Unit 1: Introduction to Chemistry's Measurement and Problem Solving
Part 1 – Notes: Identifying Significant Figures

Objectives: Identify the number of significant figures in a measurement.
Compare relative uncertainties of different measurements.
Text Reference: Section 3.2 (part) – pages 54–58(top)

Any measurement involves an estimate and that means that there is some uncertainty in that measurement. When a measurement is recorded, it includes all the digits that are certain plus the single estimated digit. **These certain digits plus one uncertain (estimated) digit are referred to as *significant figures.*** The certainty of a particular measurement is indicated by the number of ***significant figures*** recorded in that measurement.

*****KEY POINT***** The further the estimated digit lies to the right in a measurement, the less relative uncertainty there is, so the more reliable the measurement.

Look at the measurements: 34.3 mL 34 mL 34.46 mL 34.458 mL

Which of the measurements has the *least* amount of relative uncertainty?
Which of the measurements is the least certain (least reliable)?

Rules for Identifying Significant Figures

1. *Nonzero Integers.* Nonzero integers ***always*** count as significant figures. For example, the measurement 1479 m has 4 nonzero integers, all of which count as significant figures. 1479 m = 4 s.f.

2. *Zeros.* There are three classes of zeros:
   a. *Leading zeros* are zeros that *precede* all of the nonzero digits. These leading zeros *never* count as significant figures. For example, in the measurement 0.0025 g, the three zeros simply indicate the position of the decimal point. The measurement has only two s. f., the 2 and the 5.

   b. *Captive zeros* are zeros that fall *between* nonzero digits. They *always* count as significant digits. For example, the measurement 4.005 L has four significant figures.

   c. *Trailing zeros* are zeros at the *right* end of the number. They are significant only if the number is written with a decimal point. For example, the measurement 3.5000 g has a decimal point and the trailing zeros are significant; the measurement contains 5 s. f. The measurement 100. m is written with a decimal and contains 3 s.f while the measurement 100 m does not have a decimal and therefore contains only 1 s.f. (Note, a bar may also be placed over a zero to indicate that it is significant.)

3. *Exact numbers.* Often calculations involve numbers that were not obtained using measuring devices but were determined by counting: 7 beakers, 3 apples, 8 molecules. Such numbers are called *exact numbers*. They can be assumed to have an unlimited number of significant figures. Exact number may also arise from definitions. For example, 1 inch is defined as 2.54 cm; neither 2.54 or 1 limits the number of significant figures when it is used in calculations.

Examples: How many significant figures in the following measurements and where is the uncertainty?

6072 m	0.004 040 g	6070 L	6070. L	3000. g
0.004 04 L	49 500 m	3000 m	372.00400 L	0.000000004 m

Unit 1: Introduction to Chemistry's Measurement and Problem Solving
Part 1 – Assignment: Identifying Significant Figures

Indicate the number of significant figures in each measurement. Then circle the uncertain digit.

1.	679.98 g	2.	50 002 L	3.	4.000 m	4.	0.000 9 m
5.	3001 g	6.	3000 m	7.	0.000 400 m	8.	0.000 06 g
9.	10 001 m	10.	10 000 L	11.	100.001 g	12.	0.010 10 mg
13.	0.809 00 nm	14.	10 030 L	15.	0.000 300 4 dam	16.	34 578 pg
17.	40 908 mL	18.	5.80 hL	19.	0.009 00 km	20.	3000. kg

Unit 1: Introduction to Chemistry's Measurement and Problem Solving
Part 2 – Notes: Mathematical Operations and Significant Figures – 1

Objectives: Correctly record an answer to a single-step math problem with the appropriate number of significant figures.
Determine, utilize, and explain the rule for the number of significant figures in a sum or a difference.
Determine, utilize, and explain the rule for the number of significant figures in a product or quotient.
Explain how the rules for determining the number of significant figures in single-step math problems are derived.

Text Reference: Section 3.2 – pages 54-62

Adding and Subtracting Significant Figures

When adding or subtracting measurements, the **sum or difference can only be as certain as the least certain measurement.**

Remember, properly recorded measurements are recorded using significant figures, all certain digits plus a single uncertain estimated digit.

Example 1: Add the following and report the answers to the correct number of significant figures.

