

Module 1: Measurement and Units

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August 22 – 29, 2025

1 Overview

1. **Matter** is anything that has *mass* and takes up space
2. **Units of Measurement** make numbers meaningful
3. **The Metric System** is designed to make units measurement consistent and simple
4. **Unit Conversion** is based on multiplying fractions
5. **Significant Figures** are a convention for maintaining precision in measurements
6. **Scientific Notation** allows us to represent numbers “naturally”
7. **Density** is a measure of how tightly mass is packed into volume

2 The Metric System

The metric system is an attempt to make measurements more standardized and easier to use in mathematical equations than our “English” or “Imperial” system.

- each physical quantity has a *base unit* for that quantity [Wile, 2003, p. 5] (see Table 1, page 5)
- each base unit can be scaled up or down through *metric prefixes* that represent powers of 10 [Wile, 2003, p. 7] (see Table 2, page 5)
- note: don’t make the mistake of thinking that *mass* and *weight* are the same thing¹

¹*Mass* is how much of you there is, *weight* is how hard gravity pulls you down. In a constant gravitational field (*e.g.* most places on earth), weight is a useful proxy for mass, but don’t let’s get them confused.

3 Significant Figures

- we need a way to communicate uncertainty in our measurements
- any number we record contains its own precision claims
- when we record any measurements, we need to be careful of our significant figures

3.1 Rules for Significant Figures

Per our textbook [Wile, 2003, p. 21], a digit is a significant figure if:

1. it is non-zero, or
2. it is a zero between two significant figures, or
3. it is a zero to the right of a decimal point

3.2 Examples

Example 1 2.80 has three significant figures

Example 2 0.0028 has two significant figures

Example 3 28000 has two significant figures

Example 4 5.56 has three significant figures

Example 5 0.44 has two significant figures

3.3 Mathing with Significant Figures

- when adding or subtracting, round your answer to the same precision as your least precise measurement ([Wile, 2003, p. 25])
- when multiplying or dividing, round your answer to the same number of significant figures as your least precise measurement ([Wile, 2003, p. 26])
- digits in fractions don't count: ignore them!
- digits in definitions are considered to have infinite precision, so we can safely ignore them

3.4 Examples

Example 1 What is $5.56mm + 9mm$?

$$\begin{aligned} 5.56mm + 9mm &= 14.56mm \\ &= 15mm \end{aligned} \tag{1}$$

Example 2 What is $2.80cm + 3.6678cm$?

$$\begin{aligned} 2.80cm + 3.6678cm &= 6.4678cm \\ &= 6.47cm \end{aligned} \tag{2}$$

Example 3 What is $17.000ft \times 3.001lb$?

$$\begin{aligned} 17.000ft \times 3.001lb &= 51.017ft \cdot lb \\ &= 51.02ft \cdot lb \end{aligned} \tag{3}$$

Example 4 What is $5.243mol \div 17.32543L$?

$$\begin{aligned} 5.243mol \div 17.32543L &= \frac{5.243mol}{17.32543L} \\ &= 0.30261875 \frac{mol}{L} \\ &= 0.3026 \frac{mol}{L} \end{aligned} \tag{4}$$

4 Scientific Notation

- we can represent numbers as a number $\{x|1 \leq x < 10\}$ multiplied by a power of 10
- this is particularly helpful with very large or very small numbers.
- the rules for scientific notation are:
 1. the first number is between 1 and 10
 2. the power of 10 is the number of places you move the decimal on the original number to the *left* (e.g. 1000 is 10^3)²

²“Left” and “right” here get tricky: we’re counting spaces we need to move the decimal *of the original number* in order to get something between 1 (inclusive) and 10 (exclusive).

3. if you move the decimal to the *right*, then the power of 10 is negative (*e.g.* 0.01 is 10^{-2})
- notice that it's very easy to track significant figures with scientific notation!
 - notice that the metric system is just scientific notation with pretentious names! (see Table 3, page 5)

4.1 Examples

Example 1 we can write 1000 as 1×10^3

Example 2 we can write 256 as 2.56×10^2

Example 3 we can write 0.000002341 as 2.341×10^{-6}

5 Unit Conversion

- we use the idea of fraction multiplication to convert measurements between units.
- **we can multiply any number by 1 without changing it!**

5.1 Examples

Example 1 How many yards in a mile?

We begin with what we know:

- $1\text{mile} = 5280\text{ft}$
- $1\text{yd} = 3\text{ft}$

Metric Base Unit	Physical Quantity	Approximate Size
meter	length	just over 1 yard (39.37 inches)
gram	mass	one penny is about 2 grams
liter	volume	just over 1 quart
second	time	about 1 second

Table 1: Common metric base units

Metric Prefix	Power of 10
<i>giga</i> (G)	1,000,000,000
<i>mega</i> (M)	1,000,000
<i>kilo</i> (k)	1,000
<i>hecta</i> (H)	100
<i>deca</i> (D)	10
<i>deci</i> (d)	0.1
<i>centi</i> (c)	0.01
<i>milli</i> (m)	0.001
<i>micro</i> (μ)	0.000001
<i>nano</i> (n)	0.000000001

