1 • Introduction
The Scientific Method
(1 of 20)

This is an attempt to state how scientists do science. It is necessarily artificial. Here are MY five steps:

- Make observations
  the leaves on my plant are turning yellow
- State a Problem to be solved
  how can I get my plants healthy (non-yellow)
- Form a hypothesis
  maybe they need more water
- Conduct a controlled experiment
  water plants TWICE a week instead of once a week
- Evaluate results
  if it works, good... if not, new hypothesis (sunlight?)

1 • Introduction
Observations and Measurements
Qualitative, Quantitative, Inferences
(2 of 20)

Step 1 of the Scientific Method is Make Observations. These can be of general **physical** properties (color, smell, hardness, etc.) which are called **qualitative** observations.

These can be **measurements** which are called **quantitative** observations.

There are also statements that we commonly make based on observations. *"This beaker contains water"* is an example. You **infer** (probably correctly) it is water because it is a clear, colorless liquid that came from the tap. The **observations** are that it is clear, it is colorless, it is a liquid, and it came from the tap.

1 • Introduction
Significant Digits I
What do they mean?
(3 of 20)

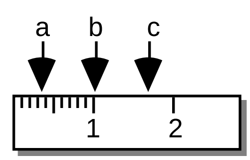
Consider: 16.82394 cm
In a measurement or a calculation, it is important to know which **digits** of the reported number are **significant**.

That means… if the same measurement were repeated again and again, some of the numbers would be **consistent** and some would simply be **artifacts**.

All of the digits that you are absolutely certain of plus one more that is a judgment are significant.

If all the digits are significant above, everyone who measures the object will determine that it is 16.8239 cm, but some will say …94 cm while others might say …95 cm.

1 • Introduction
Significant Digits II
Some examples with rulers.
(4 of 20)



(A composite ruler)

a- No one should argue that the measurement is between 0.3 and 0.4. Is it exactly halfway between (.35 cm)… or a little to the left (.34 cm)? The last digit is the judgment of the person making the measurement. The measurement has 2 significant digits.
b- The same ruler… so the measurement still goes to the hundredths place… 1.00 cm (3 significant digits).
c- A ruler with fewer marks reads 1.6 cm (2 sig digits).

1 • Introduction
Significant Digits III
Rules for Recognizing Sig. Digits
(5 of 20)

In a number written with the correct number of sig. digits...
- All non-zero digits are significant. 523 grams (3)
- 0's in the MIDDLE of a number are ALWAYS significant.
  5082 meters (4) 0.002008 L (4)
- 0's in the FRONT of a number are NEVER significant.
  0.0032 kg (2) 0.00000751 m (3)
- 0's at the END of a number are SOMETIMES significant.
  - Decimal point is PRESENT, 0's ARE significant
    2.000 Liters (4) 0.000500 grams (3)
  - Decimal point is ABSENT, 0's are NOT significant
    2000 Liters (1) 550 m (2)

NOTE: textbook values are assumed to have all sig. digits

1 • Introduction
Scientific Notation
Useful for showing Significant Digits
(6 of 20)

Scientific notation uses a number between 1 and 9.99 x 10 to some power. It's use stems from the use of slide rules.

Know how to put numbers into scientific notation:

$5392 = 5.392 \times 10^3$ $0.000328 = 3.28 \times 10^{-4}$

$1.03 = 1.03$ $550 = 5.5 \times 10^2$

Some 0's in numbers are placeholders and are not a significant part of the measurement so they disappear when written in sci. notation. Ex: 0.000328 above. In scientific notation, only the three sig. digits (3.28) are written.

Scientific Notation can be used to show more sig. digits. Values like 550 (2 sig. digits) can be written 5.50×10^2 (3)

1 • Introduction
Significant Digits IV
Significant Digits in Calculations
(7 of 20)

When you perform a calculation using measurements, often the calculator gives you an incorrect number of significant digits. Here are the rules to follow to report your answers:

x and ÷: The answer has the same # of sig. digits as the number in the problem with the least number of sig. digits.
example: 3.7 cm x 8.1 cm = 29.97 ≈ 30. cm^2 (2 sig. digits)

+ and –: The last sig. digit in the answer is the largest uncertain digit in the values used in the problem.
example: 3.7 cm + 8.1 cm = 11.8 cm (3 sig. digits)

Know **how** to **ilustrate why** these rules work.