(A) 23.0042 m + 9.0 m (B) 12.7 g + 3.3 g (C) 19.38 cm + 2.4 cm (D) 9.1 m + 11.01 m + 10 m

The position of the uncertain digit with relation to the decimal point in the least certain measurement determined the position of the uncertain digit with relation to the decimal point in the answer. For our purposes, the uncertain digit should be rounded up or rounded down, just as in math.

In your own words, what's the rule for adding and subtracting?

Multiplying and Dividing Significant Figures

When multiplying and dividing measurements, the **answer may only contain as many significant figures as does the measurement used with the least amount of significant figures.**

Example 2: Find the products of the following and report the answers to the correct number of significant figures.

(A) 42.0 m x 4.0 m (B) 1.78 cm x 0.05 cm (C) 1001 mm x 40 mm (D) 1001 mm x 40. mm

Note that unlike addition and subtraction, the position of the uncertain digit with relation to the decimal point in the answer does not have to be in the same position as the position of the uncertain digit in any measurement.

In your own words, what's the rule for multiplying and dividing?

Unit 1: Introduction to Chemistry's Measurement and Problem Solving
Part 2 – Assignment: Mathematical Operations and Significant Figures – 1

Solve the following. Show all work and intermediate steps and record the answers to the correct number of sig figs.

1. 23.67 m + 45.7 m

2. 67.2 g – 30.4 g

3. 56 m x 20.4 m

4. 1020 m ÷ 20 s

5. 10 000.0 m + 1 m

6. 10 000 m + 1 m

7. 1020 m ÷ 20.0 s

8. 5.67 m + 10.0 m

9. 5.67 m + 10 m

10. 1000 g + 5 g

11. 15 m x 5 m

12. 15 m x 5.0 m

13. 3500 m^2 ÷ 700 m

14. 300 g – 6 g

15. 4000.9 g + 1 g

16. 4500.5 m – 0.5 m

17. Water with a mass of 35.4 g is added to an empty flask with a mass of 87.432 g. The mass of the flask and the water is 146.72g after a rubber stopper is added. Express the mass of the stopper to the correct number of significant figures.

18. Criticize this statement: "When two measurements are added together, the answer can have no more significant figures than the measurement with the least number of significant figures."

Unit 1: Introduction to Chemistry's Measurement and Problem Solving
Part 3 – Notes: Mathematical Operations and Significant Figures – 2

Objectives: Correctly record an answer to multiple-step math problems, averages, and conversions with the appropriate number of significant figures. Additional objective – see objectives from Unit 1 Part 2.

Text Reference: Section 3.2 – pages 54-62

Rule 1: ***Multiplication, Division, and Exponent Operations, only:***
When multiplying, dividing, and using exponents only, perform all operations first and then apply the rule for significant figures to the answer.
Example 1: $(2.48\ m)^2\ (2.58\ m) \div 5.78\ m$

Rule 2: ***Addition and/or Subtraction in combination with other operations:***
When addition or subtraction operations are used in conjunction with other operations, follow the algebraic order of operations, applying the rules of significant figures after addition/subtraction is completed. Then perform the multiplication/division as necessary and then apply the rules for significant figures, again.
Example 2: (4.238 m + 5.97 m) ÷ 2.5 s

Example 3: [(4.56 m + 5.679 m) (2.00 m + 9 m)] ÷ 27.9 m

Rule 3: ***Numbers with NO uncertainty:***
All conversion fractions, constants, and numbers with no uncertainty are treated as if they have an infinite number of significant figures. More on this part later.
Example 4: Convert 14.75 hours to seconds.

Rule 4: ***Averages and uncertainty:***
When finding the average (mean) of measurements, the mean cannot be more certain than the least certain measurement. Also, keep in mind that you are adding and dividing (two different significant figure rules).
Example 5: Calculate the mean of the following measurements: 2.784 g and 9.863 g.

Unit 1: Introduction to Chemistry's Measurement and Problem Solving
Part 3 – Assignment: Mathematical Operations and Significant Figures – 2

Perform the following operations and record your answers to the correct number of sig figs. Show all intermediates.

