Welcome to CHEM 1101

Lecture Times: Tuesday and Thursday 9 to 9:50 a.m.

Instructor: Dr. Muhannad Amer

Office Location: 44 staff Bldng

Required Materials

- *Text: Chemistry,* 7th ed., Zumdahl

Please see course web page in blackboard to access worksheets that will be used in the Discussion Sections.

Quiz scores

On topics from Lecture and Discussion that week

Each quiz is worth 15 points –
keep 10 highest
Total quiz score: 150 pts

On-line Homework

13 Homework assignments on

Please check due dates in the on-line *Syllabus*

Credit is rewarded for all problems **completed by deadline**

- **Beyond the chemistry theory of this class taking this class will enable you to:**

- Apply knowledge to solve new problems.
- Analyze information you have gathered.
- Work with and delegate responsibility to others.
- Have confidence in yourself and your work.
- Be organized in your thoughts and actions.
- Ask and answer questions.

Scientific Method

- The scientific method provides the method by which scientists solve problems.
- Chemists use this method to understand matter at the atomic or molecular level.

observation

hypothesis

(explanation of observation)

prediction

Carrying out experiment A will prove the hypothesis by giving result B

experiment

•Observations

•Observations can be quantitative or qualitative.

•Measurement of physical phenomenon

•All measurements <u>MUST</u>
•consist of a number and a unit!

•**Example**: charge of an electron is 1.60 x 10-19 coulombs

•Scientific notation

•1.60 x 10-19 = 0.000000000000000000160

Scientific Notation

number x 10^n ← integer

↓

1-9

$1.60 \times 10^0 = 1.60 \times 1 = 1.60$

$1.60 \times 10^1 = 16.0$

$1.60 \times 10^{-1} = 0.160$

1.60 or 1.6 or 1.600 can be used

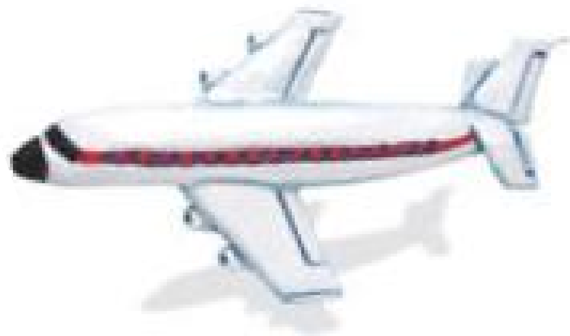
Are Units of Measurement that Important?

July 23rd, 1983: **Gimli Glider**, an Air Canada aircraft ran out of fuel

Needed for trip: 22,300 **kg** of fuel

Used to fill plane: 22,300 ***pounds*** of fuel (10,115 kg !)

Not enough fuel!



Important SI (International system) base units

Quantity	SI Base Unit
Length	meter (m)
Mass	kilogram (kg)
Time	second (s)
Temperature	Kelvin (K)
Amount	mole (mol)

Volume = length^3

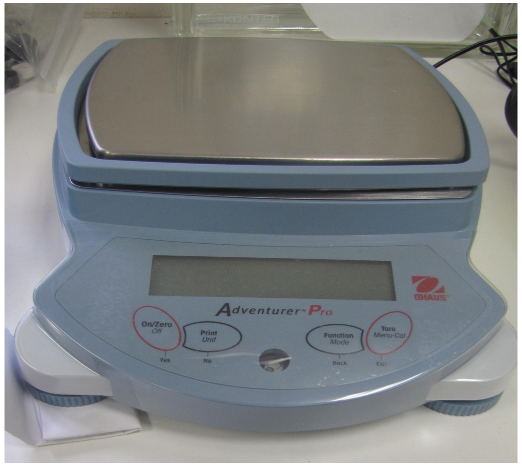
1L = 1 dm^3 = 1000 cm^3 = 10^{-3} m^3 = 1000 ml

