CHEM 103
Measurement in Chemistry

Lecture Notes
January 26, 2006
Prof. Sevian

Agenda

- Calculations skills you need:
  - Dimensional analysis
  - Significant figures
  - Scientific notation
- Group problem #1

Some Important Chemistry Skills

- Using dimensional analysis to solve problems

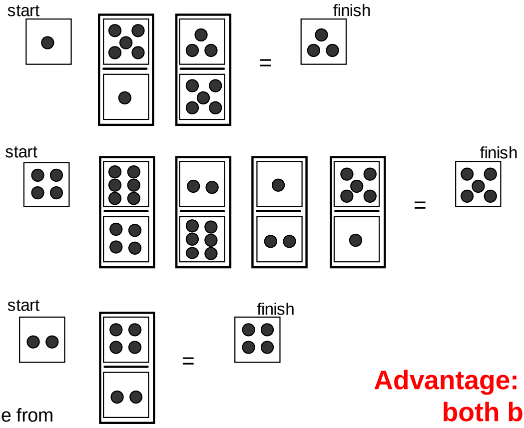
  Why is it useful?
  - Ensures correct units
  - Can find missing information

- Keeping track of significant digits in calculations

  Why is this important?
  - Measurements taken by instruments limit accuracy of information

Dimensional Analysis

There is only one rule

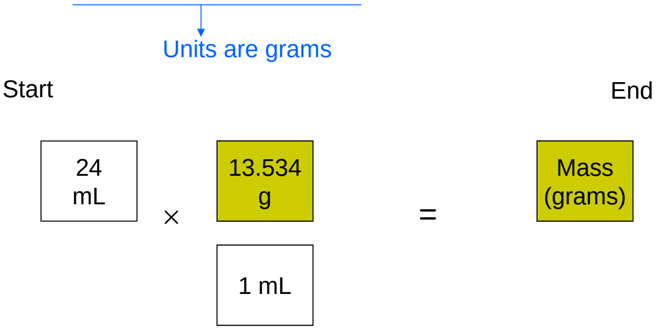


Dominoes image from Sevian et al, ***Active Chemistry*** (2004)

Advantage: You can work both backwards and forwards to solve a problem

Example: Density Problem

Mercury has a density of 13.534 g/mL. What is the mass of 24 mL of mercury?



Units are grams

Start

End

$$24\ \text{mL} \times \frac{13.534\ \text{g}}{1\ \text{mL}} = \text{Mass (grams)}$$

Dimensional analysis

What to look for:

1. What data are given?
2. What quantity do you need?
3. What conversions are available to take you from start to end?

What you need to be able to do:

- Recognize dimensions by both names and units
  - Volume or (mL, L, cm^3; and mL is the same as cm^3)
  - Mass or (g, kg)
  - Density or (units of mass/units of volume)
- Determine starting and ending information
- Work toward the middle from both ends
- Do the calculations properly

Another density example

volume

12. cm^3

A 2.0 cm x 2.0 cm x 3.0 cm rectangular piece of metal has a mass of 32.4 g. What is the metal's density?

mass

Units of mass/volume

Start

End

$$\frac{32.4\ g}{12\ cm^3} = \frac{2.7\ g}{cm^3}$$

Significant Digits

Why are they important?

- Tell you to what extent you can “trust” the data
- Tell you how reliable is the least reliable instrument that was used in determining the data

What goes into any calculation?

Three kinds of information

1. Measured values
2. Exact values
3. Derived (calculated) values

The least reliable one determines the reliability of the final calculated result.

Measured Values

Measurement	Units most often used	Instrument often used
Length Wavelength	m, cm nm	Ruler (spectrometer)
Mass	g, kg	Balance
Volume	L, mL, m^3, cm^3	Ruler (to measure dimensions), graduated cylinder (liquids)
Time	s	Clock
Temperature	ºC, ºF, K	Thermometer
pH	No units	pH gauge
Pressure	kPa, atm, mmHg	manometer

Exact Values

- Integer-based values
  - Fractions: ½, ¾
  - Counting numbers: 2 electrons
  - Metric conversions: 1 meter = 100 cm
  - Two important exact conversions you need to remember
    - Distance: 1 inch = 2.54 cm
    - Energy: 1 calorie = 4.184 Joules
- Constants of nature are usually treated as exact values
  - Speed of light in vacuum, c = 2.99792458 x 10^8 m/s
  - Pi, π = 3.141592654… (no units)
  - Planck's constant, h = 6.62617636 x 10^{-34} J·s
  - Gas constant, R = 8.3144126 J/mol·K

Derived (Calculated) Values

Value	Units most often used
Amount of a substance	mol
Energy	Joules, cal, kcal
Solution concentration	Molarity (M) = mol/L
Density	g/mL or g/cm^3
Velocity	m/s
Molar mass	g/mol

A Note About SI

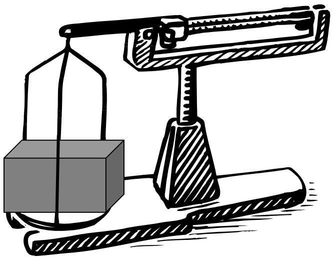
SI = *Le Système International d'Unités* (International system of units)

- If you use SI units exclusively in a calculation, then the answer will always come out in SI units. This is why people memorize SI units (makes life easier).
- Unfortunately, in many cases, convention (or ease) is to not use SI units, so you have to convert.