Table 2: Most common metric prefixes

Pretentious Name	Scientific Notation Equivalent
<i>giga</i> (G)	$\times 10^9$
<i>mega</i> (M)	$\times 10^6$
<i>kilo</i> (k)	$\times 10^3$
<i>hecta</i> (H)	$\times 10^2$
<i>deca</i> (D)	$\times 10^1$
<i>deci</i> (d)	$\times 10^{-1}$
<i>centi</i> (c)	$\times 10^{-2}$
<i>milli</i> (m)	$\times 10^{-3}$
<i>micro</i> (μ)	$\times 10^{-6}$
<i>nano</i> (n)	$\times 10^{-9}$

Table 3: Scientific Notation and Metric Prefixes

$$\begin{aligned}
 1\text{mile} &= \left(\frac{1\text{mile}}{1}\right)\left(\frac{5280\text{ft}}{1\text{mile}}\right)\left(\frac{1\text{yd}}{3\text{ft}}\right) \\
 &= \frac{(\cancel{1\text{mile}})(5280\cancel{\text{ft}})(1\text{yd})}{(\cancel{1\text{mile}})(3\cancel{\text{ft}})} \\
 &= \frac{5280\text{yd}}{3} \\
 &= \frac{5280}{3}\text{yd} \\
 &= 1760\text{yd}
 \end{aligned}
 \tag{5}$$

Example 2 How many cups are in 5 liters?

We begin with what we know:

- $1\text{quart} = 2\text{pints}$
- $1\text{pint} = 2\text{cups}$
- $1\text{quart} = 0.946353\text{L}$

$$\begin{aligned}
5L &= \left(\frac{5L}{1}\right)\left(\frac{1qt}{0.946353L}\right)\left(\frac{2pint}{1qt}\right)\left(\frac{2cup}{1pint}\right) \\
&= \left(\frac{5\cancel{L}}{1}\right)\left(\frac{1\cancel{qt}}{0.946353\cancel{L}}\right)\left(\frac{2\cancel{pint}}{1\cancel{qt}}\right)\left(\frac{2cup}{1\cancel{pint}}\right) \\
&= \frac{5 \cdot 2 \cdot 2cup}{0.946353} \\
&= \frac{5 \cdot 4 \cdot 2}{0.946353}cup \\
&= \frac{20}{0.946353}cup \\
&= 21.1338cup
\end{aligned}
\tag{6}$$

6 Density

1. density is a measurement of how tightly packed matter is in an object
2. it's related to both *mass* (how much matter there is) and *volume* (how much space something takes up)
3. when you pick up an object and think, "it didn't look that heavy," you're thinking in terms of density
4. lead (*Pb*) is more dense than aluminum (*Al*), so 1g of *Pb* is much smaller ("takes up less space") than 1g of *Al*
5. we calculate density as $\rho = \frac{m}{V}$ where ρ is density, m is mass, and V is volume
6. ρ (density) can be measured in $\frac{kg}{L}$, or $\frac{kg}{m^3}$, or $\frac{g}{mL}$, or $\frac{g}{cm^3}$, or ...

6.1 Examples

Example 1 1.00kg of Carbon occupies 0.4411L. What is the density of Carbon in $\frac{kg}{L}$?

$$\begin{aligned}
 \rho &= \frac{m}{V} \\
 &= \frac{1kg}{0.4411L} \\
 &= 2.26705962 \frac{kg}{L} \\
 &= 2.27 \frac{kg}{L}
 \end{aligned}
 \tag{7}$$

Example 2 1.260kg of Copper occupies 0.140625L. What is the density of Copper in $\frac{kg}{L}$?

$$\begin{aligned}
 \rho &= \frac{m}{V} \\
 &= \frac{1.260kg}{0.140625L} \\
 &= 8.96 \frac{kg}{L} \\
 &= 8.960 \frac{kg}{L}
 \end{aligned}
 \tag{8}$$

Example 3 The density of Silicon is 1.00g per 0.429mL. What is the mass of 3.24L of Silicon?

$$\begin{aligned}
 \rho &= \frac{m}{V} \\
 (\rho)(V) &= \left(\frac{m}{V}\right)(V) \\
 (\rho)(V) &= \left(\frac{m}{\cancel{V}}\right)(\cancel{V}) \\
 (\rho)(V) &= m \\
 m &= (\rho)(V) \\
 &= \left(\frac{1.00g}{0.429mL}\right)(3.24L) \\
 &= \left(\frac{1.00g}{0.429mL}\right)\left(\frac{3.24L}{1}\right)\left(\frac{1000mL}{1L}\right) \\
 &= \left(\frac{1.00g}{0.429\cancel{mL}}\right)\left(\frac{3.24\cancel{L}}{1}\right)\left(\frac{1000\cancel{mL}}{1\cancel{L}}\right) \\
 &= 7552.447552g \\
 &= 7550g
 \end{aligned}
 \tag{9}$$

7 Homework

Review Problems: p. 35 # 1–10

Practice Problems: p. 36 # 1–10 (due 2025-09-05)

References

[Wile, 2003] Wile, J. L. (2003). *Exploring Creation with Chemistry*. Apologia Educational Ministries, Inc., 2 edition.