CHAPTER 2
Data Analysis

2.1 Units of Measurement

- The standard of measurement used in science are those of the metric system.
- All the units are based on 10 or multiples of 10.
- **SI Units**: The International System of Units is a revised version of the metric system.
- There are seven base SI units.

SI Base Units

Quantity Measured	Unit (symbol)
Length	meter (m)
Mass	kilogram (kg)
Time	second (s)
Electric current	ampere (A)
Temperature	Kelvin (K)
Substance amount	mole (mol)
Light intensity	candela (cd)

Metric Prefixes used with SI Units

Prefix	*Symbol*	*Multiple*	*Standard*
giga	G	1 000 000 000	10^{9}
mega	M	1 000 000	10^{6}
kilo	k	1 000	10^{3}
hecto	h	100	10^{2}
deca	da	10	10^{1}
deci	d	0.1	10^{-1}
centi	c	0.01	10^{-2}
milli	m	0.001	10^{-3}
micro	μ	0.000001	10^{-6}
nano	n	0.000000001	10^{-9}

Greater than 1

Less than 1

Derived Units

- These are not measured directly.
- They are a combination of base units
- Examples: volume, density, pressure, speed

How is volume measured?

- The space occupied by any sample of matter is called its **volume.**
- The SI unit of volume is the amount of space occupied by a cube that is 1 meter along each edge. **1 m^3**
- A more convenient unit of volume for everyday use is the liter (L)
  - (1 dm^3 = 1 L) (1 cm^3 = 1 mL)
- The units milliliter and cubic centimeter are used interchangeably.
- The volume of any solid, liquid, or gas will change with temperature.

How is mass measured?

- Mass is the quantity of matter an object contains.
- Mass is not weight, weight is a force. *(Gravity required)*
- The weight of an object can change with its location, although its mass remains constant regardless of its location. (On the moon vs. Earth)
- The mass of an object is measured by comparing it to a standard mass of 1 **kilogram (kg),** the basic SI unit of mass.
- A **gram** is 1/1000 of a kilogram and is a more commonly used unit of mass because a kilogram is relatively large.

Measuring Temperature

- Almost all substances expand with an increase in temperature.
- Most substances also contract as the temperature decreases.
- Heat moves from the object at the higher temperature to the object at the lower temperature.
- Temperature is the degree of hotness or coldness of an object.

Measuring Temperature

- Several temperature scales have been devised. Two readily determined temperatures are used as reference temperature values. 1.) freezing point and the 2.) boiling point of water,
- On the **Celsius temperature** scale, the freezing point of water is taken as 0 °C and the boiling point of water as 100 °C.
- On the **Fahrenheit temperature** scale, the freezing point of water is taken as 32 °F and the boiling point of water as 212 °F.

Kelvin Scale

- Another temperature scale used in the physical sciences is the Kelvin scale, or absolute scale.
- On the Kelvin temperature scale, the freezing point of water is 273 K and the boiling point is 373 K.
- Notice that with the Kelvin scale, the degree sign is not used.
- The zero point on the Kelvin scale, or **absolute Zero,** is -273°C.
- K = °C + 273 or °C= K-273

Temperature Scales

Fahrenheit, Celsius, Kelvin

- Kelvin = Celsius + 273
- Celsius = (°F - 32) x 0.56
- Fahrenheit = (1.8 x °C) + 32
- Celsius = Kelvin - 273

Fahrenheit	Celsius	Kelvin
76°F		
	45°C	
		315 K
	-16°C	
-145°F		

1. Which is warmer, 0° Celsius or 0° Fahrenheit?
- *To determine the answer, you must convert one of them so both are the same temperature scale and then compare!!!*

2. Which is colder, 223 Kelvin or -50° Fahrenheit?

-50°C -46°C

223 K - 273 = -50 °C (-50°F - 32) x 0.56 = -46 °C

Units of Energy

Energy is the capacity to do work or to produce heat. SI unit is the joule.

