Welcome to AP Chemistry 2025!

Please complete the following **FIVE** assignments according to **due dates**---Mrs. Lawrence

Assignment #1: Letter of Introduction.

We are going to spend a lot of time together this year, so it's best if we get a head start on learning a bit about you. Your first assignment is to successfully send an e-mail letter to me by **August 1st,** at jlawrence@powerscatholic.org. Be sure to watch for a reply e-mail to confirm receipt of this! Draft the e-mail following these guidelines:

a. Use clearly written, full sentences. Do not abbreviate words like you are texting with a friend. Use spell check! This is a professional communication like you would have with a college professor, so let's practice for your rapidly nearing future!
b. Address it to the appropriate teacher: **jlawrence@powerscatholic.org**
c. Make the **Subject**: "**AP Chemistry: Introduction to <Insert Your Name Here>**" (Do not include the quote marks or the brackets, just the words)
d. Begin the e-mail with **a formal salutation**, like "Mrs. Lawrence," or "Dear Mrs. Lawrence,"
e. Now introduce yourself (your name) and tell me a little bit about yourself:
   - What do you like to do (hobbies, sports, music, interests, etc.)?
   - Do you have a job? Volunteer? Intern? Capstone ideas??
   - Tell me a little bit about your family (Mom? Dad? Guardian? Siblings? Pets?)
   - Was there anything that you liked/disliked about your earlier science classes?
   - What is your strengths and weaknesses as a student? Why?
   - What has been your favorite science lab thus far? Why?
   - What are you plans for the future?
   - Why are you taking AP Chemistry?
   - What are you looking forward to the most in AP Chemistry?
   - What are you most anxious about in AP Chemistry?
   - Is there anything else I should know about you as we begin this journey?
   - Lastly, include a **picture** so I can recognize you on Day 1!
f. End the e-mail with a **formal closing**: "Cordially", "Sincerely", "Warm regards", etc. and add your name (as you would like to be called) as if you signed a letter.

Assignment #2: Textbook and Introduction Reading/Notes

Download the textbook for this year: **OpenStax. Chemistry 2e**. Become familiar with the setup/arrangement of the book. Read and take **notes** over the following sections: 1.4 Measurements, 1.5 Measurement Uncertainty, Accuracy and Precision and 1.6 Mathematical Treatment of Measurement Results by **August 17th**. We will upload these into Google Classroom on the first day of class.

Assignment #3: Math Fundamentals

We will be completing many lab measurements, math problems and conversions throughout this course. Please read and complete the assigned worksheets "Rules for Uncertainty, Sig Figs and Scientific Notation" and "Rules for Unit Conversions and Dimensional Analysis" found below by **August 17th.** We will upload these into Google Classroom on the first day of class.

Assignment #4: T-Shirt Design

If you don't know, we create a class T-shirt every year for this class. Have an original, full page T-shirt design (front and back with color) that you would like to have for our class by **August 17th**. We will vote on the design the first week of school. We will upload these into Google Classroom on the first day of class.

Assignment #5: AP Chemistry supplies

Please collect the following supplies by **August 17th** as we will be starting with content and lab ASAP: lab coat and goggles (you may purchase new or gently used if needed), binder, graphing calculator (the one you use for math class is just fine), and an attitude of enthusiasm!

AP Chemistry is a rigorous course that requires commitment and GRIT---I am excited to get started! Are you?

AP Chemistry Summer Review Part I: Uncertainty in Measurement and Calculations:

1. Exact Numbers:

Counted numbers and *definitions* do not involve any measurement and are considered as exact numbers with an infinite number of significant figures. Do not consider them when determining significant figures for your final answer.

Definitions: 1 week = 7 days.
1 mile = 5,280 feet
1 yard = 3 feet

Counted: 5 Players on the basketball court.
23 students in a room
25 pennies used by a class in an experiment.

2. Measured Numbers:

All ***measured numbers*** have some degree of uncertainty.

When recording measurements, ***record only the significant figures***. Record measurements to include one decimal estimate beyond the smallest increment on the measuring device.

Examples (consider a measuring instrument like a ruler):

- If smallest increment = 1m, then record measurement o 0.1m (i.e. 3.1***m***)
- If smallest increment = 0.1m, then record measurement to 0.01m (i.e. 5.67 ***m***)
- If smallest increment = 0.01m, then record measurement to 0.001m (i.e.12.675 ***m***)

c. Unless otherwise stated the uncertainty in the last significant figure *(the uncertain digit)* is assumed to be ±1 unit. Modern digital instruments and many types of volumetric glassware will state the level of uncertainty.