1 • Introduction
Accuracy vs. Precision
(8 of 20)

Accuracy refers to how close a measurement is to some **accepted** or **true** value (a **standard**).

Ex: an experimental value of the density of Al° is 2.69 g/mL. The accepted value is 2.70 g/mL. Your value is accurate to within 0.37%

% error is used to express accuracy.

Precision refers to the **reliability**, **repeatability**, or **consistency** of a measurement.

Ex: A value of 2.69 g/mL means that if you repeat the measurement, you will get values that agree to the tenths place (2.68, 2.70, 2.71, etc.)

± and sig. digits are used to express precision.

1 • Introduction
Metric System
(9 of 20)

We generally use three types of measurements:

volume	Liters	(mL)
length	meters	(km, cm and mm)
mass	grams	(kg and mg)

We commonly use the prefixes:

centi-	$^1/_{100}$th
milli-	$^1/_{1000}$th
kilo-	1000

Occasionally you will encounter micro(μ), nano, pico, mega, and giga. You should know where to find these in chapter 1. Know that 2.54 cm = 1 inch and 2.20 lb = 1 kg

1 • Introduction
% and ppm
(10 of 20)

Percentage is a mathematical tool to help compare values. Two fractions, 3/17 and 5/31 are difficult to compare:
If we set up ratios so we can have a common denominator:

$$\frac{3}{17} = \frac{x}{100} = \frac{17.65}{100} \qquad \frac{5}{31} = \frac{x}{100} = \frac{16.13}{100}$$

so… we can see that $\frac{3}{17} > \frac{5}{31}$.

There are 17.65 **parts per 100** (Latin: parts ***per centum***) or 17.65 **percent** (17.65 %)… the **%** is a **"1 0 0"**

ppm (parts per million) is the same idea, (use 1,000,000 instead of 100) $\frac{3}{17} = \frac{x}{1\,000\,000} = 176{,}470$ ppm

1 • Introduction
Unit Analysis
Converting between English and Metric Units
(11 of 20)

Consider the metric/English math fact: 2.54 cm = 1 inch
This can be used as the "conversion factor":

$$\frac{2.54\text{ cm}}{1\text{ inch}} \quad \text{or} \quad \frac{1\text{ inch}}{2.54\text{ cm}}$$

You can convert 25.5 inches to cm in the following way:
Given: 25.5 in
Desired: ? cm $\quad 25.5\text{ in} \times \frac{2.54\text{ cm}}{1\text{ in}} = 64.77\text{ cm} \approx 64.8\text{ cm}$

This is the required way to show your work. You have **two** jobs in this class, to be able to **perform** the conversions and to be able to **prove** that you know why the answer is correct.

1 • Introduction
Temperature Scales
(12 of 20)

The important idea is that temperature is **really a measure of something**, the average **motion** (**kinetic energy**, **KE**) of the molecules.

Does **0°C** really mean **0 KE**? **nope**... it simply means the freezing point of water, a **convenient** standard.

We have to cool things down to –273.15°C before we reach 0 KE. This is called 0 Kelvin (0 K, note: NO ° symbol.)

For phenomena that are proportional to the KE of the particles (pressure of a gas, etc.) you must use temperatures in K. **K = °C + 273 °C = K – 273**

1 • Introduction
Mass vs. Weight
Theory, Measuring, Conversions
(13 of 20)

mass is the **amount** of something...
weight is how much gravity is pulling on the mass.
(Weight will be proportional to the mass at a given spot.)

Mass is what we REALLY want to use... measured in grams. You use a balance to measure mass... you compare your object with objects of known mass.

Weight is measured with a **scale** (like your bathroom scale or the scale at the grocery store). If there is no gravity, it doesn't work. Note: electronic balances are really scales!

You convert mass / weight using: $\frac{1\ \text{kg}}{2.205\ \text{lbs}}$ or $\frac{2.205\ \text{lbs}}{1\ \text{kg}}$

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1 • Introduction
Potential Energy (PE) and Kinetic Energy (KE)
(14 of 20)

You can **calculate** the **KE** of an object: $\mathbf{KE = \frac{1}{2}mv^2}$

m = mass, v = velocity [Note units: $1\ \text{J} = 1\ \text{kg}\cdot\text{m}^2\cdot\text{s}^{-2}$]
Temperature is a measure of the **average** kinetic energy.