1 Joule =0.2390 calorie

1 calorie = 4.184 Joule

DENSITY

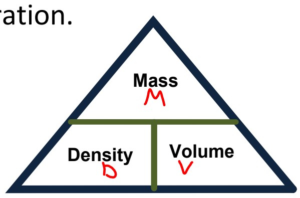
- The ratio of an object's mass to its volume

Density = mass ÷ volume
- Answer expressed in g/ml or g/cm^3.

Mass = density x volume
- Answer expressed in grams, kg, etc.

Volume = mass ÷ density
- Answer expressed in ml, cm^3, L, etc.

Cover what you are looking for and perform the operation.



D = M = 312.0 - g V = 1.5 - mL	D = 3.21 g/mL M = V = 2.81 - L	D = 1.89 g/mL M = 321.2 - g V =
D =	**M =**	**V =**

How does mass & volume change density?

Mass stays the same:
- V-increase, D-decrease
- V-decrease, D-increase

Volume stays the same:
- M-increases, D-increases
- M-decreases. D-decreases

When do substances sink in water?

Liquids

Water = 1.0 g/mL *(you must memorize this value)*

Sub A = 1.7 g/mL

Sub B = 0.5 g/mL

Sub C = 4.6 g/mL

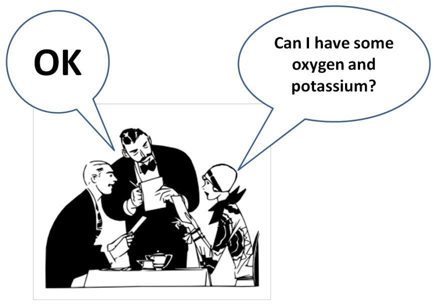
SeaWater = 1.025 g/mL

http://www.youtube.com/watch?v=_Ww6BIy3nc0

http://www.bing.com/videos/search?q=youtube+density&FORM=VIRE3#view=detail&mid=227066B76C55BC2C7183227066B76C55BC2C7183

How does Temperature affect density?

- The volume of a substance usually increases as the temperature is increased, though its mass remains the same.
- Thus, the density of a substance usually decreases as its temperature increases.
- Water is an important exception. Below 4ºC the volume of water increases as its temperature decreases. Ice floats because it is less dense than water. *(H Bonding)*



Determining Error

- In order to evaluate the accuracy of a measurement, you must be able to compare it to the true or accepted value.
- The accepted value is the true or correct value based on reliable references.
- The experimental value is the measured value determined in the experiment in the laboratory.
- **Error** = accepted value - experimental value

Percent Error

The error, divided by the accepted value, expressed as a percentage of the accepted value.

- $\% \text{ error} = \left| \dfrac{\text{Accepted - Experimental}}{\text{Accepted value}} \right| \times 100$

Error vs. Percent Error

- Error can be positive or negative, but it does not matter.
- Percent error will always be a positive value since the absolute value of the error is used.
- Percent error is the more valuable calculation, since the size of the error is relative to what you are measuring.
- An error of 10.0 is bad if the actual measurement is 2.0, however it is good if the actual measurement is 2,55.0

Example Problems

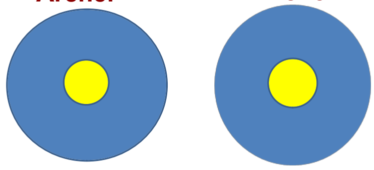
Problem 1	Problem 2
Accepted Value = 137.7-mL	Accepted Value = 7.14 g/mL
Experimental Value = 131.2-mL	Experimental Value = 9.27 g/mL
Error =	Error =
% Error =	% Error =

2.3 Reliability of Measurement

- **Accuracy** is how close a measurement comes to the actual dimension or true value of whatever is measured.
  - More decimal places usually means more accurate
- **Precision** is how close several measurements are to one another or consistent measurements.
  - They don't have to be accurate.

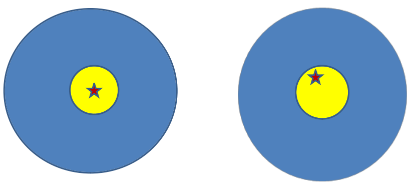
Which archer do you want to shoot the apple off of your head?

