DATA ANALYSIS Using the Metric System

Overall Learning Goal  Students will be able to construct and interpret graphs based on lab data recorded in metric units to the appropriate degree of accuracy.

Units of Measurement System Internationale’ (SI) Metric System

Prefixes smaller than 1

\[
\begin{array}{|c|c|c|}
\hline
\text{Prefix} & \text{Symbol} & \text{Factor} \\
\hline
deci & d & 10^{-1} \\
centi & c & 10^{-2} \\
milli & m & 10^{-3} \\
\mu & \mu & 10^{-6} \\
nano & n & 10^{-9} \\
pico & p & 10^{-12} \\
\hline
\end{array}
\]

0.1
0.01
0.001
0.000 001
0.000 000 001
0.000 000 000 001

Prefixes larger than 1

\[
\begin{array}{|c|c|c|}
\hline
\text{Prefix} & \text{Symbol} & \text{Factor} \\
\hline
\text{kilo} & k & 10^3 \\
\text{mega} & M & 10^6 \\
\hline
1000 \\
1000 000 \\
\hline
\end{array}
\]

Volume is not a base unit. Volume units are derived from length.

A cube can hold liquid. It has to be 3 dimensional. A cube is described using length, width, and height.

1 liter is a cubic decimeter.\( \text{dm}^3 \)

1ml is a cubic centimeter \( \text{cm}^3 \) or cc

Converting from one unit to another using dimensional analysis

Write the relationship between the units.

Example, \( 1 \text{ liter} = \frac{1000 \text{ milliliters}}{1 \text{ liter}} \)

Any relationship can be written in two ways. These are called conversion factors.

\[
\begin{array}{c}
1 \text{ liter} \\
1000 \text{ milliliters} \\
\hline
1000 \text{ milliliters} \\
1 \text{ liter} \\
\hline
\end{array}
\]

What conversion factor would you use to convert 4.2 liters to milliliters?

\[
4.2 \text{ liters} \times \frac{1000 \text{ ml}}{1 \text{ liter}} = 4200 \text{ ml}
\]

Write the conversion factors between meters and centimeters.

\[
\begin{array}{c}
1 \text{ m} \quad \text{or} \\
100 \text{ cm} \\
\hline
100 \text{ cm} \\
1 \text{ m} \\
\hline
\end{array}
\]

Which one would you use to convert 350 cm to meters?

\[
350 \text{ cm} \times \frac{1 \text{ m}}{100 \text{ cm}} = 3.5 \text{ m}
\]

Temperature scales Three scales: Fahrenheit, Celsius, Kelvin

\[
\begin{array}{c}
\text{Celsius} = \text{Kelvin} - 273 \\
\text{Kelvin} = \text{Celsius} + 273 \\
\end{array}
\]

Precision and Accuracy

Accurate close to true (accepted) value
Precise close together, not necessarily close to true value

Percent Error

How accurate (or inaccurate) is a measurement? Calculate the % error.

\[
\text{Accepted value} - \text{Experimental value} \times 100 = \% \text{ error}
\]

Accepted value
A student’s conducts an experiment which should produce 0.10 grams of solid copper from a copper chloride solution. The final mass of copper is 0.08 grams. Calculate the percent error.

\[
\frac{0.10 - 0.08}{0.10} \times 100 = \% \text{ error}
\]

**Scientific Notation Rules**

1. The **coefficient** must be greater than or equal to 1 and less than 10.
2. The **base** must be 10.
3. The **exponent** must show the number of decimal places that the decimal needs to be moved to change the number to standard notation.

Examples:

- \(3.8 \times 10^{-8}\) Numbers > 1 have a positive exponent
- \(3.8 \times 10^{-8}\) Numbers < 1 have a negative exponent

**Significant Figures**

There are 2 different types of numbers: exact and measured.

- **Measured numbers** are obtained with a measuring device so these numbers have ERROR.
- When you use your calculator your answer can only be as accurate as your worst measurement...

- **Exact Numbers** are obtained when you count objects or use a defined relationship.

**Measurement and Significant Figures**: Every experimental measurement has a degree of uncertainty.

The last digit in any measurement is always an estimate – it is not certain.

To indicate the precision of a measurement, the value recorded should use all the digits known with certainty and the last digit which is estimated. These measured and estimated digits are the significant figures of any measurement.

**Rules for significant figures** When reading a measured value, all nonzero digits should be counted as significant.

**Rules for Zeros**

1. Non-zero digits are always significant.
2. Any zeros between two significant digits are significant.
3. A final zero or trailing zeros are significant **ONLY** if they are in the decimal portion.

**The Box-and-Dot Method: A Simple Strategy for Counting Significant Figures.**

In this method a “dot” is the decimal point.

Step 1 - Draw a box around all nonzero digits, beginning with the leftmost nonzero digit and ending with the rightmost nonzero digit in the number, including any sandwiched zeros.

- \(0.001230400\)

Step 2 - If a dot is present, draw a box around any trailing zeros.

- \(0.00\overline{1230400}\)

Step 3 - All the boxed digits are significant. There are 7 significant digits.
**Multiplying and dividing with significant figures**

Inspect the numbers in the calculation; find which has the least number of significant figures. This is the number of significant figures you should have in your answer.

#1 Find the area of a rectangle 2.1 cm by 3.24 cm.
Solution: Area = 2.1 cm \times 3.24 cm = 6.804 cm²
2.1 has two significant figures, 3.24 has 3. The answer should have two significant figures. Answer is 6.8 cm.

**Adding and Subtracting with Significant Figures**

Write down the numbers, keep like decimal places under each other, and add (or subtract).

Next, find which column contains the first estimated figure. This column determines the last place of the answer. Round the answer off in this column.

#2

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>42.56 g</td>
</tr>
<tr>
<td>39.460 g</td>
<td>4.1 was estimated in the first decimal place; answer is rounded to 86.1 g</td>
</tr>
<tr>
<td>4.1 g</td>
<td></td>
</tr>
<tr>
<td>86.120 g</td>
<td></td>
</tr>
</tbody>
</table>

**Scientific Method** How do we apply knowledge of chemistry to problems?

**Steps in the scientific method**

Observation - a problem that you want to solve
Hypothesis - a possible solution to a problem, based on knowledge and research.
Experiment - the part of the scientific method that tests your hypothesis.
Data & Analysis - information collected during the experiment
Conclusion - a summary of the experiment's results, and how those results match up to your hypothesis.

**Variables**

- independent variable - the one you change in an experiment.
- dependent variable - changes depend on the changes in the independent variable.
- controls - kept the same throughout the experiment

**Qualitative data**. Words

**Quantitative data**. Numbers

**Law or Theory?**

A scientific law is a statement of fact that describes an action or set of actions. It is generally accepted to be true in any situation (universal).

A scientific theory is NOT a guess. A scientific theory is a well substantiated explanation for a set of observations. A scientific theory is the result of many observations and experiments by many scientists. It can be used to make testable predictions.