Chapter 1

Chemical Foundations

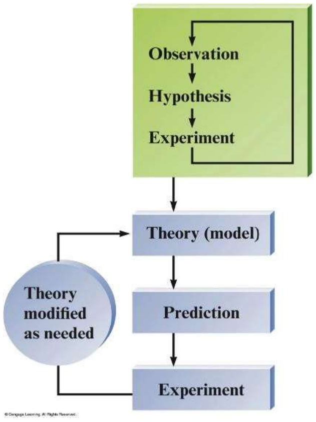
Chemistry is the study of **change**

Not the kind of change that involves rearranging objects

It's about changes in the nature of a substance.

Fundamental Steps of the Scientific Method

Process that lies at the center of scientific inquiry.



Observation is often a chance occurrence

- Not necessarily planned for

Experiments: Observations that are planned, under controlled circumstances, with an effort to study the effects of different variables on an outcome

Hypothesis: A (hopefully) reasonable explanation of what has been observed. It may also be in the form of a proposed "law".

- Has to be testable (can be shown to be wrong)

Observation → Qualitative
Observation → Quantitative

Qualitative: Non-numerical

- gold is yellow
- lemon is sour
- air is a gas

Quantitative: Numerical measurements

- Gold surface absorbs more than 60% of visible light with wavelengths smaller than 520 nm
- Lemon juice has a pH of 2
- Air density at 1 atm pressure and 15° C is 1.225 kg/m^3

Which of the following is an example of a quantitative observation?

a) Solution A is a darker red color than solution B
b) The grass is green
c) Substance A has a greater mass than substance B
d) The temperature of the water is 45°C.

Interpreting the Behavior of Nature Scientifically
-- Law, Theory, Hypothesis

Law

A summary of repeatable observed (measurable) behavior.

- Tells us "**what** happens"
- Makes predictions
  - ➢Therefore can be shown to be wrong (at least under some conditions)

Theory (Model)

An explanation of an existing law (or laws) (i.e. observed behavior) based on more fundamental laws

- An explanation of the "**how**" and "**why**"
- Best explanation of observed behavior
- Regarded as legitimate even by scientists who don't favor it
- The closest that science comes to "the truth"
- Makes predictions
  - ➢Therefore can be shown to be wrong (at least under some conditions)

Hypothesis

- May hypothesize the existence of a certain law (if it survives enough testing, it becomes a "law")
- May provide a possible explanation for a "law" (if it survives enough testing, it becomes a "theory")
- As the "immature" version of a potential theory or a law, it makes predictions
  - ➢Therefore it can be shown to be false
- Needs to be verified

A theory explains the "why" and the "how" of a law in terms of more fundamental laws

- ➢ The more fundamental laws can in turn be explained by more fundamental theories

In practice what is effectively a law might not be called a "law". We may simply know "what happens" without naming it as a "law".

In practice a theory might be called a "model". "Model" and "theory" are used interchangeably.

- Both laws and theories allow us to make predictions.
- When a prediction turns out to be wrong under certain conditions, we don't necessarily discard a law or a theory:
  - ➢ We simply note the exception to the "law" or "theory"
  - ➢ More importantly, we can improve our **theory** and get:
    - ➢ A deeper understanding of nature
    - ➢ A more useful tool make predictions about the behavior of nature

      -- so we can design machines or structures that can operate under a wider range of conditions

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Measurement

- Quantitative observation.
- Has two parts:
  - ▪ number
  - ▪ scale (unit)

For example:

20 grams

45.8 liters

We will first focus on the number part

Significant Figures and Uncertainty

“**He weighs 215 pounds**”

Why wouldn’t it say 215.34706 pounds?

- The balance’s precision is limited to the nearest pound
  - The instrument itself may have fundamental limitations
  - It may be designed in a way that makes the person’s movements change the readings
- If the measurement can’t reasonably capture digits beyond whole pounds, those digits should not be reported

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“**He weighs 215 pounds**”

Does it mean the reported 215 pounds is absolutely correct to the nearest pound? No.

- The last digit is still uncertain.
- If we are lucky, it is off only by 1, making the uncertainty ±1
- But the uncertainty in the last digit may be any single-digit number, up to ±9

The manual that comes with the balance might state its precision. For example ±2 pounds

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What if the balance is rated at a precision of ±0.3 pounds? Should it still report your weight to the nearest pound?

- No, it would be wasting its precision.
- It should report the weight to nearest tenth (0.1) of a pound.
  - Even though the last digit it reports will have an expected “error” (i.e. “uncertainty”) in it

In fact, the reported digits should push right up to the limit where we have some idea of the last digit’s value, but we are <u>not totally certain</u>.

- That maximizes the amount of meaningful information conveyed

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So an engineer devises a spectacular new balance that has supreme precision, and is totally immune to the wiggling of the person being weighed as well. It can measure your weight to the nearest 0.00001 pound.

Should it report the weight as 215.34706 pounds?

No.

As the person stands, water is evaporating from the body, not to mention the water vapor and carbon dioxide (produced by the metabolism) exhaled with every breath.

The thing measured would constantly be changing to a degree (more than 0.00001 pound during any weighing) that would make that kind of precision meaningless.

Precision and Accuracy

The terms “precision” and “accuracy” in daily usage may be used interchangeably.

In science and technology, they are very distinct concepts.

Precision and Accuracy

Accuracy:

Agreement of the measurement(s) with the true value

Precision:

Agreement of the measurement(s) with one another

When do we call a number "precise" or "accurate"?

It's almost always left vague. Not here.

- Whether a number is "accurate" depends on its precision as well
- Whether a number is "precise" depends on its precision as well as that of the known true value

We will see how to properly think about those.

If the precision of the measurement matches the precision with which the "true value" is known, then we call it "precise"

A measurement is "accurate" if it agrees with the "true value" to within the claimed uncertainties of the measurement and the true value.

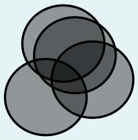
But first let's look at "precision" and "accuracy"

Precision

Degree of agreement among measurements of the same quantity.

For a set of measurements, the claimed precision of each must allow them to be consistent with one another.

3.5±0.3
3.6±0.3
3.4±0.3

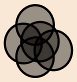


These measurements are reported with a precision that allows them to overlap

These measurements should claim less precision (larger value for the ± uncertainty) to allow overlap with one another

3.37±0.02
3.53±0.02
3.45±0.02

Like here:



3.4±0.1
3.5±0.1
3.5±0.1

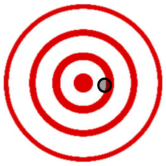
Accuracy

Degree of agreement of measurement(s) (within the claimed precision) with the true value. An accurate value may be imprecise.

Measurement: 3±1
True value: 2.4±0.2

A cannonball may be "accurate"
(it captures the center of the target)

Measurement: 2.8±0.1
True value: 2.4±0.2



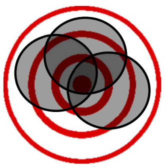
While a more precise bullet may be "not accurate"

Accuracy

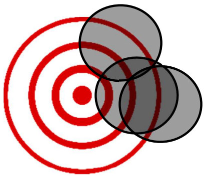
So, it's easier to be technically "accurate" when the precision is low.

—The guess that somebody's weight is "around 150 lb, give or take 50 lb" is technically "accurate" if the true weight is 114 lb.

—The relatively precise measurement of 117±1 lb is "not accurate" because it does not agree with the true value to within its **claimed** precision (the lowest consistent true value would be 116 lb)



Not precise but accurate



Not precise and not accurate

Precise* and accurate

Precise* but not accurate

* their precision matches that of the target value

Note that we took care to overlap the little circles (the measurements) in the target pictures.

You will often see pictures where the "precise" measurements are not overlapping. In other words, their claimed uncertainty (size of the circles) is smaller (i.e. higher precision) than their true uncertainty (they are not as precise as they claim).

"Precise" (?)

Maybe not after we make the circles large enough to overlap with one another

They would be significantly bigger than the "target" circle (the uncertainty in the true value)

Example

Measurements (meters)
121 ± 3
122 ± 3
119 ± 3

True length is 120.65 meters

- Are these measurements consistent with one another, given their stated precision (indicated by the ±uncertainty) Yes
- Are these measurements, individually, "accurate"? Yes
- If the true length is 120.5 meters, are these individual measurements "precise"? No

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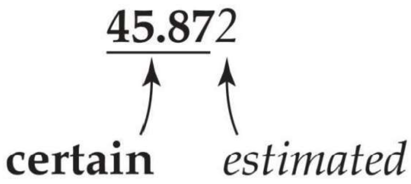
The boiling point of a liquid was measured in the lab:

Trial	Boiling Point
1	22.0 ˚C ± 0.2
2	22.1 ˚C ± 0.2
3	21.9 ˚C ± 0.2

The actual boiling point of the liquid is 28.7 ˚C. The results of the determination of the boiling point are:

a) accurate and precise.
b) precise but inaccurate. ⬅
c) accurate but imprecise.
d) inaccurate and imprecise.