PE = the **potential to do work** which is due to an object's **position in a field**. For example, if I hold a book 0.5 m above a student's head it can do some damage... 1.0 m above her/his head, more work can be done.

Important ideas:
Objects tend to change from high PE to low PE (**downhill**).
High PE is **less stable** than low PE.

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1 • Introduction
Mass, Volume, and Density
Intensive vs. Extensive Properties
(15 of 20)

Extensive properties depend on the **amount** of substance. We **measure** these properties frequently... (mass & volume... mostly).

Intensive properties are **independent** of the size of the sample. These are useful for **identifying** substances... (melting point, boiling point, density, etc.)

It is interesting that an intensive property, density = $\frac{\text{mass}}{\text{volume}}$ is the ratio of two extensive properties... the size of the sample sort of "cancels out." Be able to do density problems (3 variables) and know the usefulness of specific gravity.

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1 • Introduction
Calorimetry
(16 of 20)

Heat is the **total** KE while **temperature** is the **average** KE.

A way to measure heat is to measure the temperature change of a substance... often water. It takes 1 calorie of heat energy (or 4.184 J) to heat 1 gram of H_2O by 1 °C.

The specific heat of water = $1\ \frac{\text{cal}}{\text{g}\cdot°\text{C}} = 4.184\ \frac{\text{J}}{\text{g}\cdot°\text{C}}$

$$\mathbf{heat = specific\ heat \times mass\ H_2O \times \Delta T}$$

You can heat other substances as well, you just need to know their specific heats. Notice that this is simply heating or cooling a substance, not changing its phase.

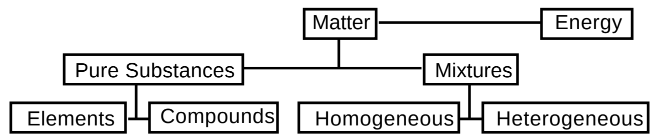
1 • Introduction
Physical and Chemical Properties
Physical and Chemical Changes
(17 of 20)

Equations to symbolize changes: **reactants → products**

Physical Properties can be measured from a sample of the substance **alone**... (density, MP, BP, color, etc.)
Chemical Properties are measured when a sample is **mixed with another chemical** (reaction with acid, how does it burn in O_2)

Physical Changes imply that no new substances are being formed (melting, boiling, dissolving, etc.)
Chemical Changes imply the substance is forming new substances. This change is accompanied by heat, light, gas formation, color changes, etc.

1 • Introduction
Pure Substances, Elements, & Compounds
Homogeneous & Heterogeneous Mixtures
(18 of 20)



This chart should help you sort out these similar terms.
Be able to use chemical symbols to represent elements and compounds. For example...
$CuSO_4 \cdot 5H_2O$, a hydrate, contains 21 atoms & 4 elements.

Memorize the 7 elements that exist in diatomic molecules: HONClBrIF or BrINClHOF or “H and the 6 that make a 7 starting with element #7”

1 • Introduction
Separating Mixtures by Filtration, Distillation, and Chromatography
(19 of 20)

Mixtures are substances the are NOT chemically combined... so if you want to separate them, you need to **exploit** differences in their **PHYSICAL** properties.

Filtration:
some components of the mixture dissolve and some do not. The filtrate is what passes through the filter.

Distillation:
some components vaporize at different temperatures or one component may not vaporize at all (e.g.: salt+water) complete separation may not be possible.

Chromatography:
differences in solubility vs. adhesion to the substrate. Substaratemay be filter paper (paper chromatography),

1 • Introduction
Early Laws: the Law of Definite Composition & the Law of Simple Multiple Proportions
(20 of 20)

Definite Composition:
samples of the same substance from various sources (e.g. water) can be broken down to give the same %’s of elements. *Calculation: percent composition*

Multiple Proportions:
samples of 2 substances made of the same 2 elements... (e.g. CO_2 & CO or H_2O and H_2O_2 or CH_4 and C_3H_8) if you break down each to give equal masses of one element, the masses of the other element will be in a simple, whole-number ratio.
Calculation: proportions to get equal amounts of one element and then simple ratios.