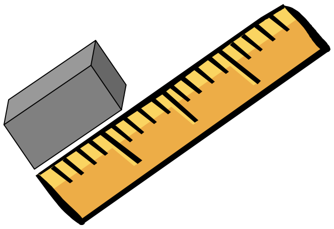
Not SI (must convert)	SI units
Grams (g)	Kilograms (kg)
Centimeters (cm)	Meters (m)

Example

Determining the density of a rectangular block of aluminum

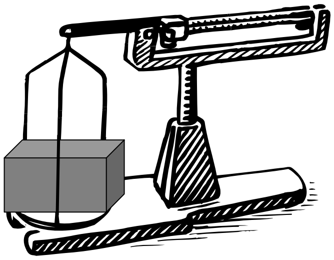


Mass is 32.4030 grams



Dimensions are

2.0 cm x 2.0 cm x 3.0 cm



Mass is 32.4030 grams

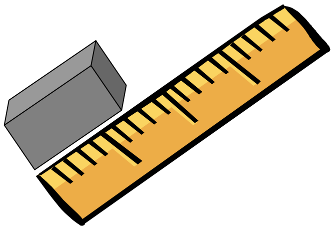
How many digits are *significant*?

Significant = to have meaning

Mass measurement

32.4030 grams

Six digits of information



How many digits are *significant*?

Length measurement

Dimensions are

2.0 cm x 2.0 cm x 3.0 cm

Need to calculate volume from this

2.0 cm

Two digits of information

Same number of digits of information for the other two length measurements

Rules for Sig Figs

- Generally, count the digits
- Zeroes written to the left don't count

  0.00056 has 2 sig figs
- Zeroes written to the right do count

  81.00 and 0.0008100 both have 4 sig figs
- Convention for numbers not containing a decimal point

  7200 has 2 sig figs, 7200. has 4 sig figs
- See pp. 22-23 in the text for rules

Keeping Track of Sig Figs

Density calculation:

1. Calculate volume from length measurements
2. Calculate density from mass and volume

Keeping Track of Sig Figs

Volume calculation

V = L x W x H

2 sig figs **2 sig figs** **2 sig figs**

= (2.0 cm) x (2.0 cm) x (3.0 cm)

= 12 cm **Final answer must be rounded to 2 sig figs**

Keeping Track of Sig Figs

Density calculation

6 sig figs

$$D = \frac{m}{V} = \frac{32.4030\ \text{g}}{12\ \text{cm}^3}$$

2 sig figs

Calculator gives 2.70025

$$= 2.7\ {}^{\text{g}}\!/_{\text{cm}^3}$$

Must round to 2 sig figs

Rules for Sig Figs

- Multiplication/division rule
  - The measurement with the least total sig figs wins
- Addition/subtraction rule
  - The measurement with the least decimal places (compared to the decimal point) wins
- Other rules can wait until you need to do more complicated calculations

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Example of Addition Rule

Problem

Find the sum: 28.6 + 8.289 + 0.003 + 1007.56

Solution

Line up the numbers at the decimal point, compare, and cut off at least significant (compared to decimal point)

$$\begin{array}{r@{.}l} 28&6 \\ 8&289 \\ 0&003 \\ +\ 1007&56 \\ \hline 1044&452 \end{array} \xrightarrow{\text{rounds to}} 1044.5$$

Some Measurements and Conversions You Need to Know

Two types of conversions

- Proportional
  - Time
  - Length or distance
  - Volume
  - Mass
- Equations
  - Temperature

Time Conversions

$$2.5\ \text{hours} \times \frac{?\ \text{something}}{?\ \text{hours}} \times \frac{?\ \text{seconds}}{?\ \text{something}} = ?\ \text{seconds}$$

$$2.5\ \text{hours} \times \frac{60\ \text{minutes}}{1\ \text{hour}} \times \frac{60\ \text{seconds}}{1\ \text{minute}} = 9000\ \text{seconds}$$

2 sig figs

2 sig figs

How can you write 9000 so it's clear it has 2 sig figs?

HOLD THAT THOUGHT…

Metric Prefix Meanings for Conversions

Prefix	Meaning	Example
Centi (c)	1/100th of	1 cm = 0.01 m 100 cm = 1 m
Milli (m)	1/1000th of	1 mL = 0.001 L 1000 mL = 1 L
Kilo (k)	1000 of	1 kg = 1000 g
Micro (μ)	10^{-6} of	1 μmol = 10^{-6} mol 1,000,000 μmol = 1 mol
Nano (n)	10^{-9} of	1 nm = 10^{-9} m

Conversions Using Metric System

1) How many moles are in 12.2 mmol?

3 sig figs

$$12.2 \text{ mmol} \times \frac{1 \text{ mol}}{1000 \text{ mmol}} = 0.0122 \text{ mol}$$

exact conversion

3 sig figs

2) Red light has a wavelength of 630 nm. How many meters is that?