Reporting Scientific Numbers



- The last significant (i.e. reported) digit is estimated.
- The last digit still contains information; it just isn't perfect
- Any further digits are regarded as "garbage", and should not be reported

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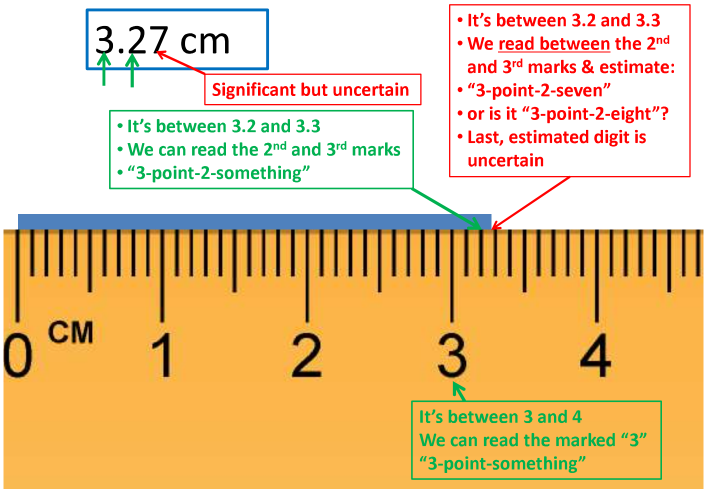
Estimating the last digit? How?

- An electronic instrument will estimate the last measurable digit automatically, but …
- If a human is making the measurement, we read down to the finest marks available, and then …
- We go one step further & literally "read between the lines"
  - We mentally divide the space between the finest marks into 10
  - Then we estimate the reading from these imaginary lines

Why not 5 lines?

Because we have a decimal system

Anatomy of a measurement



3.27 cm

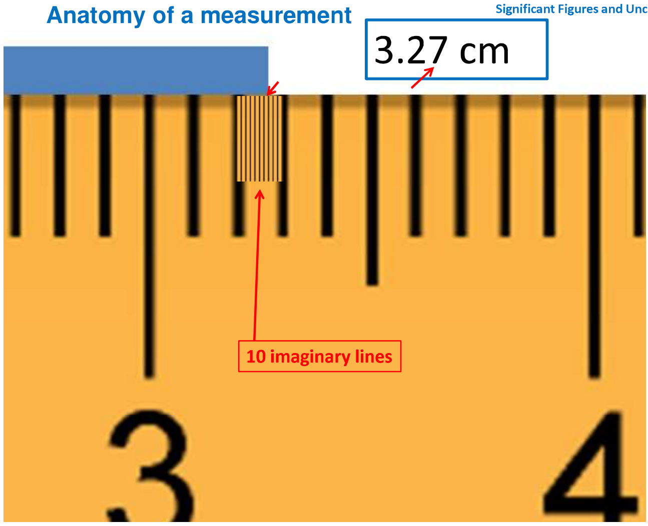
Significant but uncertain

- It's between 3.2 and 3.3
- We read between the 2nd and 3rd marks & estimate:
- "3-point-2-seven"
- or is it "3-point-2-eight"?
- Last, estimated digit is uncertain

- It's between 3.2 and 3.3
- We can read the 2nd and 3rd marks
- "3-point-2-something"

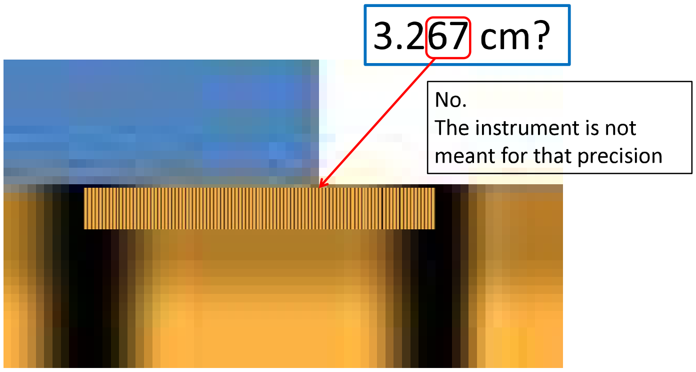
It's between 3 and 4
We can read the marked "3"
"3-point-something"

Anatomy of a measurement



But I zoomed in with my phone. Can I divide it into 100 lines then?

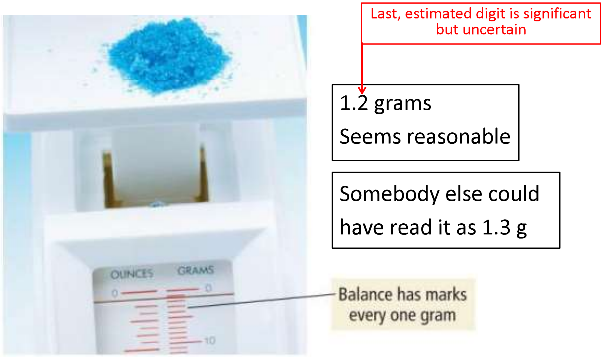
- The instrument is most likely not precise enough to justify it
- If it were, another level of finer marks would have been there



No.
The instrument is not meant for that precision

A weight measurement to tenth of a gram

- This balance has markings every 1 g.
- So we can estimate to the tenths place (i.e. 0.1 g)
- To estimate between markings, mentally divide the space into 10 equal spaces and estimate the last digit.



Last, estimated digit is significant but uncertain

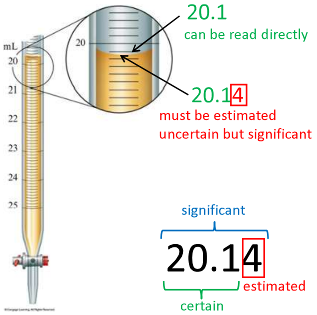
1.2 grams
Seems reasonable

Somebody else could have read it as 1.3 g

Balance has marks every one gram

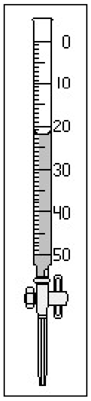
Measurement of Volume Using a Buret

- The volume is read at the bottom of the liquid curve (meniscus).
- First read all the digits from the markings
  - These digits are certain
  - And significant
- Then estimate one more digit
  - That digit is uncertain
  - Still significant

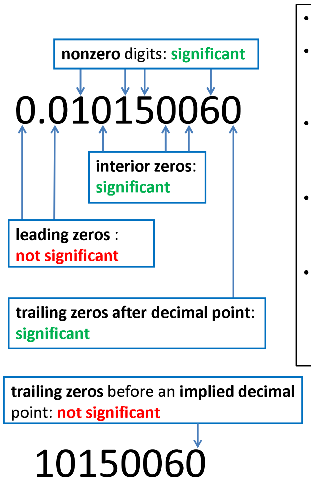


The buret is filled to the zero mark (at the top) with a solution and the solution is transferred to a beaker. What volume of transferred solution should be reported?

a) 20 mL
b) 22 mL
c) 22.0 mL ⬅
d) 22.00 mL
e) 25 mL



Significant Figures ("sig figs") in a Measurement



- **All nonzero** digits **are significant.**
- **Interior zeros** (zeros between two nonzero numbers **are significant.**
  - Regardless of where the decimal point is
- **Trailing zeros** (zeros to the right of a nonzero number) that fall after a decimal point **are significant.**
- **Leading zeros** (zeros to the left of the first nonzero number) **are NOT significant.** They serve only to locate the decimal point.
- **Trailing zeros at the end** of a number, but before an implied (but not written) decimal point, should be assumed not-significant.
  - If a trailing zero before the last zero is significant, there is no way to indicate it!

trailing zeros before an **actual decimal point**: significant

10150060.

Practice -- Counting Significant Figures

How many significant figures are in each number?

0.0035

1.080

2371

100.00

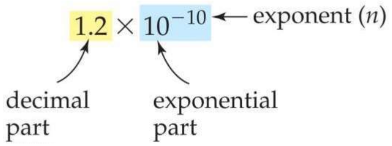
100,000

100,000.

When you have:
Leading zeros, ambiguous trailing zeros, or just too many zeros ….