1. 46.98 m ÷ (3.5 s x 5.78 s)
2. (56.7 s + 12 s) (35.46 s)
3. (5678.1 m x 2.0 m) ÷ 5 m
4. (12.3 m + 15 m) ÷ 2 s
5. $43400\ m^3 \div (4.334\ m + 44.0002\ m - 0.982\ m)$
6. Convert 5 789 seconds to days.
7. Average: 5.0 g and 7.89 g
8. Average: 9.1 g, 11.0 g, and 10 g
9. Average: 2.345 g and 8.349 g
10. Average: 5.00 g, 55.0 g, and 200 g

Unit 1: Introduction to Chemistry's Measurement and Problem Solving
Part 4 – Notes: Scientific Notation

Objectives: Convert a number in standard notation to correct scientific notation, and vice versa.
Identify the part of a number in scientific notation and state appropriate values for each part.
Distinguish between qualitative and quantitative measurements.
Perform simple math problems with number in scientific notation without a scientific calculator.

Text Reference: Section 3.1 – pages 51-53

Qualitative:

Quantitative:

Measurements in chemistry range from incredibly tiny to almost unimaginably large. The number of iron atoms that would fit side by side on a line 1 cm in length is more than 80 million. The number that could be packed into a cube with a volume of 1 cm^3 is $(80 \text{ million})^3$ – or about 500 thousand billion billion! The size of an iron atom is almost unimaginably small.

Scientists use large and small numbers frequently, but with so many zeros, problems are frequently encountered. Writing large numbers of zeros is time consuming and there is an increased chance of an error being made when writing many zeros. Scientists handle large and small numbers using ***scientific notation***, also called *standard exponential form*. A number written in scientific notation has the following appearance and parts:

$$a.bc \times 10^d$$

Part	**What is it?**	**Acceptable values**
a.bc		
x10		
d		

A **positive exponent** indicates. . .

A **negative exponent** indicates . .

Converting a number form standard form to scientific notation:

1. Write the decimal point directly to the right of the first nonzero digit, creating a number that is greater than or equal to one but is less than ten.
2. Count how many places from your new decimal to the original decimal place. This becomes the exponent.
3. If the number was larger than one, it remains a positive exponent. If the number was smaller than one, it becomes a negative exponent.

Keep in mind, this is not magic and it is not something new and different. All you have done is rewrite a number is different form!!!

Example 1: Write the following numbers in scientific notation.

(a) 12340000000000000 (b) 45060000000000 (c) 0.00000000000000000234 (d) 0.000000000002000015

Example 2: Convert the following from scientific notation to standard notation.

(a) $2.36\text{x}10^4$ (b) $2.654498\text{x}10^3$ (c) $3.659\text{x}10^{-7}$ (d) $6.019\text{x}10^{-3}$

Changing the Form of Exponential Numbers

You can ***increase*** either the coefficient or the exponent of a number scientific notation by any factor without changing the overall *value* of the number as long as we ***reduce*** the other portion by the same factor. In other words, as one part increases, the other part decreases. (Remember: Elevator go up, elevator go down.)

Example 3: Change the following to correct scientific notation:

(a) $10 \text{x} 10^{7}$ (b) $0.0050 \text{x} 10^{6}$

(c) $303 \text{x} 10^{-4}$ (d) $0.012 \text{x} 10^{-7}$

Example 4: Change the following coefficients so they match the exponents that are given.

(a) $4.98 \text{x} 10^{6}$ = _______________ $\text{x} 10^{3}$ (b) $5.686 \text{x} 10^{8}$ = _______________ $\text{x} 10^{10}$

(c) $6.37 \text{x} 10^{-4}$ = _______________ $\text{x} 10^{-7}$ (d) $3.587 \text{x} 10^{-7}$ = _______________ $\text{x} 10^{-5}$

Multiplication and Division with Scientific Notation

To multiply numbers in scientific notation, you need to multiply the coefficient portions and the exponent portions separately. To multiply the exponential portions, the exponents are ***algebraically added***. *(Be aware of signs!)* For example:

$(2.0 \text{x} 10^{3}) \text{ x } (4.0 \text{x} 10^{2}) =$

To divide number in scientific notation, you divide the coefficients portions and the exponent portions separately. To divide the exponential portions, the exponents are ***algebraically subtracted***. *(Be aware of the signs!)* For example:

$(6.0 \text{x} 10^{5}) \div (2.0 \text{x} 10^{3}) =$

Often, when dividing and multiplying, the product or quotient will be a number greater than or equal to ten or less than one. When this happens, you need to adjust the coefficient so it is a number greater than or equal to 1 but less than 10. You then need to adjust the exponent so the value of the answer is not changed. Your final answer always needs to be in proper scientific notation.