3. Rules for counting Significant Figures.

a.***Non-Zero Numbers***: Non-zero numbers are always significant.
b. ***Zeros:***

1: **Leading zeros** that come before the first non-zero number are ***never*** significant
2. **Captive zeros** (*sandwich zeros*) that fall between two non-zero digits are ***always*** significant.
3. **Ending zeros** that appear after the last non-zero digit are significant only when a decimal point appears somewhere in the number.

Examples:

Number	0.005	5005	5005.00	500.	0.0050
Sig Figs	1	4	6	3	2

c. Scientific Notation: Significant figures are recorded in the mantissa (*number* $1 \le x < 10$*)* **Examples:**

Number	3.0×10^{3}	5.998×10^{5}	6.00000×10^{-23}	0.5×10^{4}
Sig Figs	2	4	6	1

4. Rules for Using Significant Figures in Calculations

(a) Multiplication, Division, Powers and Roots:-**"*LEAST SIG.FIG RULE*"**

1. The result should be reported to the same number of significant figures as the measured number having the ***least number of significant figures.***

2. Only consider the number of significant figures in each of the ***measured numbers! (not constants)***

Example 1:

$2.3 \times 5.78 - \textit{Calculator returns } 13.294$

$2.3 \textit{ has 2 sig.fig}$

$5.78 \textit{ has 3 sig.fig.}$

$2.3 \times 5.78 = 13 \quad \textit{The answer must be rounded to show 2 sig.fig}$

Example 2.

$$\frac{1.67x10^{5} \times 0.00045}{2x10^{-23}} = \textit{calculator returns } 2.505000000x10^{24}$$

$1.67x10^{5}$ *has 3 sig.figs*

0.00045 has 2 sig.figs

$2x10^{-23}$ *has 1 sig.fig*

$$\frac{1.67x10^{5} \times 0.00045}{2x10^{-23}} = 3x10^{24} \textit{ (rounded to 1 sig.fig)}$$

Example 3

$\sqrt{2.3} = \textit{calculator returns } 1.516575089$

$2.3 \textit{ has 2 sig.figs}$

$\sqrt{2.3} = 1.5 \quad \textit{round answer to 2 sig.figs}$

(b) Addition and Subtraction: "***LEAST PRECISE DECIMAL RULE***"

1. The result should be reported with the same decimal precision as the measured number having the uncertain digit in the ***least precise decimal place.***

2. Only consider the decimal precision in each of the ***measured numbers! (not constants)***

Example 4: a – c

a. $123cm + 5.35cm - 128cm \textit{ (rounded to } 10^{0})$

b. $1.0001m + 0.0003m = 1.0004m \textit{ (rounded to } 10^{-4})$

c. $1.002s - 0.998s - 0.004s \textit{ (rounded to } 10^{-3})$

Example 5: Watch for numbers ending with zero!

$10 + 0.0110 = \textit{calculator returns } 10.0110$

$10: \textit{the uncertain digit appears in the } 10^{1} \textit{ place}$

$0.0110: \textit{the uncertain digit appears in the } 10^{-4} \textit{ place}$

$10 + 0.0110 - 10 \quad \textit{round answer to the } 10^{1} \textit{ place}$

Rationale: The uncertainty in the measured number 10 is ±1. The uncertainty alone in the first number (10) is greater than the entire second number (0.0110).

Problems

1. How many significant figures in the following numbers:

1. ______ 1,245m
2. _______ 0.030m
3. _______ 10,000m
4. _______ 1.340×10^{23}m
5. _______ 3.02003×10^{14}m
6. _______ 0.0000001m
7. _______ 1,000.
8. _______ 0.10000010

2. Convert the following numbers into standard scientific notation:

a. $96.3 \times 10^4\ g$ ____________________

b. $0.05 \times 10^{23}\ s$ ____________________

c. $123 \times 10^{-7}\ m$ ____________________

Problems 3-6: Perform the following calculations and record your answers in the proper number of significant figures and units.