A better way to write a number: Scientific Notation

- A **decimal part**:
  a number that is between 1 and 10 (but less than 10).
  - Shows only the significant figures
- An **exponential part**:
  10 raised to an integer power
  - Shows only the size (“order of magnitude”) of the number



The anatomy of 10^n

- A positive exponent n means: 1 multiplied by **10** n times
- A negative exponent n means: 1 divided by **10** n times

$$10^0 = 1 \qquad 10^{-1} = \frac{1}{10} = 0.1$$
$$10^1 = 1 \times 10 = 10 \qquad 10^{-2} = \frac{1}{10 \times 10} = 0.01$$
$$10^2 = 1 \times 10 \times 10 = 100 \qquad 10^{-3} = \frac{1}{10 \times 10 \times 10} = 0.001$$
$$10^3 = 1 \times 10 \times 10 \times 10 = 1000$$

How many significant figures does 10^n have?

- The 10 in 10^n is an exact number ---> infinite number of sig. figs.
- An exact number, multiplied by itself n times,
  still has an infinite number of significant figures.

Wait!
You wrote 10^n and then said “for a negative exponent n” Shouldn’t it be 10^{-n} ?

- We could’ve avoided this conversation at this point, but …
- Many students never quite connect algebraic thinking to the practice of technical thinking, …
  - And suffer.
- So we might as well get that done right here
- Let’s get used to the idea of “symbols”, just like those x’s and y’s in math classes
  - The symbol (i.e. “variable”) x in $y=x^2$ can be positive or negative
  - We still write the equation the same way, regardless, without a negative sign

If n=–6, 10^n is 10^{-6}

Decimal Format to Scientific notation

0.000001234 ---> 1.234×10^{-6}

To get a number with only one nonzero digit before decimal point

- The decimal point in 0.000001234 had to move to the **right** by 6 positions
- Or, there are 6 zeros (including the one before the decimal point) before the first nonzero digit
- So the exponent is –6

Implied decimal point

1234000 ----> 1.234×10^6

To get a number with only one nonzero digit before decimal point

- The (implied) decimal point in 1234000 had to move to the **left** by 6 positions
- So the exponent is +6 (i.e. just 6)

Decimal Format to Scientific Notation
-- Another couple of examples

$$5983 = 5.983 \times 10^3 \qquad 0.00034 = 3.4 \times 10^{-4}$$

Again:

- If the decimal point had to be moved to the **left**, the exponent is **positive**.
- If the decimal point had to be moved to the **right**, the exponent is **negative**.

Calculations with exponents

For any number (10 here) raised to various powers

- **Multiplication**: add powers

  $10^7 \times 10^8 = 10^{7+8} = 10^{15}$

- **Division**: subtract powers

  $10^7 / 10^8 = 10^{7-8} = 10^{-1}$

- Power to another power (**exponentiation**): multiply powers

  $(2 \times 10^4)^3 = 2^3 \times (10^4)^3 = 8 \times 10^{4\times 3} = 8 \times 10^{12}$

  Both terms are raised to the power that applies to the parentheses

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Practice

Perform the following the calculations without using a calculator for the exponent. Report your answer in scientific notation.

$(1 \times 10^7)/(1 \times 10^8)$

$(1 \times 10^2)^4$

$(1 \times 10^4) \times (1 \times 10^{-3}) / (1 \times 10^{-5})$

Practice

Convert to scientific notation.

Remember: You should only include the significant digits in the decimal part

350078000

350078000.

0.00043210

22.6×10^{-3}

0.00856×10^7

Practice

Convert to Standard (i.e. "decimal") notation.

Remember: You need to keep all the significant digits in the decimal part

5.19×10^3

5.190×10^3

5.1900×10^3

7.473×10^{-2}

0.7473×10^{-2}

Exact numbers

- They have an unlimited (infinite) number of significant figures
- For "Exact Numbers", we don't need to identify the significant digits
- All listed digits are significant, and ….
- We can assume an infinite number of implied, significant zeros wherever needed

  If 2.54 is supposed to be exact, then it's like

  2.540000000000……….

What numbers are "exact"?

- Exact counts (not estimates of a count) of discrete objects
  - 7 pencils → 7.00000000……..
- Integral numbers (integers) that are part of an equation
  - The "2" in y = 2x → 2.000000000……..
- Defined quantities
  - Speed of light is now actually defined and fixed: 299,792,458 meters/second
  - It has infinite number of significant figures, not just 9
  - Can be used like 299792458.0000000000000000……. meters/second
- *Some conversion factors are defined quantities, while others are not.*
  - For example, 1 in. = 2.54 cm exactly
  - It has infinite number of significant figures, not just 3
  - Can be used like 1 in. = 2.54000000000000000000……. cm

By the way, the "1" in conversion relationships is a "count", So it is exact

But be careful with exact numbers

- What if I am given a defined quantity with less digits than it was defined with?
- Such as, if I am given the speed of light as 3.00x10^{8} meters/second?
  - ➢ Then we treat it as a normal, not-exact number

- 1 pound (lb) is defined to be exactly 0.45359237 kilograms
- If we express the relationship to show how many lb there are in 1 kg:

  1 kg = 2.2046226218487758072297380134503……. lb !!!!!!!!

  (it's the reciprocal of 0.45359237, i.e. 1/0.45359237, and goes on forever)
  - ➢ In this form, there is no exact conversion factor.
- So if we are given, instead, 1 kg = 2.2046 lb, the number "2.2046" has only 5 significant figures, not infinite

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More Examples

Estimate the number of sig. figs:

42.301 ounces

0.000960 seconds

8.91500×10^{3} meters

103052.0 gallons

28 students

Value of a dollar bill

0.75 mile

Significant Figures in Calculations

Rules for Rounding:

- When numbers are used in a calculation, the result is rounded to reflect the significant figures of the *data*.

  -- We can of course be asked or required to round off to some other level of precision

  -- But normally we must get rid of all (and only) non-significant figures
- For calculations involving multiple steps, round only the final answer—*do not round off between steps*.

  This practice prevents small rounding errors from affecting the final answer.
  - This is why sometimes the last digit of the "correct answer" on a test may be off a little

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Rules for Rounding:

- Round down if the <u>first digit dropped</u> is 4 or less

- Round up if the <u>first digit dropped</u> is 6 or more

- Round up if the <u>first digit to be dropped</u> is 5, with at least one nonzero digit anywhere to its right

- If the first digit to be dropped is 5 and any digits to its right are zeros or there are no more digits

  -- We are used to always rounding it up

  -- But it's more complicated than that. 🤔

The proper way to round off if the first digit to be dropped is 5 and any digits to its right are zeros or there are no more digits e.g. 4.5, 0.250, 11.500

"Round-to-even": Round up or down in order to make sure the last digit kept or obtained is even.

If we kept rounding up whenever the original number is exactly midway between rounded-up and rounded-down versions, we would introduce an upward bias into our reported numbers!

BUT DON'T OVER-APPLY THIS RULE!

Do this only when:

The last digit to be kept is followed by 5
<u>and</u> there are *no nonzero digits after that 5*

Why bother with a "round-to-even" (or odd) rule?

Suppose we have the following measurements that come with two digits after they are multiplied by a conversion factor:

4.5, 4.5 , 5.5, 3.5, 4.5

but each number actually have only one significant digit. So we round to one digit before recording them.

- If we round **up** each number, we get:

  5, 5, 6, 4, 5

with an average of **5**, even though only one data point was larger than 5.

- If we apply the "round-to-even" rule, we get:

  4, 4, 6, 4, 4

with an average of **4.4**, which is much closer to the average of the original set, **4.5**

Applying the round-to-even convention:

2.350 ↑ last digit to keep

8.895 ↑ last digit to keep

2.05000 ↑ last digit to keep

1.95 ↑ last digit to keep

6.5 ↑ last digit to keep

We round up or down to produce a last digit that is even.
No net upward bias.

2.350 → 2.4
8.895 → 8.90
2.05000 → 2.0
1.95 → 2.0
6.5 → 6.

Convention could easily be "round to odd".
We just need to follow one convention consistently, and …
"Round to even" is the usual convention.

If you are still confused or scared by the "round-to even when exactly mid-way" rule:

Ignore it; you'll be fine

Don't spend too much time and effort on it

You can revisit it at higher level classes

After performing a calculation in the lab, the display on your calculator reads "0.023060070". If the number in the answer is to have **five** significant figures, what result should you report?

a) 0.0230
b) 0.00231
c) 0.023060
d) 0.2367
e) 0.02306

Practice

Round off the following numbers

1.237651 to 4 sig. figs

0.77555 to 3 sig. figs.

13.4219 to 2 sig. figs

114.1 to 2 sig figs

How many sig figs in the result of a calculation?

