CHAPTER ONE

Measurements
Math Review and Measurements

- Make measurements to understand the environment
  - Humans—sight, taste, smell, sound...
    - Limited and biased
  - Use instruments—meter sticks, thermometers, balances
    - More accurate and precise
- All measurements have units—METRIC SYSTEM vs. British System
Units of Measurement

Definitions

- **Time**
  - Interval or duration of forward events

- **Mass**
  - Measure of the quantity of matter in a body

- **Weight**
  - Measure of the gravitational attraction for a body ($w=mg$)
Units of Measurement

Definitions

- **Length**
  - Measure of space in any direction

![Meterstick](image)
- 1 meter = 39.4 inches

![Yardstick](image)
- 1 foot = 12 inches
Units of Measurement

Volume

- Amount of space occupied by a system
- Derived unit, mL or cc, cm³

Volume = 1 cm × 1 cm × 1 cm
= 1 cm³ = 1 mL

Volume = 10 cm × 10 cm × 10 cm = 1000 cm³
= 1000 mL
= 1 L
## Fundamental Quantities in Chemistry

<table>
<thead>
<tr>
<th>Quantity</th>
<th>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>temperature</td>
<td>Kelvin</td>
<td>K</td>
</tr>
<tr>
<td>amt. substance</td>
<td>mole</td>
<td>mol</td>
</tr>
</tbody>
</table>
## Measurements in Chemistry

Use Metric System

<table>
<thead>
<tr>
<th>Name</th>
<th>Symbol</th>
<th>Multiplier</th>
</tr>
</thead>
<tbody>
<tr>
<td>mega</td>
<td>M</td>
<td>$10^6$ (1,000,000)</td>
</tr>
<tr>
<td>kilo</td>
<td>k</td>
<td>$10^3$ (1,000)</td>
</tr>
<tr>
<td>deka</td>
<td>da</td>
<td>10</td>
</tr>
<tr>
<td>deci</td>
<td>d</td>
<td>$10^{-1}$ (0.1)</td>
</tr>
<tr>
<td>centi</td>
<td>c</td>
<td>$10^{-2}$ (0.01)</td>
</tr>
</tbody>
</table>
### Measurements in Chemistry

#### Metric Prefixes

<table>
<thead>
<tr>
<th>Name</th>
<th>Symbol</th>
<th>Multiplier</th>
</tr>
</thead>
<tbody>
<tr>
<td>milli</td>
<td>m</td>
<td>$10^{-3}(0.001)$</td>
</tr>
<tr>
<td>micro</td>
<td>µ</td>
<td>$10^{-6}(0.000001)$</td>
</tr>
<tr>
<td>nano</td>
<td>n</td>
<td>$10^{-9}$</td>
</tr>
<tr>
<td>pico</td>
<td>p</td>
<td>$10^{-12}$</td>
</tr>
<tr>
<td>femto</td>
<td>f</td>
<td>$10^{-15}$</td>
</tr>
</tbody>
</table>
Use of Numbers

- Exact numbers
  - 1 dozen = 12 things for example

- Measured Numbers—
  - use rules for significant figures
  - Use scientific notation were possible

- Accuracy
  - how closely measured values agree with the correct value

- Precision
  - how closely individual measurements agree with each other
Use of Numbers

- Piece of Black Paper – with rulers beside the edges
Use of Numbers

Piece of Paper Side B – enlarged

- How long is the paper to the best of your ability to measure it?

---

12  13  14
Use of Numbers

- Piece of Paper Side A – enlarged
  - How wide is the paper to the best of your ability to measure it?
Use of Numbers

- Significant figures
  - digits believed to be correct by the person making the measurement
- Exact numbers have an infinite number of significant figures

12.000000000000000 = 1 dozen
because it is an exact number
Significant figures

- Calculators give 8+ numbers
- People estimate numbers differently
- Scientists develop rules to help you determine which digits are “significant”
- Dictated by the precision (graduation) on your measuring device
- In the lab, the last significant digit is the digit you have to estimate
Use of Numbers

**Significant Figures - Rules**

- Non zero numbers are significant
- Leading zeroes are never significant
  - 0.000357 has three significant figures
- Imbedded zeroes are always significant
  - 3.0604 has five significant figures
- Trailing zeroes may be significant
  - Must specify significance by how the number is written
  - 1300 nails - counted or weighed?
  - 1.30000 - what was the precision of the measurement?
Use of Numbers

- Use scientific notation to remove doubt

  Express answers as powers of 10 by moving the
decimal place right (-) or left (+)

  2000 \rightarrow 2 \times 10^3
  15000 \rightarrow 1.5 \times 10^4
  0.004 \rightarrow 4 \times 10^{-3}
  0.000053 \rightarrow 0.0053 \times 10^{-3}

