.

**Crystallization** : Separation is based upon *differences in solubility* of components in a mixture.

**Distillation** : Separation is based upon *differences in volatility*.

**Extraction** : Separation is based upon *differences in solubility* in different solvents (major material).

**Chromatography** : Separation is based upon *differences in solubility* in a mobile solvent versus a stationary phase.
An **element** is a substance that **cannot** be separated into simpler substances by **chemical means**. There are 117 elements known, of which 82 elements occur naturally on Earth (e.g., gold, aluminum, lead, oxygen, carbon) and 35 elements have been created by scientists (e.g., technetium, americium, seaborgium)

### Some Common Elements and Their Symbols

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

<table>
<thead>
<tr>
<th>Substance</th>
<th>Mixture</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Matter with constant composition</strong></td>
<td><strong>Matter with variable composition</strong></td>
</tr>
<tr>
<td><strong>Element</strong></td>
<td><strong>Compound</strong></td>
</tr>
<tr>
<td>Made up of only one type of atom</td>
<td>Made up of two or more types of atoms chemically united together</td>
</tr>
<tr>
<td><strong>Examples</strong></td>
<td><strong>Examples</strong></td>
</tr>
<tr>
<td>gold, silver, carbon, oxygen and hydrogen</td>
<td>water, carbon dioxide, sodium bicarbonate, carbon monoxide</td>
</tr>
</tbody>
</table>
# Three States of Matter

<table>
<thead>
<tr>
<th>State</th>
<th>Shape</th>
<th>Volume</th>
<th>Compressible?</th>
<th>Flow?</th>
</tr>
</thead>
<tbody>
<tr>
<td>Solid</td>
<td>Rigid</td>
<td>Maintains</td>
<td>No</td>
<td>No</td>
</tr>
<tr>
<td>Liquid</td>
<td>Assumes Container’s</td>
<td>Maintains</td>
<td>No</td>
<td>Yes</td>
</tr>
<tr>
<td>Gas</td>
<td>Assumes Container’s</td>
<td>Assumes Container’s</td>
<td>Yes</td>
<td>Yes</td>
</tr>
</tbody>
</table>

- Evaporation or reduction in Pressure
- Liquefaction or increase in pressure
- Melting
- Freezing
- Sublimation
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. Boiling, condensing, melting, freezing, subliming, dissolving

A chemical change alters the composition or identity of the involved substance(s); e.g., burning of H₂ in air to form water, rusting of iron, fading of dyes or bleaching of colors

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, color, 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)

$1 \text{ kg} = 1000 \text{ g} = 1 \times 10^3 \text{ g}$

**weight** – force that gravity exerts on an object

$\text{weight} = c \times \text{mass}$

- on earth, $c = 9.81 \text{ m s}^{-2}$
- on moon, $c = (9.81/6) \text{ m s}^{-2}$

**Mass is measured using a BALANCE (Kg).**

**Weight is measured using SCALES (Kg m s$^{-2}$ or Newton)**

A 1 kg bar will weigh

- 9.81 kg ms$^{-2}$ on earth
- 9.81/6 kg ms$^{-2}$ on moon

---

**Measurement**

All measurements have three parts:

- **Examples:** 33.2 mL, 426 kg, 72.36 mm, 31 people
- **A value:** 26.9762 g
- **An Uncertainty**
- **Units**
The SI System - Système International d’Unitès

A complete system of units adequate for the entire realm of physical science.

### SI Base Units

<table>
<thead>
<tr>
<th>Base Quantity</th>
<th>Name of Unit</th>
<th>Symbol</th>
</tr>
</thead>
<tbody>
<tr>
<td>Length</td>
<td>meter</td>
<td>m</td>
</tr>
<tr>
<td>Mass</td>
<td>kilogram</td>
<td>kg</td>
</tr>
<tr>
<td>Time</td>
<td>second</td>
<td>s</td>
</tr>
<tr>
<td>Electrical current</td>
<td>ampere</td>
<td>A</td>
</tr>
<tr>
<td>Temperature</td>
<td>kelvin</td>
<td>K</td>
</tr>
<tr>
<td>Amount of substance</td>
<td>mole</td>
<td>mol</td>
</tr>
<tr>
<td>Luminous intensity</td>
<td>candela</td>
<td>cd</td>
</tr>
</tbody>
</table>