General rule: ***The lousiest one wins!***
(least precise)

"least precise" means different things for multiplication/division versus addition/subtraction:

multiplication & division

Precision = number of sig figs

123.4 more precise than 12.3
4 sig.fig. 3 sig.fig.

addition & subtraction

Precision = the last significant decimal place

3.45 more precise than 123.4
↑ ends @ 0.01 place ↑ ends @ 0.1 place

even though 123.4 has more sig figs

How many sig figs in the result of a calculation?

For Multiplication and Division:

The result of multiplication or division carries the same number of significant figures as the factor (number) with the fewest significant figures.

Significant Figures in Multiplication & Division

> "The result of multiplication or division carries the same number of significant figures as the factor (number) with the fewest significant figures"

$$\underset{\text{(3 sig. figures)}}{5.02} \times \underset{\text{(5 sig. figures)}}{89.665} \times \underset{\text{(2 sig. figures)}}{0.10} = 45.0118 = \underset{\text{(2 sig. figures)}}{45}$$

In this particular example:

The result (in blue) is rounded to two significant figures to reflect the least precisely known factor (0.10), which has two significant figures.

Significant Figures in Multiplication & Division

> "The result of multiplication or division carries the same number of significant figures as the factor (number) with the fewest significant figures"

$$\underset{\text{(4 sig. figures)}}{5.892} \div \underset{\text{(3 sig. figures)}}{6.10} = 0.96590 = \underset{\text{(3 sig. figures)}}{0.966}$$

In this particular example:

The result (in blue) is rounded to three significant figures to reflect the least precisely known factor (6.10), which has three significant figures.

Practice

Report the result of the following operations in the correct number of sig. figs

23.6×2.789

0.77555×2.89

$3.22 \times 10^{-2} \times 9.8$

Practice

Perform the following the calculations without using a calculator for the exponent. Report your answer in scientific notation.

Remember: You can treat the decimal and exponential parts separately because there is no subtraction or addition involved

$(5.3 \times 10^{7})/(6.1 \times 10^{8})$

$(3.92 \times 10^{2})^{4}$

$(4 \times 10^{4}) \times (2.5 \times 10^{-3}) / (8.98 \times 10^{-5})$

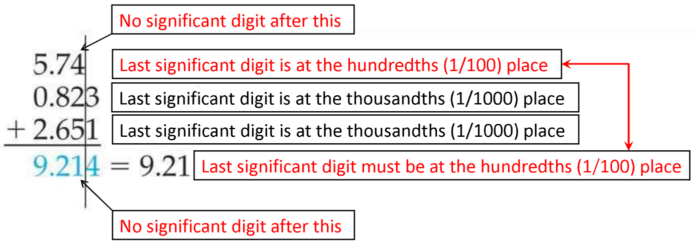
Significant Figures in Addition & Subtraction

For Addition and Subtraction:

The last decimal place of the result corresponds to that of the number whose last significant decimal place is the highest.

1. Find the number that ends at the highest-value (leftmost) decimal place
2. The result ends at that decimal place

> "For addition and subtraction, the last decimal place of the result corresponds to that of the number whose last significant decimal place is the highest"



No significant digit after this

5.74 — Last significant digit is at the hundredths (1/100) place

0.823 — Last significant digit is at the thousandths (1/1000) place

$+\,2.651$ — Last significant digit is at the thousandths (1/1000) place

$9.214 = 9.21$ — Last significant digit must be at the hundredths (1/100) place

No significant digit after this

We round the answer (in blue) to two decimal places after the point because 5.74 is the quantity whose last significant decimal is the largest, and ends at the second decimal place after the decimal point.

For addition and subtraction:

The number that ends at the leftmost (highest value) decimal place determines the rightmost decimal place allowed in the result.

$$\begin{array}{r} 4.8 \\ -\,3.965 \\ \hline 0.835 = 0.8 \end{array}$$

Ends at the “tenths” (1/10 = 0.1) place

Ends at the “thousandths” (1/1000 = 0.001) place

Ends at the “tenths” (1/10 = 0.1) place

You might see a different rule about addition & subtraction:

> *“The result has as many decimal places as the quantity with the fewest decimal places”*

But that applies only if:

- all the quantities have explicit decimal points!
- and “fewest decimal places” refers to after the decimal point

So that’s basically wrong.

Example:

Last significant decimal place is the 10s (tens) place (“lousier” one wins)

130 is the “lousier” one even if it has more decimal places than 12

$$130 + 12 = 1\underline{4}2 \rightarrow 140$$

Last significant decimal place is the 10s (tens) place

Last significant decimal place is the 1s (ones) place

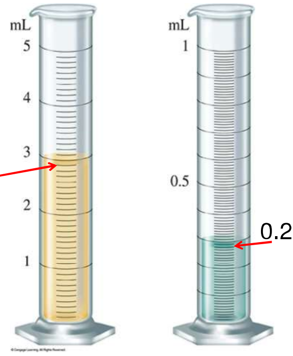
CONCEPT CHECK!

You have water in each graduated cylinder shown. You then add both samples to a beaker (assume that all of the liquid is transferred).

How would you write the number describing the total volume?

3.18 mL

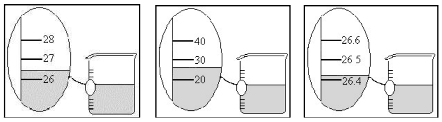
What limits the precision of the total volume?



$2.9\underline{0}$ mL

$0.28\underline{0}$ mL

The beakers below have different precisions.



All numbers are in “milliliters” (mL)

You pour the water from these three beakers into one container. What is the volume in this container reported to the correct number of significant figures?

a) 78.817 mL
b) 78.82 mL
c) 78.8 mL
d) 79 mL

Multiplication/Division & Addition/Subtraction Combined

In calculations involving both multiplication/division and addition/subtraction,

1. Do the steps in parentheses first;

   “parentheses” may not be explicit;
   operations in the numerator or the denominator are same as parentheses

2. If a number is sandwiched between a multiplication (or division) sign and an addition (or subtraction) sign, do the multiplication (or division) first

3. Determine the correct number of significant figures in the intermediate answer without rounding; then do the remaining steps.

   -- Put a bar under the last (and least) significant digit to keep track

---

In the calculation $3.489 \times (5.67 - 2.3)$

- Do the step in parentheses first: $5.67 - 2.3 = 3.\underline{3}7$
- Use the subtraction rule to determine that the intermediate answer has only one significant digit after the decimal point
- To avoid small errors, it is best not to round at this point; instead, underline the least (last) significant figure as a reminder.

$$3.489 \times 3.\underline{3}7 = 1\underline{1}.758 = 12$$

- Use the multiplication rule to determine that the answer (11.758) rounds to two significant figures (12) because it is limited by the two significant figures in $3.\underline{3}7$.

$2.01 \times \underbrace{(9.99 - 9.98)}_{0.01} + 0.123$

$2.01 \times 0.01 = 0.0\underline{2}01$

$0.0\underline{2}01 + 0.123 = 0.1\underline{4}31 \rightarrow 0.14$

Practice

Report answer with the correct sig. figs

- 2.23 – 0.1
- $8.12 \times 10^{-2} + 7.8$
- 7.8965 – 1.2
- To a beaker weighing 263.2 grams, you add 87.10 grams of water and 0.549 grams of sugar. Determine the combined mass of the three.

Practice

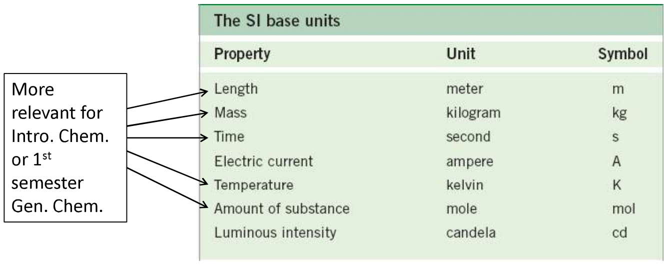
Report with the correct sig figs

$(3.1 \times 2.4367) - 2.34$

$(81.9 + 15.6) / 4.7$

The Basic Units of Measurement

The unit system for science measurements, based on the metric system, is called the International System of Units (*Système International d'Unités*) or ***SI* units**.



More relevant for Intro. Chem. or 1st semester Gen. Chem.