1cm^3 = 1ml

Common Prefixes used to adjust the size of Base Units

Prefix	Meaning	Abbreviation	Exponential Notation
deci-	tenth of	d	10^{-1}
Mega-	million	M	10^{6}
kilo-	thousand	k	10^{3}
centi-	hundredths of	c	10^{-2}
milli-	thousandths of	m	10^{-3}
micro-	millionths of	μ	10^{-6}
nano-	billionths of	n	10^{-9}
pico-	trillionths of	p	10^{-12}

Uncertainty in Measurement



The number obtained in measurement is obtained using a measuring device that introduces **some degree of uncertainty to this measurement and this must be indicated.**

Uncertainty in the measurement lies in the **last digit and is assumed to be +1 or -1**

Recorded measurement of 0.0508 g

Actual mass is 0.0507 g – 0.0509 g

$$= 5.07 \times 10^{-2} - 5.09 \times 10^{-2}\ g$$

Significant Figures

The recorded certain and the first uncertain digit or estimated number of a measurement are called its significant figures.

Rules for Significant Figures

1. Digits from 1-9 are always significant.

Example: 26981 has 5 significant figures

2. Zeros between two other significant digits are <u>always</u> significant. **Example: 1023 has 4 significant figures**

3. One or more additional zeros to the right of both the decimal place and another significant digit <u>are</u> significant.

Example: 5.00 and 500. both have 3 significant figures

Significant Figures

4. Zeros used solely for spacing the decimal point (placeholders) are <u>not</u> significant.

Example: 0.000231 has 3 significant figures

5. The **absence** of a decimal point means terminal zeros are NOT significant.

Example: 600 has 1 significant figure

6. **Exact numbers** have an infinite number of significant figures. They are obtained via counting, e.g. 1 dozen eggs, or by definition, e.g. the 2 in $2\pi r$. When used in calculations, exact numbers do **not** limit the number of significant figures.

Multiplying and Dividing Significant Figures

22.37 x 3.10 x 85.75 = 5946.50525 = 5950

4 sig. figs 3 sig. figs 4 sig. figs

Seen on calculator but **not** to be recorded as the answer

Least number of significant figures dictates the number of significant figures to be stated in the calculated answer

5946.50525
9 sig. figs

5950
3 sig. figs

Rounding $\geq$ 5 round **up** < 5 round **down**

Calculated results are <u>never</u> more reliable than the measurements they are obtained from.

Adding and Subtracting Significant Figures

3.76 + 14.83 + 2.1 = 20.69

2 dec. places | 2 dec. places | 1 dec. place

Seen on calculator but **not** to be recorded as the answer.

Least number of decimal places dictates the number of decimal places to be stated in the calculated answer.

20.69
2 dec. places

Rounding to one dec. place

20.7
1 dec. place

Calculated results are <u>never</u> more reliable than the measurements they are obtained from.

Addition (subtraction) with Multiplication (Division)

$$\frac{732.11 + 6.3}{760.00}$$

do addition (subtraction) **first**

$732.11 + 6.3 = 738.41$ (738.4)

2 decimal place   1 decimal place

NEVER round intermediate results for multistep calculations

do division (multiplication) **last**

$$\frac{738.4}{760.00}$$

4 sig fig

5 sig fig

$$\frac{738.41}{760.00} = 0.971592105$$

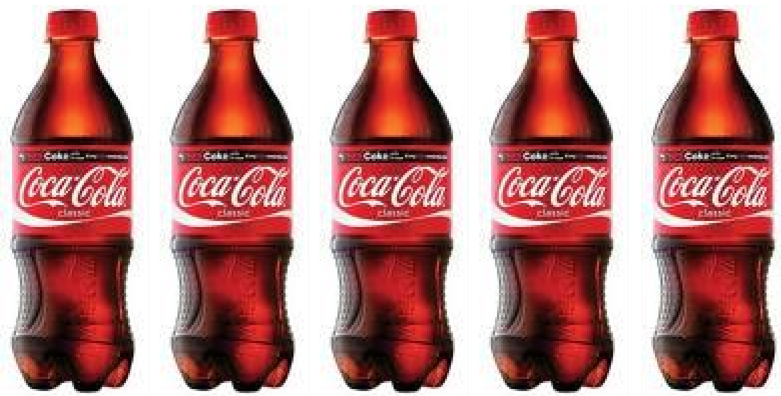
(Not to be recorded as the answer)

Answer: 0.9716 (4 sig fig)

Precision vs. Accuracy of Calculated Results

Precision	Accuracy
❖ reproducibility ❖ check by repeating measurements ❖ poor precision results from poor technique.	❖ correctness ❖ check by using a different method ❖ poor accuracy results from procedural or equipment flaws.