### TABLE 1.2 SI Prefixes

<table>
<thead>
<tr>
<th>Multiple</th>
<th>Prefix</th>
</tr>
</thead>
<tbody>
<tr>
<td>$10^{18}$</td>
<td>exa (E)</td>
</tr>
<tr>
<td>$10^{15}$</td>
<td>peta (P)</td>
</tr>
<tr>
<td>$10^{12}$</td>
<td>tera (T)</td>
</tr>
<tr>
<td>$10^{9}$</td>
<td>giga (G)</td>
</tr>
<tr>
<td>$10^{6}$</td>
<td>mega (M)</td>
</tr>
<tr>
<td>$10^{3}$</td>
<td>kilo (k)</td>
</tr>
<tr>
<td>$10^{2}$</td>
<td>hecto (h)</td>
</tr>
<tr>
<td>1</td>
<td>deca (da)</td>
</tr>
<tr>
<td>$10^{-1}$</td>
<td>deci (d)</td>
</tr>
<tr>
<td>$10^{-2}$</td>
<td>centi (c)</td>
</tr>
<tr>
<td>$10^{-3}$</td>
<td>milli (m)</td>
</tr>
<tr>
<td>$10^{-6}$</td>
<td>micro (μ)(^a)</td>
</tr>
<tr>
<td>$10^{-9}$</td>
<td>nano (n)</td>
</tr>
<tr>
<td>$10^{-12}$</td>
<td>pico (p)</td>
</tr>
<tr>
<td>$10^{-15}$</td>
<td>femto (f)</td>
</tr>
<tr>
<td>$10^{-18}$</td>
<td>atto (a)</td>
</tr>
</tbody>
</table>

\(^a\)The Greek letter \(μ\) (pronounced “mew”).
SI System of Measurement

Rules for Using the SI Systems

1- Use only singular form of units and do **NOT** use a period after the symbol for the unit. (15.1 m length **NOT** 15.1 m. length)

2- Use a dot on the base line for the decimal point. 23.6 m **NOT** 23,6 m

3- Group digits in threes around the decimal point and do **NOT** use commas. 1 000 000.000 003 km **NOT** 1,000,000.000003

4- Do **NOT** use spaces for four-digit measurements. 1645 mL or 0.2367 mg **NOT** 164 5 mL or 0.236 7

5- Do **NOT** use the degree sign (°) for temperature recorded for the Kelvin temperature scale. 78.6 K **NOT** 78.6 °K (Use it for the °C & °F scales)
<table>
<thead>
<tr>
<th>Quantity</th>
<th>English to SI Equivalent</th>
</tr>
</thead>
</table>
| **Length** | 1 mile = 1.609 km  
1 yard = 0.9144 m  
1 foot (ft) = 0.3048 m  
1 inch = 2.54 cm (exactly!) |
| **Volume** | 1 cubic foot = 0.0283 m³  
1 gallon = 3.785 dm³  
1 quart = 0.9464 dm³ (Lt.)  
1 quart = 946.4 cm³  
1 fluid ounce = 29.6 cm³ |
| **Mass** | 1 pound (lb) = 0.4536 kg  
1 pound (lb) = 453.6 g  
1 ounce = 28.35 g |
**Volume** – SI derived unit for volume is cubic meter (m³)

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

\[
1 \text{ dm}^3 = (1 \times 10^{-1} \text{ m})^3 = 1 \times 10^{-3} \text{ m}^3
\]

\[
1 \text{ L} = 1000 \text{ mL} = 1000 \text{ cm}^3 = 1 \text{ dm}^3
\]

\[
1 \text{ mL} = 1 \text{ cm}^3
\]

A 10-cm cube contains 1000 1-cm cubes

**Figure 1.10** Comparison of two volumes, 1 mL and 1000 mL.

Ashraf A. Mohamed, Prof. Dr.
Density – SI derived unit for density is kg/m³

\[ 1 \text{ g/cm}^3 = 1 \text{ g/mL} = 1 \text{ kg/L} = 1000 \text{ kg/m}^3 \]

\[ \text{Density} (d) = \frac{\text{Mass} (m)}{\text{Volume} (V)} \]