The SI base units

Property	Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Electric current	ampere	A
Temperature	kelvin	K
Amount of substance	mole	mol
Luminous intensity	candela	cd

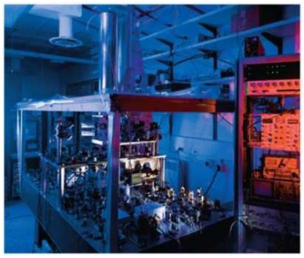
Units used to be tied to a tangible object

- A standard meter stick one could hold and measure
    - Would change over time, as the material "aged"
    - Would change with temperature; requiring great control
- A standard kilogram one could grab and calibrate a balance
    - Would change over time, as the material "aged"
- One 86400th of a day
    - Would change as Earth's rotation slows down
        - don't worry; very gradually
- Pro: Objects and concepts everyone understands
- Con: Not truly constant standards; so not good standards!

Modern "standards are tied to various fundamental physical constants.

The standard of time

- The second is defined, using an atomic clock, as the duration of 9,192,631,770 periods of the radiation emitted from a certain transition in a cesium-133 atom.
- No longer "one-86400th of a day"



The standard of length

- The definition of a **meter**, established by international agreement in 1983, is the distance that light travels in vacuum in 1/299,792,458 second.
- Because:
  The speed of light in vacuum is now **<u>defined</u>** to be 299,792,458 m/s)
- It also relies what a "second" is exactly
  - That's another standard
- There is no longer a "standard meter" kept in a vault

The standard of mass

- As of May 20, 2019, the definition of the **kilogram** is based on a fixed assignment of the Planck constant as 6.62607015 × 10^{-34} kg·m^2·s^{-1} (a fundamental constant of the universe), yielding a definition of the kilogram in terms of the <u>second</u> and the <u>meter</u> in a consistent system of units based on physical constants only.
- There is no longer a "standard kilogram" kept in a vault

- Modern standards (calibrations) for units are more abstract, and harder for "ordinary" people to understand
- But needed for the great precision modern science and technology achieves and requires

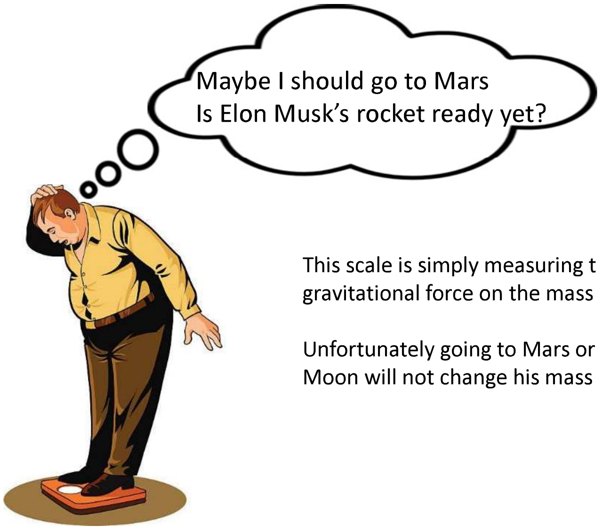
---

Standard foot kept in a vault in Subway headquarters

Just in case you took it seriously: **Just kidding**

Weight vs. Mass

- The kilogram is a measure of mass, which is different from weight.
- The **mass** of an object is a measure of the quantity of **matter** within it.
- The weight of an object is a measure of the gravitational pull on that matter.
- Consequently, **weight** depends on **gravity** while mass does not.



This scale is simply measuring the gravitational force on the mass (weight)

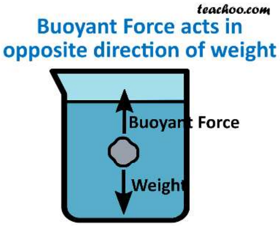
Unfortunately going to Mars or the Moon will not change his mass …

- Modern electronic scales and old-fashioned “spring” scales measure “weight” (gravitational force on the mass) directly, and give mass indirectly.
- Move them to a place where Earth’s effective gravity is slightly different, they will show a slightly different number

---

Unlike modern electronic scales, or “spring” scales, a traditional balance scale actually does **measure mass** because the gravitational force on the standard “weights” used change in the same way as the object weighed

---



Even the buoyant force due to air (pretty small, but not zero) will make a difference. Measurements for low-density objects may have some error.

- ➢ Less error with traditional balances
  - ✓ No error if the object has the same density as the standard weights used on the other side of the scale (same buoyant force)

Using Dimensional Analysis to Convert Between Units

What is Dimensional Analysis?

Using units as a guide to solving problems

- We will use it for unit conversions first
- It will be extremely handy in solving quantitative problems
  - ➢once you learn to see relationships as “conversion factors”
  - ➢and expand your idea of what could be a “unit”

Using Dimensional Analysis to Convert Between Units

- Units are multiplied, divided, and canceled like any other algebraic quantities.
- Always write every number with its associated unit.
- Always include units in your calculations, dividing them and multiplying them as if they were algebraic quantities.
- Do not let units appear or disappear in calculations. Units must flow logically from beginning to end.

Converting between units

For most conversion problems, we are given a quantity in some units and asked to convert the quantity to another unit. These calculations take the form:

$$\text{information given} \times \text{conversion factor(s)} = \text{information sought}$$

$$\cancel{\text{given unit}} \times \frac{\text{desired unit}}{\cancel{\text{given unit}}} = \text{desired unit}$$

Converting Between Units

A "conversion factor" is a ratio of two quantities known to be equivalent.

For example, if we know that 1 in. = 2.54 cm

we can construct the ratio as:

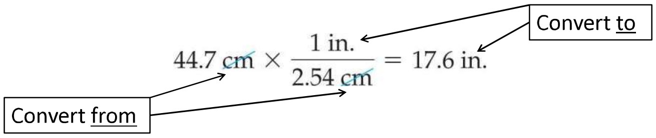
$$\frac{1\text{ in.}}{2.54\text{ cm}}$$

or as:

$$\frac{2.54\text{ cm}}{1\text{ in.}}$$

depending on which unit we need to convert to.

- Make sure the unit you want to convert from (i.e. "get rid of") is in the denominator (bottom part) of the conversion factor, so it cancels the starting unit and
- The unit you want to convert to is in the numerator (top part) of the conversion factor

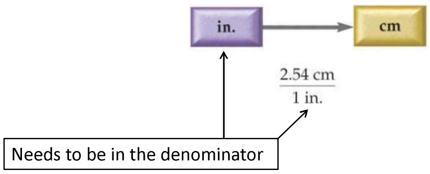


$$44.7\,\cancel{\text{cm}} \times \frac{1\text{ in.}}{2.54\,\cancel{\text{cm}}} = 17.6\text{ in.}$$

Convert to

Convert from

- To convert from inches to centimeters:

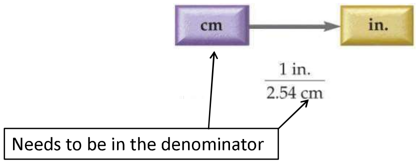


in. → cm

$$\frac{2.54\text{ cm}}{1\text{ in.}}$$

Needs to be in the denominator

- To convert from centimeters to inches:



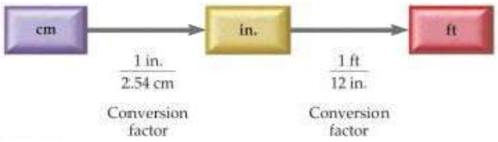
cm → in.

$$\frac{1\text{ in.}}{2.54\text{ cm}}$$

Needs to be in the denominator

Solving Multistep Unit Conversion Problems

- Each step should have a conversion factor with the units of the previous step in the denominator and the units of the following step in the numerator.
- If we are given the conversion between inch and cm, and inch and ft, but **not** between cm and ft, we construct the following:



cm → in. → ft

$$\frac{1\text{ in.}}{2.54\text{ cm}} \qquad \frac{1\text{ ft}}{12\text{ in.}}$$

Conversion factor

Conversion factor

SOLUTION

$$194\,\cancel{\text{cm}} \times \frac{1\,\cancel{\text{in.}}}{2.54\,\cancel{\text{cm}}} \times \frac{1\text{ ft}}{12\,\cancel{\text{in.}}} = 6.36\text{ ft}$$

SI Prefix Multipliers

Prefix	Symbol	Meaning	Multiplier	
tera-	T	trillion	1,000,000,000,000	(10^{12})
giga-	G	billion	1,000,000,000	(10^{9})
mega-	M	million	1,000,000	(10^{6})
kilo-	k	thousand	1,000	(10^{3})
hecto-	h	hundred	100	10^{2}
deca-	da	ten	10	10^{1}
deci-	d	tenth	0.1	(10^{-1})
centi-	c	hundredth	0.01	(10^{-2})
milli-	m	thousandth	0.001	(10^{-3})
micro-	μ	millionth	0.000001	(10^{-6})
nano-	n	billionth	0.000000001	(10^{-9})
pico-	p	trillionth	0.000000000001	(10^{-12})
femto-	f	quadrillionth	0.000000000000001	(10^{-15})