Addition and Subtraction with Scientific Notation

When we add or subtract numbers in exponential notation, *the exponents must be the same*. This rule is the same as making sure the decimals of your numbers you add or subtract are aligned. The answer is the sum or difference of the coefficients with the same exponent as each number in the problem. Remember, the final answer needs to be in correct scientific notation form.

Example 5: Perform the following operations.

(a) $(3.0 \text{x} 10^{6})(5.0 \text{x} 10^{3}) =$ (b) $(4.25 \text{x} 10^{-5})(7.7 \text{x} 10^{2}) =$

(c) $(5.0 \text{x} 10^{3}) \div (6.75 \text{x} 10^{5}) =$ (d) $(2.5 \text{x} 10^{-5}) \div (5.6 \text{x} 10^{-2}) =$

(e) $(3.45 \text{x} 10^{6}) + (4.56 \text{x} 10^{7}) =$ (f) $(2.45 \text{x} 10^{-4}) + (3.75 \text{x} 10^{-2}) =$

(g) $(2.98 \text{x} 10^{5}) - (3.25 \text{x} 10^{4}) =$ (h) $(9.68 \text{x} 10^{-3}) - (4.67 \text{x} 10^{-4}) =$

Unit 1: Introduction to Chemistry's Measurement and Problem Solving
Part 4 – Assignment: Scientific Notation

1. Which of the following are NOT in proper scientific notation:

a.	$2.3x10^{-4}$	b.	$45.7x10^{6}$	c.	$0.23x10^{3}$	d.	$1.234x10^{-3}$
e.	$0.004x10^{-9}$	f.	$1x10^{8}$	g.	$12.3x10^{6}$	h.	$0.48x10^{-7}$

2. Write the following numbers in correct scientific notation.

a. 1010000000000000000000000000

b. 120

c. 0.00000000000987

d. 0.00000000005600007

e. 6.87

f. one billion

g. 0.00000090

h. 700200000

Perform the following operations ***without a scientific calculator***. Show your work, any intermediate steps, and clearly indicate your final answer.

3. $(6.883x10^{-4})$ $(7.883x10^{-7})$ =

4. $(3.9857x10^{7}) \div (7.395x10^{11})$ =

5. $(4.23x10^{8}) \div (7.65x10^{-3})$ =

6. $7.993x10^{-8} + 8.883x10^{-10}$ =

7. $8.997x10^{-7} – 9.987x10^{-6}$ =

8. $7.856x10^{6} + 9.546x10^{8}$ =

9. Identify the following as qualitative or quantitative:

a. A flame is hot.

b. A candle has a mass of 90 g.

c. Wax is soft.

d. A candle's height decreases 4.2 cm/hour.

10. Rank these numbers from smallest to largest:

(a)	$5.3x10^{4}$	(b)	$57x10^{3}$	(c)	$4.9x10^{-2}$
(d)	0.0057	(e)	$5.1x10^{-3}$	(f)	$0.0072x10^{2}$

__________ __________ __________ __________ __________ __________

smallest largest

Unit 1: Introduction to Chemistry's Measurement and Problem Solving
Part 5 – Class: Let's Start Problem Solving

Objectives: Develop and utilize a method for problem solving.
Text Reference: Section 4.1 – pages 83-88

How would you solve the following problems? Show ALL work, steps, units, etc.

1. A car travels 270 miles in 5 hours. How many hours will it take the car to go 150 miles. (Assume it travels at a constant velocity.)

2. Convert 3000 meters into kilometers.

3. The density of a particular material is 2.3 g/cm^3. What is the volume of 800 g of this material?

4. Let x = number of meters and y = number of kilometers. An equation using the symbols x and y and the number 1000 which expresses the relationship between the number of meters and the number of kilometers is:

   (A) $1000x = y$ (B) $1000y = x$ (C) $x + y = 1000$ (D) $xy = 1000$

5. Information: M is a concentration unit equal to *moles ÷ volume (in Liters).*
   How many grams of solid sodium chloride are needed to prepare 2.5 L of a 0.5M solution of sodium chloride? The mass of 1 mole of sodium chloride is 58.5 g.

6. a. Using only the following set of circles, show how you would represent the operation 6 ÷ 2:

   O O O O O O

   b. Based on the definition of division, draw a picture that justifies the answer to the following problem. What is the answer to the following problem: 2 ÷ 1/2?