- For equal volumes, denser object has larger mass
- For equal masses, denser object has smaller volume
- Heating an object generally causes it to expand (except for polymers), therefore the density changes with temperature

A piece of platinum metal with a density of 21.5 g/cm³ has a volume of 4.49 cm³. What is its mass?

\[ d = \frac{m}{V} \]

\[ \therefore m = d \times V = 21.5 \text{ g/cm}^3 \times 4.49 \text{ cm}^3 = 96.5 \text{ g} \]

Example: Decide if a ring with a mass of 3.15 g that displaces 0.233 cm³ of water is platinum

\[ d = \frac{m}{V} \]

\[ \therefore d' = \frac{3.15 \text{ g}}{0.233 \text{ cm}^3} = 13.5 \text{ g/cm}^3 \]

Density of platinum = 21.4 g/cm³, therefore not platinum
# Derived Units

<table>
<thead>
<tr>
<th>Base Quantity</th>
<th>Common Units</th>
</tr>
</thead>
<tbody>
<tr>
<td>Volume</td>
<td>dm³</td>
</tr>
<tr>
<td>Density</td>
<td>kg/dm³</td>
</tr>
<tr>
<td>Acceleration</td>
<td>m/s²</td>
</tr>
<tr>
<td>Force</td>
<td>kg x m/s²</td>
</tr>
</tbody>
</table>
$K = ^{0}\text{C} + 273.15$

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

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

$0\ K = \text{absolute zero}$

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

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

$212\ ^{0}\text{F} = 100\ ^{0}\text{C}$
# Comparison of Temperature Scales

<table>
<thead>
<tr>
<th>Set Points</th>
<th>Fahrenheit, °F</th>
<th>Celsius, °C</th>
<th>Kelvin, K</th>
</tr>
</thead>
<tbody>
<tr>
<td>water boils</td>
<td>212</td>
<td>100</td>
<td>373</td>
</tr>
<tr>
<td>body temperature</td>
<td>98.6</td>
<td>37</td>
<td>310</td>
</tr>
<tr>
<td>water freezes</td>
<td>32</td>
<td>0</td>
<td>273</td>
</tr>
<tr>
<td>absolute zero</td>
<td>-460</td>
<td>-273</td>
<td>0</td>
</tr>
</tbody>
</table>

Convert 172.9 °F to degrees Celsius.

\[
\begin{align*}
0^\circ F &= \frac{9}{5} \times 0^\circ C + 32 \\
0^\circ F - 32 &= \frac{9}{5} \times 0^\circ C \\
\frac{5}{9} \times (0^\circ F - 32) &= 0^\circ C \\
0^\circ C &= \frac{5}{9} \times (0^\circ F - 32) \\
0^\circ C &= \frac{5}{9} \times (172.9 - 32) = 78.3
\end{align*}
\]
# Temperature Conversion Formulas

<table>
<thead>
<tr>
<th>Conversion</th>
<th>Formula</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>Celsius to Kelvin</td>
<td>( K = ^0C + 273 )</td>
<td>( 21^\circ C = 294 , K )</td>
</tr>
<tr>
<td>Kelvin to Celsius</td>
<td>( ^0C = K - 273 )</td>
<td>( 313 , K = 40 , ^\circ C )</td>
</tr>
<tr>
<td>Fahrenheit to Celsius</td>
<td>( ^0C = \left(^0F - 32\right) \times \frac{5}{9} )</td>
<td>( 89 , ^\circ F = 31.7 , ^\circ C )</td>
</tr>
<tr>
<td>Celsius to Fahrenheit</td>
<td>( ^0F = \left(^0C \times \frac{9}{5}\right) + 32 )</td>
<td>( 50 , ^\circ C = 122 , ^\circ F )</td>
</tr>
</tbody>
</table>

Practice – Convert 0.0°F into Kelvin
Chemistry In Action (The Gimli Glider)

In December 1998, NASA launched the 125-million dollar Mars Climate Orbiter, intended as the red planet’s first weather satellite. On 9/23/1999, the orbiter entered Mars’ atmosphere about 100 km (62 mi) lower than planned and was destroyed by heat. This was due to the failure to convert English measurement units into metric units in the navigation software.