Most reliable way to construct prefixed SI unit conversions

Simply use the numerical equivalent of the prefix

- You must get used to dealing with powers of 10
  - Negative power or not (10^9 or 10^{-12})
  - In the numerator or the denominator
  - You can't put off getting competent in that; used everywhere
- Put a "**1**" in front of the prefixed unit
  - The prefixed unit already has the prefix; it's fine already
- Put the numerical equivalent of the prefix (e.g. 10^{-3} for milli) in front of the un-prefixed unit

Convert 2.5 μm to meters:

μ = "micro" = 10^{-6}

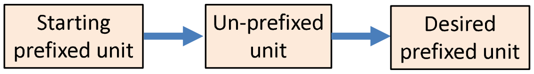
un-prefixed unit

$$2.5\ \cancel{\mu m} \times \frac{10^{-6}\ m}{1\ \cancel{\mu m}} = 2.5 \times 10^{-6}\ m$$

How to convert from one prefixed SI unit to another

Use the un-prefixed SI unit (meter, gram, second) as the intermediate unit

- First convert from the starting prefixed unit (e.g. km, pg, fs) to the un-prefixed SI unit (meter, gram, second)
- Then convert from un-prefixed SI unit (meter, gram, second) to the desired prefixed unit


Starting prefixed unit → Un-prefixed unit → Desired prefixed unit

nanometer — meter — picometer

$$0.121\ \cancel{nm} \times \frac{10^{-9}\ \cancel{m}}{1\ \cancel{nm}} \times \frac{1\ pm}{10^{-12}\ \cancel{m}} = 121\ pm$$

Practice

- How many mg does a 433 kg sample contain?

m = milli = 10^{-3}
k = kilo = 10^3

4.33×10^8 mg

- How many km is 1.25×10^8 cm?

c = centi = 10^{-2}
k = kilo = 10^3

1.25×10^3 km

Practice

A jogger runs at an average speed of 6.50 miles/hour. Convert the speed to meters/second (1 km = 0.6215 mi)

Conversion with Units Raised to a Power -- for Area

- Area has the units of length-squared.
- So we square (instead of cube) both sides to obtain the proper conversion factor for area.

$$\begin{gathered} 1\ in = 2.54\ cm \\ \downarrow \\ (1\ in)^2 = (2.54\ cm)^2 \\ \downarrow \\ 1^2\ in^2 = 6.4516\ cm^2 \\ \downarrow \\ 1\ in^2 = 6.4516\ cm^2 \end{gathered}$$

$$308\ \cancel{cm}^2 \times \frac{1\ in^2}{6.4516\ \cancel{cm}^2} = 47.7\ in^2$$

Practice

- If a room requires 25.4 square yards of carpeting, what is the area of the floor in units of ft^2? (3 ft = 1 yd exactly)
- A sheet of notebook has an area of 603.22cm^2. What is the area of this paper in nm^2?

Volume and Volume Units

- Measures how much space is taken up by a sample or an object
- An important quantity in general
- Especially useful in chemistry
- Allows easy measurement of quantity, especially of liquids
- A popular unit is “Liters” (L)
- The SI prefixes are also used with liters

Practice

- The volume of a certain raindrop is 0.0036 cL. Convert to microliters.

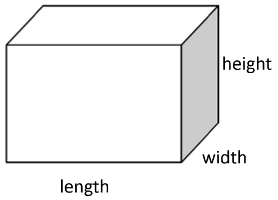
- A bottle contains 750 mL of soda. How many quarts of soda is that?
  (1 quart = 0.946353 L)

Volume as a Derived Unit

- A derived unit is formed from other units.
- Many units of **volume,** a measure of space, are derived from **length**.
- Any unit of length, when **cubed** (raised to the third power), becomes a unit of volume.

“height” *“width”*

Volume = length x length x length



Volume as a Derived Unit

- Cubic meters (m^3), cubic centimeters (cm^3), and cubic millimeters (mm^3) are all examples of SI units of volume
  - Derived from the SI units for length
- Liter (usually for liquids) is also an SI unit because

  $1\ L = 1\ dm^3$
- How do we convert between different units of volume derived from length?
  - Similar to what we did for area units
  - Same process, whether SI units or not

Conversion with Units Raised to a Power -- for Volume

- Volume has the units of length-cubed.
- So we cube both sides of the length conversion equation to obtain the proper conversion factor for volume.

$$(2.54\ \text{cm})^3 = (1\ \text{in.})^3$$

$$(2.54)^3\ \text{cm}^3 = 1^3\ \text{in.}^3$$

$$16.387\ \text{cm}^3 = 1\ \text{in.}^3$$

We can do the same thing in fractional form:

$$1255\ \cancel{\text{cm}^3} \times \frac{(1\ \text{in.})^3}{(2.54\ \text{cm})^3}$$

$$\Big\downarrow$$

$$1255\ \cancel{\text{cm}^3} \times \frac{1\ \text{in.}^3}{16.387\ \cancel{\text{cm}^3}} = 76.5851\ \text{in.}^3 = 76.59\ \text{in.}^3$$

How to obtain volume conversion factors (recap)

- Start from the length conversion relationship between the two units
- **Cube** both sides of the equation
- Any numbers in front of length units also get cubed

$$1\text{ m} = 3.28084\text{ ft}$$

$$\downarrow$$

$$(1\text{ m})^3 = (3.28084\text{ ft})^3$$

$$\downarrow$$

$$1^3\text{ m}^3 = 3.28084^3\text{ ft}^3$$

$$\downarrow$$

$$1\text{ m}^3 = 35.3147\text{ ft}^3$$

How to obtain volume conversion factors
-- important example, with SI prefixes

$$1\text{ cm} = 10^{-2}\text{ m}$$

$$\downarrow$$

$$(1\text{ cm})^3 = (10^{-2}\text{ m})^3$$

$$\downarrow$$

$$1^3\text{ cm}^3 = (10^{-2})^3\text{ m}^3$$

$$\downarrow$$

$$1\text{ cm}^3 = 10^{-6}\text{ m}^3$$

So, when we see

1 cm^3

Is it 1 centi(m^3) (one 100th of a m^3)? **No**

Or is it 1 (centimeter)^3 ? **Yes**

- "Centi" doesn't simply mean 10^{-2} here
- It leads to 10^{-6} in the case of volume as we just saw
- And we learned how to handle it

Practice

How many cm^3 are contained in $3.77 \times 10^4\text{ mm}^3$?

37.7 cm^3

Practice

The sides of a block of solid copper metal are measured to be 6.58 cm, 4.95 cm, and 0.82 cm (length, width, height). Determine the volume and report it in with the correct number of sig. figs and units.

---

1 Liter (L) is equal to 1 dm^3. Express mL in terms of cm^3.

$$1\cancel{\text{mL}}\ \mathrm{x}\ \frac{10^{-3}\ \cancel{\text{L}}}{1\ \cancel{\text{mL}}}\ \mathrm{x}\ \frac{1\ \cancel{\text{dm}^3}}{1\ \cancel{\text{L}}}\ \mathrm{x}\ \underset{\uparrow \atop \left[\frac{10^{-1}\text{m}}{1\text{ dm}}\right]^3}{\frac{10^{-3}\ \cancel{\text{m}^3}}{1\ \cancel{\text{dm}^3}}}\ \mathrm{x}\ \underset{\nwarrow \atop \left[\frac{1\text{ cm}}{10^{-2}\text{m}}\right]^3}{\frac{1\text{ cm}^3}{10^{-6}\ \cancel{\text{m}^3}}} = 1\text{ cm}^3$$

Density

- Mass of substance per unit volume
- An important characteristic of materials

We might desire high density if we need large mass in a small volume

– a styrofoam cannonball or a cotton-ball paperweight won’t be effective

We might desire low density for a given amount of strength

– A bike made of “carbon fiber” will be a lot lighter than one made of steel

Calculating Density

The density of a substance is the ratio of its mass to its volume.

We calculate the density of a substance by dividing the mass of a given amount of the substance by its volume.