Rookie Archer | Pro Archer

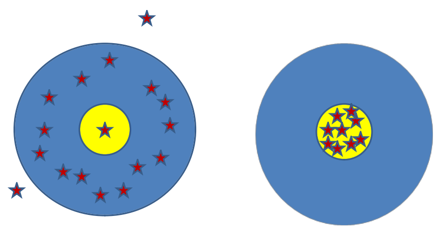


1 shot: archer shooting at the round target is more accurate

Rookie Archer | Pro Archer



After multiple shots, which archer is more precise?



TQ Accuracy vs. Precision

Which group of measurements is most accurate and which is the most precise, when the actual mass of an item is 1.22 grams?

A. 1.10, 1.20, 1.30

B. 1.29, 1.30, 1.31

C. 1.05, 1.68, 1.87

Accuracy vs. Precision

- Accuracy depends on the measuring device and can be no more reliable than the measuring instrument.
- Precision depends on the skill of the person doing the measuring.
- Regardless of the quality of the measuring device, precise measurements will be similar since precise measurements are reproducible but necessarily accurate.

Calibration

- Measuring devices need to be calibrated periodically to make sure they are maintaining accuracy.
- Even when calibrated, a poor device may only be accurate once.
- This does not increase precision or the reliability of the instrument!
- Thermometers, barometers, voltage meters, scales

SIGNIFICANT FIGURES IN MEASUREMENTS

- The significant figures in a measurement include all the digits that are known precisely plus one last digit that is estimated.
- Estimated digit is rounded.

Rules for Significant Digits

1. All numbers 1 – 9 are always significant.
2. Zeroes between significant digits are always significant.
3. Zeroes after significant digits are only significant if they follow a decimal point.
4. Zeroes in front of significant digits are never significant, even after a decimal point.
5. When using scientific notation, all numbers (including zeroes) in the coefficient are significant.

Rules for determining significant digits:

Every nonzero digit is significant.

- 24.7
- 34,000
- 500,000

When does the decimal point matter?

- The decimal point only matters if it follows a nonzero digit.
- Zeros following a nonzero and to the right of a decimal point are always significant.
  - 247.00
  - 0.00000500
- Zeros appearing in front of nonzero digits are not significant, even if they follow a decimal point.
  - 0.00247

When zeroes are between significant digits?

- Zeroes are always significant if they are between two significant digits:
  - 303
  - 2450.1
  - 650.0
  - 0.0000501

What about Scientific Notation?

- When a number is in scientific notation, all numbers in the coefficient are considered significant.
  - 2.470×10^{3} (2. 470 is the coefficient)
  - 3.003×10^{8} (3.003 is the coefficient)
  - 2.3×10^{-6} (2. 3 is the coefficient)

Determine the number of Significant Digits

200	200.0	0.0002
0.0000020	301	301,000
500,000	3.460×10^{4}	0.208
405,450	0.0000601	4030
0.310	3.09×10^{-7}	1.0000

Round to 3 Significant Digits

413,670	
0.034678	
6,908,595	
0.00000567908	
45	

When are significant figures not important?

There are two instances in which measurements have an unlimited number of significant figures.

1. Counting
2. Conversion factors

Significant Figures in Calculations

- An answer cannot be more precise than the least precise measurement from which it was calculated.
- So, we round calculations to reflect the least precise measurement. *See the rounding rules...*
- A calculator does not keep track of significant figures or round off answers correctly.
- Rounding does not change the value of your answer, it just makes it less precise by decreasing significant digits.

Addition and Subtraction

- The answer to an addition or subtraction problem can have no more digits to the right of the decimal than the number in the problem with the least numbers past the decimal.

Addition: Places after decimal

4.1 + 5.68 + 6.4370192

2.3333 + 1.12

Subtraction: Places after decimal

18 – 4.25891

154.1 – 8.555555

Multiplication and Division

- The answer to a multiplication and division problem can have no more significant digits than the number in the problem with the least number of significant digits.

Multiplication: total significant digits

3.555555 x 12.3

4.6 x 8.375 x 3300

Division: total significant digits

55.6 ÷ 2.68754

55500 ÷ 231.078

2.2 Scientific Notation

- A number is written as the product of two numbers:
  - *a coefficient and a power of 10*.
- Only significant digits are included.
  - Ex. 36,000 is written as **3.6 x 10^4**
- The coefficient is a number equal to or greater than one and less than ten. **$1.000\overline{0}$ - $9.9999\overline{9}$**

- Numbers **greater than ten** in scientific notation the **exponent is positive** and equal to the number of places that the decimal point has been moved to the left.