7. a. Consider the ratio the ratio $40 / 5 gallons. When we divide 40 by 5 we get 8. What does 8 represent in this case?

   b. Now consider 5 gallons / $40. 5/40 = 0.125. What does the 0.125 represent in this case?

8. A certain liquid costs $5 per gallon. What does the following equation represent: $40 / ($5 / 1 gallon)?

   What is the answer to the above problem? When you perform the division, what is the label on the number you obtain?

Unit 1: Introduction to Chemistry's Measurement and Problem Solving
Part 5 – Assignment: Let's Start Problem Solving

Read pages 83 – 88 in your textbook. Then answer the following questions (from your text).

1. Evaluate the following and record the answer.
   a. $89^2 + 17^3 + 5^6 =$

   b. $(35 + \sqrt{529}) / (2.9\text{x}10^{17}) =$

   c. The volume of a sphere with radius *r* is given by $(4/3)\pi r^3$. What is the volume of a sphere with a radius of 3.50 cm?

   d. Find the number of atoms in 7.00 g of gold if an atom of gold has a mass of $3.271\text{x}10^{-22}$ g.

2. The density of silicon is 2.33 g/cm^3. What is the volume of a piece of silicon that has a mass of 62.9 g?

3. A small piece of gold has a volume of 1.35 cm^3.
   a. What is the mass of the gold piece, given that the density of gold is 19.3 g/cm^3?

   b. What is the value of the gold piece if the market value of gold is $11/g?

4. A watch loses 0.15 seconds every minute. How many minutes will the watch lose in one day?

5. Earth is approximately $1.5\text{x}10^8$ km from the sun. How many minutes does it take light to travel from the sun to Earth? The speed of light is $3.00\text{x}10^8$ m/s.

Use the following procedure when solving mathematical problems in this chemistry class.

1. Read the whole problem. (Don't just start writing down numbers.)
2. Write down what you are given. (It is important that you known what you have to work with.)
3. Identify and write down what you wish to find. (You can't solve it if you don't know what you're looking for.)
4. Is there an equation that relates the given information to the unknown? Write the formula. (More than one formula may be required. Write all required formula.)
5. Substitute the known values into the equation. Make sure you are including units.
6. Write your final answer using the correct number of significant figures and a correct unit. Is it clearly recognizable as the answer or should you circle or box it?
7. Check to see that the answer makes sense. If it doesn't make sense, check over your work.

Unit 1: Introduction to Chemistry's Measurement and Problem Solving
Part 6 – Notes: The International System of Units

Objectives: List the seven basic units in the metric system, what they measure, and the abbreviations for their units.
Compare units in the metric system with units in the English system.
List the required metric prefixes, their abbreviations, and their numerical values.
Distinguish between mass and weight.
Convert between temperatures in Celsius and Kelvin scales.

Text Reference: Section 3.3 – pages 63-67 and Section 3.5 – pages 74-75

You are familiar with most common English units. ***Within the English system***, you need to know the following:

1 foot = 12 inches	1 minute = 60 minutes
1 yard = 3 feet	1 hour = 60 minutes
1 mile = 5280 feet	1 day = 24 hours
1 year = 365 days	

In the realm of science, the metric system is preferred for two different reasons:

1. Metric units for complex quantities are defined in terms of units for simpler quantities.
2. The metric system has the same numerical base as the decimal system. Every unit in the metric system is ten times the size of the next smaller unit. The units in the metric system are based on the number 10.

There are seven basic quantities that may be measured. The unit attached to a number indicates what quantity is being measured. The seven basic quantities and their base units are as follows. You need to know these.