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}} \quad \text{or} \quad d = \frac{m}{V}$$

Common units: **g/cm^3** or **g/mL**

Can be used for solids or liquids

Usually used for liquids

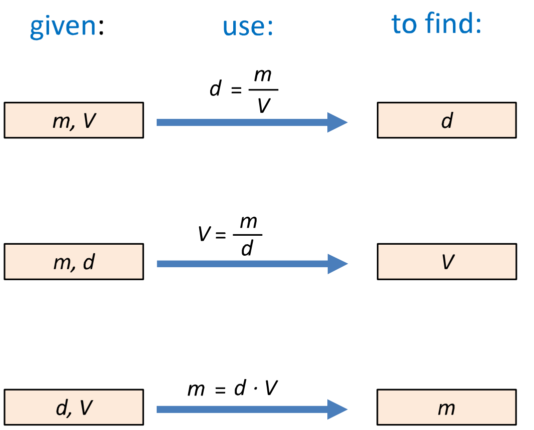
	Material	Density (g/cm^3)
	Outer space	10^{-24}
Least dense solid	Graphene Aerogel	0.00016
	Lightest Silica aerogel	0.001
	Air	0.0012
Least dense metal	Lithium	0.534
	Alcohol	0.78
	Water	1.00
	Aluminum	2.70
	Lead	11.34
Densest liquid	Mercury	13.5
Densest element	Osmium	22.59
	Neutron star	10^{14}

Example

If a sample of liquid has a volume of 22.5 mL and a mass of 27.2 g, what is its density?

$$d = \frac{m}{V} = \frac{27.2\ \text{g}}{22.5\ \text{mL}} = 1.21\ \text{g/mL}$$

A Solution Map Involving the Equation for Density



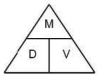
given:	use:	to find:
m, V	$d = \frac{m}{V}$	*d*
m, d	$V = \frac{m}{d}$	*V*
d, V	$m = d \cdot V$	*m*

The equations

$$d = \frac{m}{V} \qquad V = \frac{m}{d} \qquad m = d \cdot V$$

The equations all represent the same relationship, just rearranged algebraically

Wanting to avoid “algebra” to avoid mistakes in favor of pre-college “magic shortcuts” like



is counterproductive.

- You will soon have to deal with basic algebra; be ready
- Those who resort to shortcuts end up often using them incorrectly anyway (I have seen it enough times in lab reports)

Example

A certain mineral has a density of 7.57 g/cm3. What is the volume of a sample with a mass of 17.8 g?

$$d=\frac{m}{V} \quad \Rightarrow \quad V=\frac{m}{d}$$

$$V=\frac{m}{d}=\frac{17.8\,\cancel{g}}{7.57\,\cancel{g}/\text{cm}^3}=2.35\,\frac{1}{1/\text{cm}^3}=2.35\ \text{cm}^3$$

Example

What is the mass of a 49.6-mL sample of a liquid, which has a density of 0.85 g/mL?

$$d=\frac{m}{V} \quad \Rightarrow \quad d\cdot V=m \quad \Rightarrow \quad m=d\cdot V$$

$$m=0.85\,\frac{\text{g}}{\cancel{\text{mL}}}\times 49.6\,\cancel{\text{mL}}=\mathbf{42\ g}$$

Thinking of Density as a Conversion Factor

- When things are related by simple multiplication and division, i.e. by direct or inverse proportionality, we can use a ratio like density as a "conversion factor"
- Between mass (*m*) and volume (*V*) in this case

Example:

For a liquid substance with a density of 1.32 g/cm^3, what volume should be measured to deliver a mass of 68.4 g?

$$68.4\,\cancel{g}\times\frac{1\ \text{cm}^3}{1.32\,\cancel{g}}=51.8\ \text{cm}^3$$

Note how we place g and cm^3 in order to cancel g and obtain cm^3

Example:

For a liquid substance with a density of 1.32 g/cm^3, what mass is contained in a volume of 51.8 cm^3?

$$51.8\,\cancel{\text{cm}^3}\times\frac{1.32\ \text{g}}{1\,\cancel{\text{cm}^3}}=68.4\ \text{g}$$

Note how we place g and cm^3 in order to cancel cm^3 and obtain g

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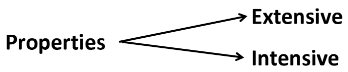
Density as a Conversion Factor

- Using density as a "conversion factor" is an example of using "dimensional analysis" to solve quantitative problems involving proportionality
  - We will use dimensional analysis that way
  - Not just with density problems
- Much more efficient and effective than setting up proportions as in elementary and middle school
  - Wean yourself off setting up proportions
- Much safer; helps you avoid mistakes
  - Especially when you are tired/sleep-deprived/panicked



Properties

Extensive

Intensive

Extensive: Proportional to amount
More of the substance, bigger the property
Examples:

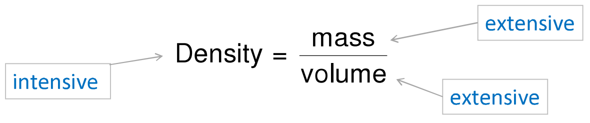
- mass
- volume
- total heat capacity

Intensive: Independent of amount
Same value regardless of amount of substance
Examples:

- density
- viscosity
- heat capacity per gram

An intensive property is basically a ratio of two extensive properties

- Easier to see that for some properties than others



$$\text{Density}=\frac{\text{mass}}{\text{volume}}$$

intensive

extensive

extensive

Temperature

-- a measure of kinetic energy per particle
-- an intensive property

It's a measure of kinetic energy per particle

Higher temperature → faster particles

- First quantification and measurements of temperature (i.e. thermometers) didn't directly use its true nature
  - Fahrenheit (˚F) and Celsius (˚C) are the best known
- They assumed that liquids expanded perfectly linearly with temperature
  - Not a bad assumption, but not exact either

Three Systems for Measuring Temperature

Fahrenheit (˚F)
Celsius (˚C) (sometimes called "centigrade")
Kelvin (K) (no "degrees"; just "Kelvins")

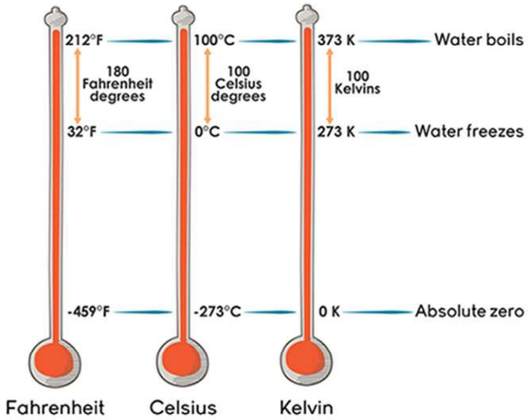
1 ˚C change is equal to 9/5=1.8 ˚F change

⇨ **˚C is a larger unit**

1 ˚C change is equal to 1 K change

⇨ **˚C and K are equal-sized units**

The Three Major Temperature Scales



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Converting Between Scales

$$T_K = T_C + 273.15 \qquad T_C = T_K - 273.15$$

$$T_C = \left(T_F - 32^\circ\text{F}\right)\frac{5^\circ\text{C}}{9^\circ\text{F}} \qquad T_F = T_C \times \frac{9^\circ\text{F}}{5^\circ\text{C}} + 32^\circ\text{F}$$

Same equation, same relationship, just rearranged

Makes sure the units come out right

$$T_C = \left(T_F - 32^\circ\text{F}\right)\frac{5\boxed{^\circ\text{C}}}{9\boxed{^\circ\text{F}}} \qquad T_F = T_C \times \frac{9\boxed{^\circ\text{F}}}{5\boxed{^\circ\text{C}}} + 32^\circ\text{F}$$

Note the relative sizes of different "degrees":

Apart from shifting the scale by 32°F, we have the conversion factor $\frac{5^\circ\text{C}}{9^\circ\text{F}}$

Which means
5°C = 9°F
°C = 9/5 °F
°C = 1.8°F But only for **changes** in temperature

Practice

At what temperature are Celsius and Fahrenheit temperatures numerically equal?