  4.3 x 10^8

- Numbers **less than one** have a **negative exponent** when you write them in scientific notation.

  3.17 x 10^{-4}

Rules for Scientific Notation

- The coefficient must be between 1.0 and 9.99.
- Your coefficient must contain all significant digits.
- Move the decimal point as many places as necessary until you create a coefficient between 1.0 and 9.99.
- The exponent will be the number of places you move your decimal point.
- **Numbers greater than 10 always have exponents that are positive.**
- **Numbers less than 1.0 always have exponents that are negative.**

Put into Scientific Notation

634,000	
0.0000346	
4,908,000	
0.005680	
122	

Put into Expanded Form

3.062×10^{11}	306200000000
2.1×10^{-8}	.000000021
9.00×10^{7}	90000000
5.098×10^{-4}	.0005098

Addition: Places after decimal

$(3.04 \times 10^{22}) + (2.457 \times 10^{15})$

Subtraction: Places after decimal

$(7.146 \times 10^{-7}) - (5.47 \times 10^{-10})$

Multiplication: total sig figs

$(2.5 \times 10^{-3}) \times (3.620 \times 10^{-5})$

Division: total sig figs

$(8.040 \times 10^{11}) \div (4.5 \times 10^{8})$

Review Problem A

$(9.99 \times 10^{12}) \div (3.33 \times 10^{2})$

Review Problem B

$(2.00 \times 10^{16}) \times (2.50 \times 10^{21})$

Review Problem C

$(8.00 \times 10^{8}) + (2.0 \times 10^{12})$

Review Problem D

$4.32 \times 10^{16} \times 7.8917 \times 10^{21} + 2.4445 \times 10^{16}$

CONVERSION PROBLEMS

- A conversion factor is a ratio of equivalent measurements
- Dimensional Analysis
- Converting between units *(measurement does NOT Change)*
- Multi-step problems
- Converting complex units
- Answer should be in scientific notation with the number of significant digits in the measurement.

When converting a measurement...

- The answer should be in scientific notation with the number of significant digits in the actual measurement.

One Step Conversions

- **6.76 L → mL**
- **7.41×10^{8} feet → miles**

Two Step Conversions

- 3.29 km → cm
- 3.2×10^{11} lbs → grams

More Two Step Conversions

- 6.28×10^{9} mm → km
- 4.587×10^{-17} µL → cm^3

Multi Step Conversions

- 2.63 hm → inches
- 6.23×10^{13} minutes → decades

Converting in Multiple Dimensions

- 4.13 g/cm^3 → cg/dm^3
- 5.12×10^{-1} cm/s → km/h

Converting in Multiple Dimensions

A car covers 175.6 miles in 2.5 hours. Calculate it's speed in meters per second?

Using new conversion factors

On your family trip, you traveled 235.5 miles over 3.5 hours? Based on your current speed, how far will you travel in the next 45 minutes?

Making your own conversion factors

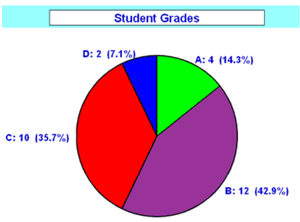
You have a dixie cup that holds 55-mL. How many dixie cups would be needed to fill a 1-kL bottle?

2.4 Representing Data

- Data is typically gathered and placed in tables, but tables are not always easy to read and make it difficult to compare.
- Tables do not make it easy to identify Patterns in the collected data either.
- Graphs make data easier to compare and is more likely going to reveal patterns in the data.