Quantity	***Modifiable Unit***	***Unit Abbreviation***
length	meter	m
mass	gram	g
time	second	s
temperature	kelvin	K
amount of substance	mole	mol
electric current	ampere	A
luminous intensity	candela	cd

In the metric system, ***every unit has a quantity that can be modified by prefixes***. There are specific prefixes used to modify the units and these prefixes have specific prefixes. You need to know the following prefixes and their values:

Prefix	***Abbrev.***	***Meaning (Words)***	***Meaning (Math)***	***Or, for the unit meter. . .***
mega-	M	one million	1 000 000	1 Mm = 1 000 000 m
kilo-	k	one thousand	1 000	1 km = 1 000 m
hecto-	h	one hundred	100	1 hm = 100 m
deka-	da	ten	10	1 dam = 10 m
deci-	d	one-tenth	0.1	1 m = 10 dm
centi-	c	one-hundredth	0.01	1 m = 100 cm
milli-	m	one-thousandth	0.001	1 m = 1 000 mm
micro-	μ	one-millionth	0.000 001	1 m = 1 000 000 μm
nano-	n	one-billionth	0.000 000 001	1 m = 1 000 000 000 nm
pico-	p	one-trillionth	0.000 000 000 001	1 m = 1 000 000 000 000 pm

Mass versus Weight:

Temperature:

Unit 1: Introduction to Chemistry's Measurement and Problem Solving
Part 6 – Assignment: The International System of Units

1. As you climbed a mountain and the force of gravity decreased, would your weight increase, decrease or remain the same? How would your mass change?

2. List these units in order from largest to smallest: 1 dm^3 1 μL 1 mL 1 L 1 cL 1 dL

________	________	________	________	________	________
largest					smallest

3. The boiling point of elemental argon is 87 K. What is argon's boiling point in degrees Celsius?

4. Order these lengths from smallest to largest: cm μm km mm m nm dm pm

________	________	________	________	________	________	________	________
smallest							largest

5. From what unit is a measure of volume derived?

6. Astronauts in space are said to have apparent weightlessness. Explain why it is incorrect to say they are massless.

7. Which would melt first. germanium with a melting point or 1210K or gold with a melting point of 1064^{o}C?

8. Which is larger? Circle the larger of the two measurements.

   a. 1 centigram or 1 milligram

   b. 1 liter or 1 centiliter

   c. 1 calorie or 1 kilocalorie

   d. 1 millisecond or 1 centisecond

   e. 1 cubic millimeter or 1 cubic decimeter

   f. 1 microliter or 1 milliliter

Unit 1: Introduction to Chemistry's Measurement and Problem Solving
Part 7 – Notes: Unit Conversions

Objectives: Perform one-step and multi-step units conversions with various units using dimensional analysis.
Text Reference: Section 4.2 – pages 89-95 and Section 4.3 – pages 97-100

Let's look at converting between different units of length.
For all your work in chemistry, you will always perform conversions using UNIT ANALYSIS. Unit analysis is a way you will solve problems; it is also called dimensional analysis, unit factoring,, the factor label method, and has a few other names.

When you solve problems using unit analysis, you begin with a known quantity and use a ratio in a fraction form to allow you to convert from your known unit to your unknown unit. When working with unit analysis, it is essential that you keep track of and record all units.

To solve a problem, you need to know the ratio that exists between your known and your unknown quantity. The ratio is an equality with your known unit on one side of the equal sign and the unknown unit of the other and the numbers in from of the units that make the equality true. For example, if I wanted to convert between meters and centimeters, I would need the following equality: 1 m = 100 cm. In some cases, you may need to use multiple ratios to solve a problem. Let's see how to use this method of problem solving.

Example 1: Convert 3.654 days into its equivalent quantity in hours.

Example 2: Convert 124.57 hours into its equivalent quantity in days.

Example 3: Convert 376 cm into its equivalent quantity in meters.

Example 4: Convert 475 μm into its equivalent quantity in hectometers.

Example 5: Convert 568 meters into its equivalent quantity in miles. Note 1 km = 0.62 mi.

Be careful when canceling units. Look at the following examples.

Example 6: Convert 0.748 m^2 into its equivalent quantity in cm^2.

Example 7: Convert 17564 mm^3 into its equivalent quantity in m^3.

Example 8: Convert 1.67 mi^2 into its equivalent quantity in m^2.

Remember: MEASUREMENT = QUANTITY + UNIT
While you pursue your work in chemistry, remember, you are using measurements. This means you will be working with values that have units attached. Keep your units with your numbers. Units will have to cancel in order for you to obtain a valid answer. In chemistry, ***<u>NO NAKED NUMBERS</u>***!!!