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.”
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.

\[
given\ unit \times \frac{desired\ unit}{given\ unit} = desired\ unit
\]
\[
given\ unit \times \frac{related\ unit}{given\ unit} \times \frac{desired\ unit}{related\ unit} = desired\ unit
\]

How many mL are in 1.63 L?

\[
1\ L = 1000\ mL \\
1.63\ L \times \frac{1000\ mL}{1\ L} = 1630\ mL
\]
The speed of sound in air is about 343 m/s. What is this speed in miles per hour?

\[
343 \, \frac{\text{m}}{\text{s}} \times \frac{1 \, \text{mi}}{1609 \, \text{m}} \times \frac{60 \, \text{s}}{1 \, \text{min}} \times \frac{60 \, \text{min}}{1 \, \text{hour}} = 767 \, \frac{\text{mi}}{\text{hour}}
\]

How many mL are in 3.0 ft\(^3\)?

\[
(3.0 \, \text{ft}^3) \frac{(12 \, \text{in})^3 (2.54 \, \text{cm})^3 (1 \, \text{mL})}{(1 \, \text{ft})^3 (1 \, \text{in})^3 (1 \, \text{cm}^3)} = 8.5 \times 10^4 \, \text{mL}
\]

How many ns are in 23.8 s?

\[
(23.8 \, \text{s}) \frac{(10^9 \, \text{ns})}{(1 \, \text{s})} = 23.8 \times 10^9 \, \text{ns} = 2.38 \times 10^{10} \, \text{ns}
\]
Scientific Notation

The number of atoms in 12 g of carbon:

602 200 000 000 000 000 000 000

6.022 x 10^{23}

The mass of a single carbon atom in grams:

0.000 000 000 000 000 000 000 000 0199

1.99 x 10^{-23}

N is a number between 1 and 10.
May be a fraction

N x 10^n

n is a positive or negative integer
NOT a fraction
568.762 \quad \text{move decimal left} \quad n > 0 \quad +ve
\quad 568.762 = 5.68762 \times 10^2

0.00000772 \quad \text{move decimal right} \quad n < 0 \quad -ve
\quad 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

\begin{align*}
4.31 \times 10^4 + 3.9 \times 10^3 &= \quad 4.31 \times 10^4 + 0.39 \times 10^4 = 4.70 \times 10^4
\end{align*}

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

\begin{align*}
(4.0 \times 10^{-5}) \times (7.0 \times 10^3) &= \quad (4.0 \times 7.0) \times (10^{-5+3}) = \\
&= 28 \times 10^{-2} = 2.8 \times 10^{-1}
\end{align*}

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

\begin{align*}
8.5 \times 10^4 \div 5.0 \times 10^9 &= \quad (8.5 \div 5.0) \times 10^{4-9} = 1.7 \times 10^{-5}
\end{align*}
Significant Figures

• More significant figures means more sensitive measurements.

• Any digit that is not zero is significant
  1.234 kg   4 significant figures

• Zeros to the left of the first nonzero digit are not significant; however, zeros to the right are significant
  0.08 L   (1 sig. fig.)  80.0 (3 sig. figs.)

• Zeros between nonzero digits and those on the right are significant
  606 m  ,  660 m   3 sig. figs.
  0.00420 g  , 0.00502 g 3 sig. figs.

• In a measurement, the last significant figure is assumed to be uncertain and its range is ±1.