In scientific notation, zeros are given if they are significant

  1.000 \times 10^3 \text{ has 4 significant figures}
  2.40 \times 10^3 \text{ has ? significant figures}
Use of Numbers

- Multiplication & Division rule
  Easier of the two rules
  Product has the smallest number of significant figures of multipliers

\[
\begin{align*}
4.242 \\
\times 1.23 \\
\hline
5.21766
\end{align*}
\]
round off to 5.22
Use of Numbers

- Multiplication & Division rule
  Easier of the two rules
  Product has the smallest number of significant figures of multipliers

\[
\begin{align*}
4.242 & \times 1.23 = 5.21766 \\
2.7832 & \times 1.4 = 3.89648
\end{align*}
\]

round off to 5.22
round off to 3.9
Use of Numbers

- **Addition & Subtraction rule**
  
  More subtle than the multiplication rule
  
  Answer contains smallest decimal place of the addends.

\[
\begin{align*}
3.6923 & \\
+1.234 & \\
+2.02 & \\
\hline
6.9463 & \\
\end{align*}
\]

round off to 6.95
Use of Numbers

- **Addition & Subtraction rule**
  - More subtle than the multiplication rule
  - Answer contains smallest decimal place of the addends.

```
3.6923
+1.234
+2.02
---
6.9463
```

```
8.7937

---

-2.123

---

6.6707
```

round off to 6.95  round off to 6.671
Units of Measurement

Common Conversion Factors (Equalities)

- **Length**
  - 1 m = 39.37 inches
  - 2.54 cm = 1 inch

- **Volume**
  - 1 liter = 1.06 qt
  - 1 qt = 0.946 liter

- See Text and Handout for more conversion factors
Using Conversion factors

- Scientists often must convert between units.
- Conversion factors can be made for any relationship.
- Use known equivalence to make a fraction that can be used to “convert” from one unit to the other.
  - Fraction equals 1, so you are multiplying by 1
Factor Label Method

- 1 inch = 2.54 cm

- Ratio \( \frac{1\text{in}}{2.54\text{cm}} = 1 \) or \( 1 = \frac{2.54\text{cm}}{1\text{in}} \)

- Use ratio to perform a calculation. Units will divide out.

- Convert 60 in to centimeters

\[
60\text{in} \times \frac{2.54\text{cm}}{1\text{in}} = 152.4\text{cm}
\]
Practice

- Using your conversion handout
  - Convert 25 g to lbs
  - Convert 1 mL to Liters
  - Convert 20 meters to cm
More practice

Express 9.32 yards in millimeters.

\[
\frac{1 \text{ ft}}{12 \text{ in}} \quad \frac{12 \text{ in}}{1 \text{ ft}}
\]
More Practice

9.32 yd = ? mm

\[ 9.32 \text{ yd} \left( \frac{3 \text{ ft}}{1 \text{ yd}} \right) \]
More practice

9.32 yd = ? mm

\[ 9.32 \text{ yd} \left( \frac{3 \text{ ft}}{1 \text{ yd}} \right) \left( \frac{12 \text{ in}}{1 \text{ ft}} \right) \]
More Practice

9.32 yd = ? mm

\[ 9.32 \text{ yd} \left( \frac{3\text{ ft}}{1\text{ yd}} \right) \left( \frac{12\text{ in}}{1\text{ ft}} \right) \left( \frac{2.54\text{ cm}}{1\text{ in}} \right) \]
9.32 yd = ? mm

\[
9.32 \text{ yd} \times \left( \frac{3 \text{ ft}}{1 \text{ yd}} \right) \times \left( \frac{12 \text{ in}}{1 \text{ ft}} \right) \times \left( \frac{2.54 \text{ cm}}{1 \text{ in}} \right) \times \left( \frac{10 \text{ mm}}{1 \text{ cm}} \right) = 8.52 \times 10^3 \text{ mm}
\]
9.32 yd = ? mm

\[
9.32 \text{ yd} \left( \frac{3\text{ ft}}{1\text{ yd}} \right) \left( \frac{12\text{ in}}{1\text{ ft}} \right) \left( \frac{2.54\text{ cm}}{1\text{ in}} \right) \left( \frac{10\text{ mm}}{1\text{ cm}} \right) = 8.52 \times 10^3 \text{ mm}
\]
Your turn

- Example 1-2: Express 627 milliliters in gallons.

*You do it!*
Example 1-2. Express 627 milliliters in gallons.

\[
? \text{ gal } = 627 \text{ mL} \\
? \text{ gal } = 627 \text{ mL} \left( \frac{1 \text{ L}}{1000 \text{ mL}} \right) \left( \frac{1.06 \text{ qt}}{1 \text{ L}} \right) \left( \frac{1 \text{ gal}}{4 \text{ qt}} \right) \\
? \text{ gal } = 0.166155 \text{ gal} \approx 0.166 \text{ gal}
\]
Area – length x width

- Area is two dimensional thus units must be in squared terms.
- Example 1-4: Express $2.61 \times 10^4 \text{ cm}^2$ in $\text{ft}^2$. 
Area – length x width

- Area is two dimensional thus units must be in squared terms.