T_C = temperature in ˚C
T_F = temperature in ˚F

$T_C = T_F$ A specific condition imposed

substitute

$T_C = (T_F - 32)\,\frac{5}{9}$ Always true

$\Downarrow$

$T_F = (T_F - 32)\,\frac{5}{9} \xRightarrow{\text{Feeling more comfortable with } x?} x = (x - 32)\,\frac{5}{9}$

$x = T_F = T_C = -40$

So −40 ˚F = −40 ˚C

Practice

At what temperature is $T_C = -2T_F$?

substitute $T_C = -2T_F$

$$T_C = (T_F - 32)\,\frac{5}{9}$$

$$\downarrow$$

$$-2T_F = (T_F - 32)\,\frac{5}{9}$$

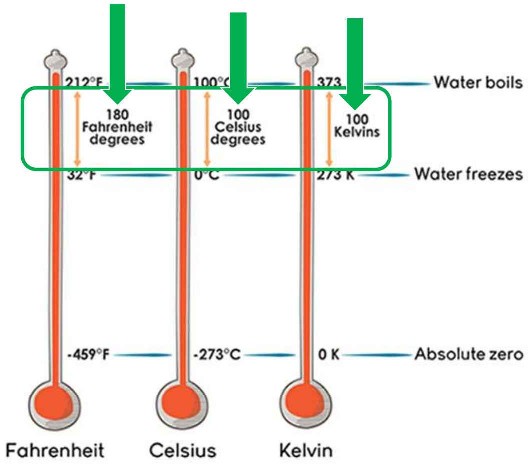
$$\Rightarrow \quad T_F = 7.0\ ^\circ\mathrm{F}$$

$$T_C = (7.0 - 32)\,\frac{5}{9} = -14\ ^\circ\mathrm{C} \qquad (T_C = -2T_F \text{ verified})$$

Temperature

Comparing temperature changes as measured in different units

- Conversion equations don't apply to temperature changes
- Only the relative size of the units matter
  - If a temperature change is 10˚C, the same change would be measured as 18˚F
    - Each 1˚C change corresponds to 1.8˚F
  - If a temperature change is 10 K, the same change would be measured as 10˚C, and as 18˚F
    - Each 1 K change corresponds to 1˚C
    - Each 1˚C change corresponds to 1.8˚F
    - Each 1 K change corresponds to 1.8˚F



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"delta"

Change (Δ) in temperature only cares about the relative sizes of the different "degrees".

$\Delta T_K = \Delta T_C$ Kelvin and Celcius "degrees" are the same size

Fahrenheit degree size is 1.8 times smaller than Celcius and Kelvin

$$\Delta T_F = 1.8\ \Delta T_C$$

$$\Delta T_F = 1.8\ \Delta T_K$$

So we need to multiply the temperature **change** measured in Celcius or Kelvin by 1.8 to get the **change** in Fahrenheit.

Practice

- Lead melts at 601.0°C. What temperature is this in °F?
- The outside air temperature is 30°F, what is the temperature in Kelvin?
- Helium boils at 4K what is the temperature in °C
- If you raise the thermostat by 5°F, by how much will the temperature change in °C?
- Temperature change between day and night on the moon is around 540 °F. What is it in Kelvins?

Matter

Anything that has mass.

Matter under "ordinary conditions" exists in three states

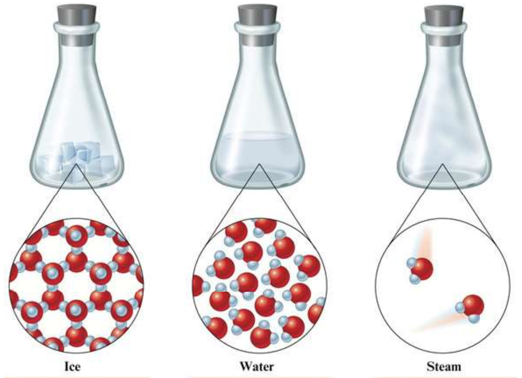
- ○ Solid
- ○ Liquid
- ○ Gas

} Stuff chemists normally deal with

There is matter in nuclear particle radiation, stars (plasma), black holes, neutron stars, etc. In fact, most "normal" matter exists in forms chemists don't deal with.

Furthermore, most of the universe is made of stuff we know nothing about: "dark matter" (and "dark energy").

The Three States of Water



Solid: The water molecules are locked into rigid positions and are close together.

Liquid: The water molecules are still close together but can move around to some extent.

Gas: The water molecules are far apart and move randomly.

Solid

- Rigid
- Has fixed volume and shape
- Particles wiggle at a constant position
  - ➢ Like restless kids sitting in chairs

Liquid

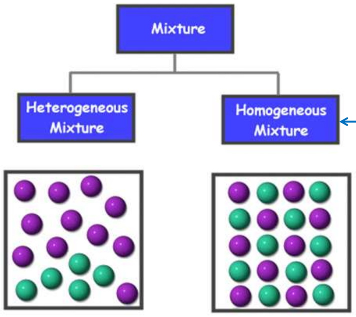
- Has definite volume but no specific shape.
- Assumes shape of container.
- Has density similar to solid
- Particles are about as close to one another as in solids, they are in contact with a lot of other particles, but they are constantly moving around
  - ➢ Like people mingling in a party

Gas

- ▪ Has no fixed volume or shape.
- ▪ Takes on the shape and volume of its container.
- ▪ Much lower density than solids and liquids
- ▪ Particles have little or no tendency to stick around; they move around and spend as much time away from one another as allowed by the volume of the container they are in
  - ➢ Like adventurers wandering the globe

Mixtures

More than one substance occupying the same volume.



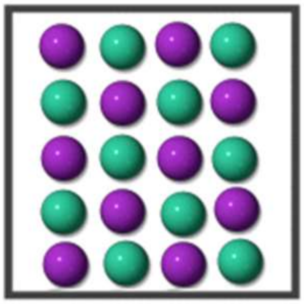
Synonymous with "solution"

Components can be solid, liquid, or gas

Homogeneous Mixture

So intimately mixed that no matter how closely we look into it, the mixture looks uniform, and we find the same proportion of the constituents, until we come to the individual, characteristic particles of the substances.

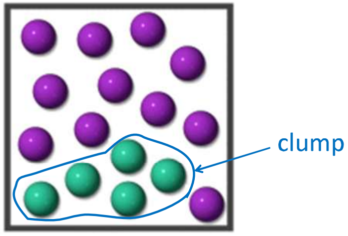
Also called a **"solution"**



Heterogeneous Mixture

Individual components clump together. If we look closely, we can see the individual components forming distinct pieces or regions.

A mixture may look homogeneous, but at close enough inspection (say, using microscopes), if we can discern regions with different compositions, we have a heterogeneous mixture.



As long as the clumps (regions) are made up of many molecules (or individual atoms as the case may be), the mixture is heterogeneous.

Visible light scatters off the “clumps” in a heterogeneous mixture when the clumps are larger than the wavelength (basically the size of the light waves)

- If a mixture is not “clear” (if it is “milky”, or “cloudy”) it is heterogeneous.
  - unless the individual molecules, even when not clumped together, are so huge that they act like “clumps” (not common; we can ignore this possibility for our purposes)
- If a mixture is “clear” (not “milky” or “cloudy”) there is a good chance it is homogeneous (therefore a “solution”)
  - but it may be a heterogeneous mixture with clumps too small for visible light to scatter off
  - so it’s not guaranteed to be truly homogeneous
  - we would need further tests to verify it is truly homogeneous

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Which of the following is a homogeneous mixture?

- Pure water
- Gasoline
- Jar of jelly beans
- Soil
- Pure copper metal

Physical Change

- Change in the form of a substance, not in its chemical composition. No chemical bonds are formed or broken.

  Example: boiling or freezing water
- Can be used to separate a mixture into pure compounds, but it will not change the compounds that make up the mixture

Separation techniques
utilize differences in physical properties

- Distillation – boiling point
- Filtration – particle size
- Chromatography – solubility in a solvent and the tendency to “stick” to a surface

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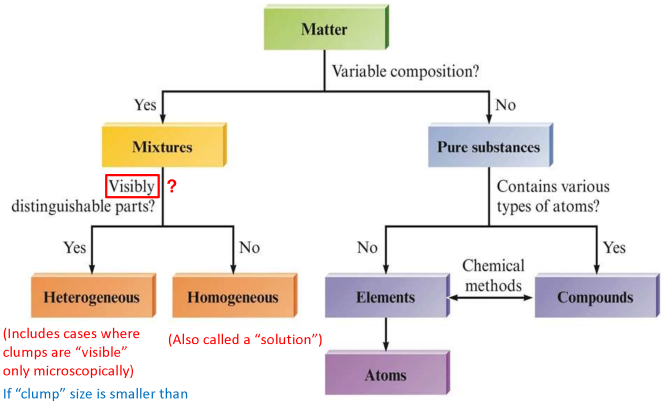
Chemical Change

- Substance(s) becomes a new substance or substances with different properties and different composition.
  - Example: Bunsen burner (methane reacts with oxygen to form carbon dioxide and water)

Bonds between atoms form or break

Which of the following are examples of a chemical change?

- Pulverizing (crushing) rock salt
- Burning of wood
- Dissolving of sugar in water
- Melting a popsicle on a warm summer day



(Includes cases where clumps are "visible" only microscopically)

(Also called a "solution")

If "clump" size is smaller than the wavelength of light, it's "invisible", but may have trillions of molecules. Still not truly "homogeneous".