2 sig figs

$$630 \text{ nm} \times \frac{10^{-9} \text{ m}}{1 \text{ nm}} = 6.3 \times 10^{-7} \text{ m}$$

exact conversion

2 sig figs

Volume Conversions

Important volume conversion to remember: $1\,\text{mL} = 1\,\text{cm}^3$

A can of soda is marked as having 258 cc of soda in it.
How many liters is this?

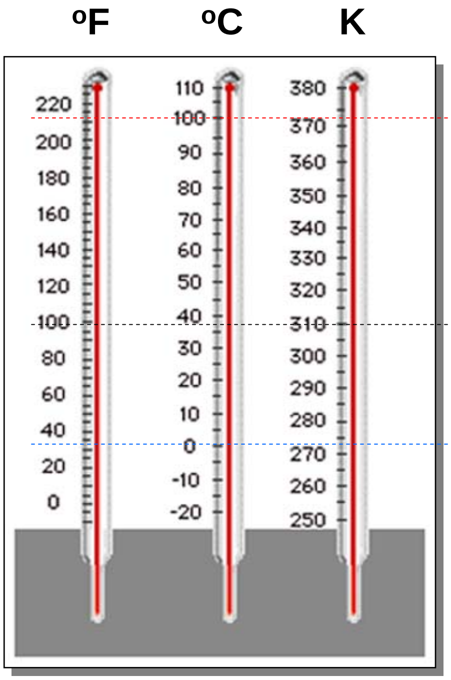
3 sig figs

$$258\ \text{cm}^3 \times \frac{1\ \text{mL}}{1\ \text{cm}^3} \times \frac{1\ \text{L}}{1000\ \text{mL}} = 0.258\ \text{L}$$

exact conversion **exact conversion**

3 sig figs

Temperature Scales



From Chemistry & Chemical Reactivity 5th edition by Kotz / Treichel. C 2003. Reprinted with permission of Brooks/Cole, a division of Thomson Learning: www.thomsonrights.com. Fax 800-730-2215.

Boiling point of water

212 °F = 100 ° C = 373 K

Human body temperature

98.6 °F = 37 ° C = 310 K

Freezing point of water
Melting point of ice

32 °F = 0 ° C = 273 K

Temperature Conversions Require Equations

Celsius (ºC) ↔ Kelvin (K)

You need to memorize this conversion!!!

$$K = {}^{o}C + 273.15 \quad or \quad {}^{o}C = K - 273.15$$

Example: A gas has a temperature of 25.8 ºC. What is the temperature in Kelvin?

$$K = {}^{o}C + 273.15$$

$$= \underbrace{25.8 + 273.15}_{\text{Sig figs: One place beyond decimal pt}} = \cancel{298.95\ K} \longrightarrow \text{Rounds to } 299.0\ K$$

Really Big Numbers

In a 22.4 liter sample of air at standard conditions, there are approximately this many particles present:

$$6\underbrace{02}_{2},\underbrace{204}_{3},\underbrace{531}_{3},\underbrace{000}_{3},\underbrace{000}_{3},\underbrace{000}_{3},\underbrace{000}_{3},\underbrace{000}_{3}$$

$$= 6.02204531 \times 10^{\boxed{23}} \text{ particles}$$

Really Small Numbers

A single snowflake has a mass of approximately

$$0.\underbrace{000\;003}_{3\quad\;\; 3}0\text{ kg}$$

$$=3.0 \times 10^{\boxed{-6}}\text{ kg}$$

Calculations Using Scientific Notation

- A typical snowflake has 100=10^2 ice crystals
- A single ice crystal has 10^{18} water molecules
- A water molecule has a mass of 3.0×10^{-26} kg
- Therefore, a typical snowflake has a mass of approximately

$$10^2\ crystals \times \frac{10^{18}\ water\ molecules}{1\,crystal} \times \frac{3.0 \times 10^{-26}\,kg}{1\ water\ molecule} = 3.0 \times 10^{-6}\,kg$$

Data taken from http://hypertextbook.com/facts/2001/JudyMoy.shtml

Scientific Notation

- A nice way to represent big and small numbers
- Makes it easy to indicate significant figures
  9000 written with two sig figs is 9.0×10^3
- Makes it easy to estimate answers
  $(3.0 \times 10^8) \times (2.0 \times 10^{-6}) = 6.0 \times 10^2$
- Scientific notation and your calculator → try the practice problems in the Assignments section on the course website to make sure you are proficient at using scientific notation in your own calculator

How group problems work

- You are assigned to a group for the entire semester.
- Group problems will be given occasionally during the last 15 minutes of class.
- I will provide 5x8 index cards on which to write your group's answers. You will turn in one index card per group. Everyone in the group will receive the same grade. Students who are not present will not receive credit.
- On one side of the index card, write your group letter and the names of every person in your group who is present. On the other side, write your group's solution to the problem. You do not need to copy the question onto the card.
- You are strongly encouraged to keep a copy of your own solutions to these problems. Every exam will contain one of the group problems.