• The result of a calculation involving measured values can not be more certain than the least certain measurement. (3 SF + 4 SF = 3 SF)

• The number of significant figures in a result depends on the number of significant figures in the measurement and on the mathematical operation being performed.
<table>
<thead>
<tr>
<th>Measurement</th>
<th>Significant Figures</th>
</tr>
</thead>
<tbody>
<tr>
<td>24 mL</td>
<td>2 SF</td>
</tr>
<tr>
<td>3001 g</td>
<td>4 SF</td>
</tr>
<tr>
<td>0.0320 m³</td>
<td>3 SF</td>
</tr>
<tr>
<td>$6.4 \times 10^4$ molecules</td>
<td>2 SF</td>
</tr>
<tr>
<td>0.04450 m</td>
<td>4 SF</td>
</tr>
<tr>
<td>5.0003 km</td>
<td>5 SF</td>
</tr>
<tr>
<td>$10 \text{ dm} = 1 \text{ m}$</td>
<td>infinite number of SF</td>
</tr>
<tr>
<td>$1.000 \times 10^5$ s</td>
<td>4 SF</td>
</tr>
<tr>
<td>0.00002 mm</td>
<td>1 SF</td>
</tr>
<tr>
<td>10,000 m</td>
<td>5 SF !!, Ambiguous, generally assume 1 sig. fig.</td>
</tr>
<tr>
<td>560 kg</td>
<td>3 SF !!, Ambiguous, generally assume 2 sig. fig.</td>
</tr>
</tbody>
</table>
Significant Figures

Addition or Subtraction

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

\[
\begin{align*}
\text{89.332} & \quad \text{one SF after decimal point} \\
+ 1.1 & \quad \text{round off to 90.4} \\
\hline
90.432 & \\
\end{align*}
\]

\[
\begin{align*}
\text{3.70} & \quad \text{two SF after decimal point} \\
- 2.9133 & \quad \text{round off to 0.79} \\
\hline
0.7867 & \\
\end{align*}
\]

\[
\begin{align*}
35.2 \text{ mL} & + 0.34 \text{ mL} = 35.5 \text{ mL} \\
1.00794 \text{ g} + 1.00794 \text{ g} + 15.9994 \text{ g} &= 18.0153 \text{ g}
\end{align*}
\]
Multiplication or Division

The number of significant figures in the result is set by the original number that has the \textit{smallest} number of significant figures.

\[ 4.51 \times 3.6666 = 16.536366 = 16.5 \]

\[ \uparrow \quad \uparrow \]

3 sig figs \quad round to \quad 3 sig figs

\[ 6.8 \div 112.04 = 0.0606926 = 0.061 \]

\[ \uparrow \quad \uparrow \]

2 sig figs \quad round to \quad 2 sig figs

Round-off Rules –
\begin{itemize}
  \item For digits 0 - 4, do not round up.
  \item For digits 6 - 9, round up.
  \item For the digit 5, round up only to make the previous digit an even number.
\end{itemize}

Report to 3 significant figures.

\begin{align*}
10.235 & \Rightarrow 10.2 \\
12.4590 & \Rightarrow 12.4 \\
19.75 & \Rightarrow 19.8 \\
15.651 & \Rightarrow 15.6
\end{align*}

Ex: density = \((9.5760 \text{ g})/(12.2 \text{ mL}) = 0.784918 \Rightarrow 0.785 \text{ g/mL} \)
Significant Figures (Continued)

**Exact Numbers**

Numbers from definitions or numbers of objects are considered to have an infinite number of significant figures (they are certain).

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

\[
\frac{6.64 + 6.68 + 6.70}{3} = 6.67333 = 6.67 \rightarrow 7
\]

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
Accuracy vs. Precision

- Suppose three students are asked to determine the mass of an object whose known mass is 10.00 g
- The results they report are as follows

**Student A**
- Trial 1: 10.49 g
- Trial 2: 9.79 g
- Trial 3: 9.92 g
- Trial 4: 10.31 g
- Average: 10.13 g
- Inaccurate, imprecise

**Student B**
- Trial 1: 9.78 g
- Trial 2: 9.82 g
- Trial 3: 9.75 g
- Trial 4: 9.80 g
- Average: 9.79 g
- Inaccurate, precise

**Student C**
- Trial 1: 10.03 g
- Trial 2: 9.99 g
- Trial 3: 10.03 g
- Trial 4: 9.98 g
- Average: 10.01 g
- Accurate, precise
Homework

• After each chapter, solve the odd or even numbered questions!
• Just reading the questions would make a student familiar with various question styles.

• If you encountered any difficulty, my office is opened.
• You are completely welcomed in the office (+ Non office) hours!