3. $0.6030s + 0.82s =$

4. $4.1m + 0.3789m - 153.22m =$

5. $\frac{0.307g}{(1.0\times 10^{-3})ml} =$

6. $\sqrt[3]{5.33\times10^{5}m} =$

Part II: Simple Metric Conversions and Consistent Units

Section 1: Metric Conversions

Fill in the chart below with the metric conversion units. Memorize the ones in bold type! An example is given:

Prefix	**Symbol**	**Power of 10**	**Meaning**
deci-	d	10^{-1}	10 times smaller than base unit
centi-			
milli-			
micro-			
nano-			
kilo-			

Make the following conversions – preserve the number of significant figures in the answer!

1. 450*nm* ____________________ *mm*

2. 34*km*___________________*cm*

3. 43 000*mm* _____________________k*m*

4. 4.0 x 10^6 *nm* ________________µ*m*

5. 3.98 x 10^{-3} km___________________µ*m*

6. 456*mm* ____________________*km*

7. 136 000*m* _____________________*km*

8. 4.89 x 10^{12} *mm* _____________*km*

9. 2.68 x 10^6 *m* __________________*km*

10. 456 000 µ*m* _______________*mm*

Unit Multiplication – Dimensional Analysis – Factor Labeling

Units:
In the world of mathematics numbers often exist as abstract and unit-less entities. However, in the world of physics and chemistry where numbers are based upon experimentation and measurement all numbers are based in a physical reality. ***As a result, every number consists of two important parts.*** The first is a **magnitude** and the second equally important part is a **unit**. It is the unit that gives physical, real-world meaning to the number. We never write one without the other!

Examples: Note that these are all "equivalence statements"!

12 ***inches*** in one ***foot***
365 ***days*** in one ***year***
7 days in one ***week***
1.0 x 10^9 ***bytes*** in one ***gigabyte***

Derived Units and Calculations
Many of the common units we use are actually derived units that result from performing mathematical operations on the basic units. ***When performing mathematical operations the units are treated and manipulated as if they were algebraic variables.*** Here are a few examples:

Area = (length -**m**) x (width -**m**) = $\mathbf{m^2}$

Volume = (length -**m**) x (width -**m**) x (height - **m**) = $\mathbf{m^3}$

Velocity = (distance traveled – **m**)/(time -s) = **m/s**

Density = (mass – **g**)/(volume - mL) = **g/mL**

Unit Conversions
It is often necessary to convert from one system of units to another. The most efficient way to do this is using a process known as "*unit multiplication*", "*factor labeling*" or "*dimensional analysis*".

"**goal posting"**
One useful version of this method is called "goal posting". ***Step 1:*** Draw a "*goal post* "with the horizontal bar extending on each side. ***Step 2:*** Place the original number and unit to the left. Place the final unit on the right. ***Step 3:*** Move the original unit (cm) from the top left (*numerator*) to the bottom of the conversion factor (*denominator*). Now there is no confusion about which form of the conversion factor you will use. If you have done this correctly the original units on the top (*cm*) will be cancelled by the same unit in the denominator of the conversion factor.

Example: Consider a car traveling at ***35 m/s*** in the metric system. What would be the corresponding length in the English system (***miles / hour***)?

Solution: Note that velocity is a derived unit and has two units that must be converted: Length (Meters □ miles) and Time (seconds □ Hours).

Step 1: The derived unit has consists of two different units – one in the numerator and one in the denominator. Place the numerator unit *together with the number* on the "top" of the goalpost. Place the denominator units on the "bottom" of the goal post.

Step 2: The top unit will be moved down and to the right, the bottom unit will be moved up and to the right.

35 m	1.094 yds	1 mile	60 s	60 minute	78 miles
s	1 m	1760 yds	1 minute	1 hour	hour

Note that the only unit not cancelled in the numerator is miles. The only unit not cancelled in the **denominator** is hours. This gives us the final unit of miles/hour which the correct unit for the result.

Dimensional Analysis Practice Problems
(These are TOUGH---but give them a try and see what you can do!)

1. I have 470 milligrams of table salt, which is the chemical compound NaCl. How many liters of NaCl solution can I make if I want the solution to be 0.90% NaCl? (9 grams of salt per 1000 grams of solution).

The density of the NaCl solution is 1.0 g solution/mL solution

2. Light travels 186 000 miles / second. How long is a light year in meters? (1 light year is the distance light travels in one year)