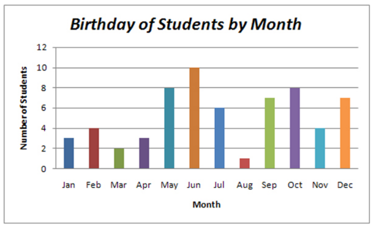
Circle Graphs

Usually used when showing percentages



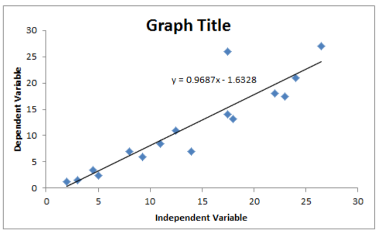
Bar Graphs

Usually used to compare quantities



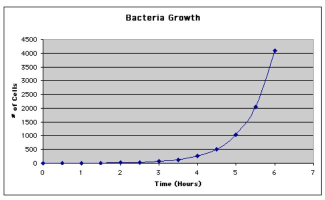
Line Graphs

Most commonly used in chemistry



Parts of a Line Graphs

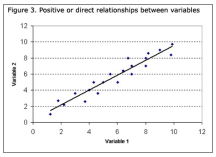
The scientist is testing the dependent variable and has no control over it. (y-axis)



Independent variable: the part of an experiment deliberately changed by the scientist or a constant factor (x-axis)

Line of Best Fit

- Used when points are scattered and a straight line cannot be drawn between all points.
- The line must be drawn so the number of points above the line equal the number of points below the line.



Straight vs. Curved Lines

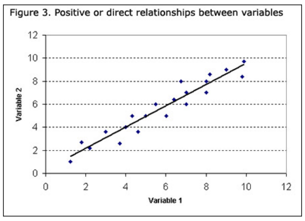
Straight Line: a constant increase or decrease

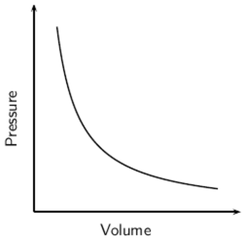
- There is a pattern between variables

Curved Line: not a constant increase or decrease

- There is not a pattern between variables

Straight vs. Curved Lines





Relationships on a Line Graph

Direct Relationship: the independent variable increases, the dependent variable increases.

Inverse Relationship: the independent variable increases, the dependent variable decreases.

Exponential Relationship: as the independent variable increases, the dependent variable increases at a factor greater than 1.

Positive vs. Negative Slope

Positive Slope: The line rises to the right.

- Dependent variable increases as the independent variable increases.

Negative Slope: The line sinks to the right.

- Dependent variable decreases as the independent variable increases.

Positive vs. Negative Slope

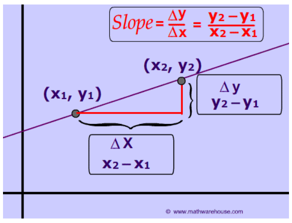
Positive Correlation

Negative Correlation

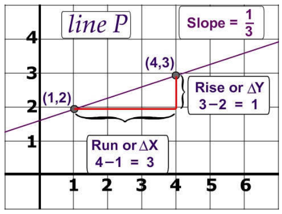
Calculating the Slope of Line

Show the relationship between the variables using the steepness, or slope, of the line.

$$m = \frac{y_2 - y_1}{x_2 - x_1}$$



Calculating the Slope of Line



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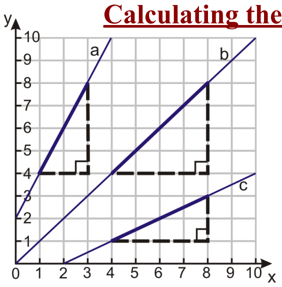

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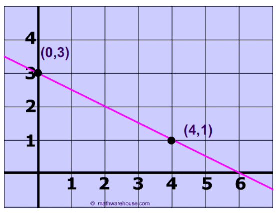

Ch. 2 Test

40 Multiple Choice:

- Temperature Conversions and comparing different temperature scales
- Determining Significant Digits
- Calculations with scientific notation & significant digits
- Scientific notation to decimal, and decimal to scientific notation
- Accuracy vs. Precision
- SI Units
- Compare prefixes of the Metric System
- What is a derived measurement compared to a direct measurement?
- Dimensional Analysis and Conversions
- How do volume changes effect density?
- Calculations with Density, Volume, and Mass
- Accepted Value vs. Experimental Value
- Percent Error vs. Error and how do determine each
- Types of graphs and variable relationships