Precision = reproducibility



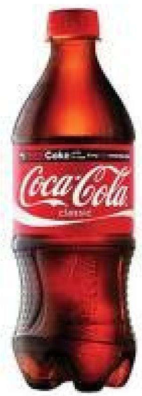
How much of a clone are you?

Accuracy = Closeness of measured value to standard value

Standard values

Sugar content: 54 grams

pH: 2.6



How do **you** measure up?

Dimensional Analysis

A problem-solving method that uses the fact that any number or expression can be multiplied by one without changing its value.

Unit factors may be made from any two terms that describe the same or equivalent "amounts" of what we are interested in.

1 inch = 2.54 centimeters

Unit factors

$$\boxed{\frac{1\ \text{inch}}{2.54\ \text{centimeters}} \quad \text{or} \quad \frac{2.54\ \text{centimeters}}{1\ \text{inch}}}$$

Steps for Using Dimensional Analysis

Steps:

1.Identify what units are required, what units have been given.

2. State the equivalent of these units.

3. Multiply the given data and its units **by the appropriate unit factors so that only the desired units are present at the end.**

Example: How many centimeters are in 6.00 inches?

Units required: centimeters

Units *given* : inches

1 inch = 2.54 centimeters

Unit factors ⇨

$$\boxed{\frac{1\ \text{inch}}{2.54\ \text{centimeters}} \quad \text{or} \quad \frac{2.54\ \text{centimeters}}{1\ \text{inch}}}$$

$$\boxed{?\ \text{cm} = 6.00\ \cancel{\text{in}} \times \frac{2.54\ \text{cm}}{1\ \cancel{\text{in}}} = 15.2\ \text{cm (to 3 significant figures)}}$$

Notice that the unit factor was chosen that allowed the **units required to remain while the other cancels** during the calculation.

Temperature Scales and Interconversions

Kelvin (K) - The "Absolute temperature scale"

At absolute zero and only has positive values.

Celsius (^oC) - Commonly used scale around the world and in laboratories.

Fahrenheit (^oF) - Commonly used scale in America for weather reports.

$T\ (K) = T\ (^oC) + 273.15$

$T\ (^oF) = 9/5\ T\ (^oC) + 32$

$T\ (^oC) = T\ (K) - 273.15$

$T\ (^oC) = 5/9\ T\ (^oF) - 32$

Density

Density is the mass per unit volume of a substance and has compound units of grams per cubic centimeter (g/cm^3)

Example: Calculate **the density** of an object that has a volume of 64 cm^3 and a mass of 34g.

Solution:

$$\text{Density} = \frac{\text{mass}}{\text{volume}} \qquad \text{Density} = \frac{34g}{64cm^3} = 0.53g/cm^3$$

What is the mass, in grams, of 1.00 gallon of water ?

The density of water is 1 g/mL (1 ml of water = 1g)

Solution

Units given: gallon, g/ml

Units required: g

1 gal = 4 qt	1.057 qt = 1 L	1 L = 1000 ml	1 g = 1 mL

given

required

All equivalent values are EXACT numbers and do **not** limit the number of significant figures in the answer.

$$1.00\,\cancel{\text{gal}} \times \frac{4\,\cancel{\text{qts}}}{1\,\cancel{\text{gal}}} \times \frac{1\,\cancel{\text{L}}}{1.057\,\cancel{\text{qts}}} \times \frac{1000\,\cancel{\text{mL}}}{1\,\cancel{\text{L}}} \times \frac{1\text{ g}}{1\,\cancel{\text{mL}}}$$

3 sig figs

$$= 3784.295 = 3780\text{ g} = 3.78 \times 10^{3}\text{ g}$$

calculator

3 sig figs