- Example 1-4: Express $2.61 \times 10^4 \text{ cm}^2$ in ft$^2$.

\[ \text{? ft}^2 = 2.61 \times 10^4 \text{ cm}^2 \left( \frac{1 \text{ in}}{2.54 \text{ cm}} \right) \]

- common mistake
Area – length x width

- Area is two dimensional thus units must be in squared terms.

- Example 1-4: Express $2.61 \times 10^4$ cm$^2$ in ft$^2$.

\[ ? \text{ ft}^2 = 2.61 \times 10^4 \text{ cm}^2 \left(\frac{1 \text{ in}}{2.54 \text{ cm}}\right) \]
Area – length x width

- Area is two dimensional thus units must be in squared terms.

Example 1-4: Express $2.61 \times 10^4$ cm$^2$ in ft$^2$.

\[ ? \text{ ft}^2 = 2.61 \times 10^4 \text{ cm}^2 \left(\frac{1 \text{ in}}{2.54 \text{ cm}}\right)^2 \]
Area – length x width

- Area is two dimensional thus units must be in squared terms.

- Example 1-4: Express $2.61 \times 10^4 \text{ cm}^2$ in ft$^2$.

$$? \text{ ft}^2 = 2.61 \times 10^4 \text{ cm}^2 \left(\frac{1 \text{ in}}{2.54 \text{ cm}}\right)^2 \left(\frac{1 \text{ ft}}{12 \text{ in}}\right)^2$$
Area – length x width

- Area is two dimensional thus units must be in squared terms.

Example 1-4: Express $2.61 \times 10^4 \text{ cm}^2$ in ft$^2$.

\[
? \text{ ft}^2 = 2.61 \times 10^4 \text{ cm}^2 \left(\frac{\text{lin}}{2.54\text{ cm}}\right)^2 \left(\frac{1\text{ ft}}{12\text{ in}}\right)^2 \\
= 28.09380619 \text{ ft}^2 \approx 28.1 \text{ ft}^2
\]
Volume – length x width x height

- Volume is three dimensional thus units must be in cubic terms—this volume is used in medical measurements--cc

- Example 1-5: Express $2.61 \text{ ft}^3$ in $\text{cm}^3$.

\[
? \text{ cm}^3 = 2.61 \text{ ft}^3 \left(\frac{12 \text{ in}}{1 \text{ ft}}\right)^3 \left(\frac{2.54 \text{ cm}}{1 \text{ in}}\right)^3
\]

\[
= 73906.9696 \text{ cm}^3 \approx 7.39 \times 10^4 \text{ cm}^3
\]
Percentage

- Percentage is the parts per hundred of a sample.
- Example 1-6: A 335 g sample of ore yields 29.5 g of iron. What is the percent of iron in the ore?

*You do it!*
Percentage

- Percentage is the parts per hundred of a sample.

- Example 1-6: A 100 kg person has 22 g of body fat. What is the percent of body fat in the person?

\[
? \% \text{ fat} = \frac{\text{kg of fat}}{\text{kg of person}} \times 100\%
\]

\[
= \frac{22 \text{ kg fat}}{100 \text{ kg person}} \times 100\%
\]

\[
= 22\%
\]
Medically Relevant Units

\[ \frac{g}{\text{tablet}} \text{ or } \frac{mg}{\text{tablet}} \text{ or } \frac{ug}{\text{tablet}} \text{ or } \frac{\text{active amount}}{\text{pill}} \]

\[ \frac{g}{dL} \text{ or } \frac{mg}{dL} \text{ or } \frac{ug}{dL} \frac{\text{amount of medicine}}{\text{amount of blood (urine, medicine)}} \]

\[ \frac{g}{cc} \text{ or } \frac{mg}{cc} \text{ or } \frac{ug}{cc} \text{ or } \frac{\text{cc}}{cm^3} = mL \]

\[ \frac{\text{mg dose}}{\text{kg body mass}} \]
A dose of 40 mg of antibiotic is required for a patient. The tablets contain 100 µg each.

How many tablets should the patient receive?
An injection requires a dose of 6 mg/kg body weight. The patient weights 200 pounds. How much medicine should she receive? The medicine is supplied as 200 mg/mL. How many mL should the injection be?
Density and Specific Gravity

- What is density?
- Density (how compact something is, mass per unit volume)
- density = mass/volume

Cork (D = 0.26 g/mL)
Ice (D = 0.92 g/mL)
H₂O (D = 1.0 g/mL)
Aluminum (D = 2.70 g/mL)
Lead (D = 11.3 g/mL)
Density and Specific Gravity

Example 1-7: Calculate the density of a substance if 742 grams of it occupies 97.3 cm$^3$.

\[
1 \text{ cm}^3 = 1 \text{ mL} \quad \therefore \quad 97.3 \text{ cm}^3 = 97.3 \text{ mL}
\]

\[
\text{density} = \frac{m}{V}
\]
Example 1-7: Calculate the density of a substance if 742 grams of it occupies 97.3 cm\(^3\).

\[
1 \text{ cm}^3 = 1 \text{ mL} \therefore 97.3 \text{ cm}^3 = 97.3 \text{ mL}
\]

\[
\text{density} = \frac{m}{V}
\]

\[
\text{density} = \frac{742 \text{ g}}{97.3 \text{ mL}}
\]

\[
\text{density} = 7.63 \text{ g/mL}
\]
Density and Specific Gravity

Example 1-8 Suppose you need 125 g of a sucrose an IV. What volume do you need?

- liquid’s density = 1.32 g/mL

You do it!
Example 1-8 Suppose you need 125 g of a sucrose an IV. What volume do you need?

- liquid’s density = 1.32 g/mL

\[
\text{density} = \frac{m}{V} \implies V = \frac{m}{\text{density}}
\]
Example 1-8 Suppose you need 125 g of a sucrose an IV. What volume do you need?

- liquid’s density = 1.32 g/mL

\[
\text{density} = \frac{m}{V} \implies V = \frac{m}{\text{density}}
\]

\[
V = \frac{125 \text{ g}}{1.32 \text{ g/mL}} = 94.7 \text{ mL}
\]
Density and Specific Gravity

\[
\text{Specific Gravity} = \frac{\text{density(\text{substance})}}{\text{density(\text{water})}}
\]

- Water’s density is essentially 1.00 at room T.
- Thus the specific gravity of a substance is very nearly equal to its density.
- Specific gravity has no units.
- Medical Importance—urine normal at 1.00-1.03
  - Too low, kidney problems
  - Too high--dehydrated
Density and Specific Gravity

Example 1-9: A 31.10 gram piece of chromium is dipped into a graduated cylinder that contains 5.00 mL of water. The water level rises to 9.32 mL. What is the specific gravity of chromium?

You do it
Example 1-9: A 31.10 gram piece of chromium is dipped into a graduated cylinder that contains 5.00 mL of water. The water level rises to 9.32 mL. What is the specific gravity of chromium?

Volume of Cr = 9.32 mL - 5.00 mL
= 4.32 mL

density of Cr = \frac{31.10 \text{ g}}{4.32 \text{ mL}}
Example 1-9: A 31.10 gram piece of chromium is dipped into a graduated cylinder that contains 5.00 mL of water. The water level rises to 9.32 mL. What is the specific gravity of chromium?

\[
density\ of\ Cr = \frac{31.10\ g}{4.32\ mL} = 7.19907\ \frac{g}{mL} \approx 7.20\ \frac{g}{mL}
\]

Specific Gravity of Cr = \[
\frac{7.20\ \frac{g}{mL}}{1.00\ \frac{g}{mL}} = 7.20
\]
Heat and Temperature

- Heat and Temperature are not the same thing
- T is a measure of the intensity of heat in a body
- 3 common temperature scales - all use water as a reference
Heat and Temperature

<table>
<thead>
<tr>
<th></th>
<th>MP water</th>
<th>BP water</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fahrenheit</td>
<td>32 °F</td>
<td>212 °F</td>
</tr>
<tr>
<td>Celsius</td>
<td>0.0 °C</td>
<td>100 °C</td>
</tr>
<tr>
<td>Kelvin</td>
<td>273 K</td>
<td>373 K</td>
</tr>
</tbody>
</table>

- **Body temperature** 37.0 °C, 98.6 °F
  - 37.2 °C and greater—sick
  - 41 °C and greater, convulsions
  - <28.5 °C hypothermia
Relationships of the Three Temperature Scales

Kelvin and Centigrade Relationships

\[ K = ^\circ C + 273 \]

or

\[ ^\circ C = K - 273 \]
Relationships of the Three Temperature Scales

Fahrenheit and Centigrade Relationships

\[ ^\circ F = 1.8 \times ^\circ C + 32 \]

or

\[ ^\circ C = \frac{^\circ F - 32}{1.8} \]
Example 1-11: Convert 211°F to degrees Celsius.

\[ ^\circ C = \frac{^\circ F - 32}{1.8} \]

\[ ^\circ C = \frac{211 - 32}{1.8} = 99.4 ^\circ C \]
Example 1-12: Express 548 K in Celsius degrees.

\[ ^\circ C = K - 273 \]

\[ ^\circ C = 548 - 273 \]

\[ ^\circ C = 275 \]