Unit 1: Introduction to Chemistry's Measurement and Problem Solving
Part 7 – Assignment: Units Conversions

Solve the following using unit analysis. Show all work, units, set-ups, etc. Convert:

1. 0.0004 kg to mg

2. 0.00000000006 Mm to km

3. 0.452 kL to cL

4. 256 dam to km

5. 3.5 hours to seconds

6. 3.76 years to minutes

7. 321 000 mm to pm

8. 12 385 884 seconds to years

9. 84 937 μL to nL

10. 0.38 km^2 to mm^2

11. 57 561 cm^3 to Mm^3

12. 0.00057 hm^3 to dm^3

Unit 1: Introduction to Chemistry's Measurement and Problem Solving
Part 8 – Notes: Density and Problem Solving

During your tour through chemistry, you will come upon numerous problems that you will be required to solve. You need to solve the mathematical problems in a logical fashion, with steps based on the scientific method. Use the following procedure when solving mathematical problems in this chemistry class.

1. Read the whole problem. (Don't just start writing down numbers.)
2. Write down what you are given. (It is important that you known what you have to work with.)
3. Identify and write down what you wish to find. (You can't solve it if you don't know what you're looking for.)
4. Is there an equation that relates the given information to the unknown? Write the formula. (More than one formula may be required. Write all required formula.)
5. Substitute the known values into the equation. Make sure you are including units.
6. Write your final answer using the correct number of significant figures and a correct unit. Is it clearly recognizable as the answer or should you circle or box it?
7. Check to see that the answer makes sense. If it doesn't make sense, check over your work.

DENSITY

What weighs more, a pound of feathers or a pound of steel?
You know the answer is "neither," both weigh the same: 1 pound. But if your instinct was to say "steel," you were probably thinking of the quantity of ***density***. Steel is more dense than feathers. What does that mean?

Density is defined as *the mass per unit volume of a substance:* Density = mass / volume
D = m/V

The dimensions (combination of units) of density involve a mass unit divided by a volume unit, such as g/mL or g/cm^3. So, to get the density of an object, simply divide it mass by its volume.

NOTE: *The density of pure water is about 1.00 g/mL or 1.00 g/cm^3.*

Let's now solve several density problems using the problem solving steps from above. Remember, use all problem solving steps.

Example 1: Calculate the density, in g/mL of wood in a desk if it has a mass of 30.0 kg and a volume of 35.0 L?

Example 2: Calculate the density of a rectangular metal bar that is 5.00 cm long, 20.0 mm wide, and 0.0100 m thick and has a mass of 23.0 g. Calculate the density in g/cm^3.

Example 3: Calculate the mass, in kilograms, of 254 mL of mercury. The density of mercury is 13.6 g/mL.

Substances generally expand when heated, and the resulting change in volume causes some change in density. Within reasonable temperature changes, the density of a substance is fairly constant. For example water varies from 0.99979 g/mL at 0^{o}C to 1.0000 g/mL at 4^{o}C to 0.95838 g/mL at 100^{o}C. You will often ignore such changes.

Density is an ***intensive property***, useful in identifying substances. For example gold may be distinguished from iron pyrite (fool's gold) by their differing densities. The density of gold is 19.3 g/mL while that of iron pyrite is only 5.0 g/mL.

The relative densities determine whether an object will float in a given liquid in which it does not dissolve. An object will float if its density is less than the density of the liquid in which it is placed. An object will sink if its density is greater than the density of the liquid in which it is placed. For example, the density of water is 1.00 g/mL and a piece of wood placed in the water has a density of 0.80 g/mL. The wood will float in the water since it has a lower density.

Example 4: Will the wood in example 1 sink or float in water? Will it sink or float in mercury?

Specific Gravity:

Unit 1: Introduction to Chemistry's Measurement and Problem Solving
Part 8 – Assignment: Density and Problem Solving

Solve the following problems. Show all work, set-ups, units, steps, etc. Use correct significant figures. BE NEAT.

1. A cube of gold that is 2.000 cm on each side has a mass of 154.4 g. What is the density of gold?

2. A piece of silver (density = 10.5 g/mL) dropped in water and displaces 21.56 mL. What is the mass, in grams, of the silver?

3. Concentrated sulfuric acid has a density of 1.84 g/mL. Calculate the mass, in grams, of 1.00 L of sulfuric acid.

4. Calculate the density, in g/cm^3, of a block of wood that has a mass of 75.0 kg and the following dimensions: length = 25.0 cm, width = 0.100 m, and depth = 550.0 mm.

5. A 1.0000 kg of metallic osmium, the "heaviest" element known, occupies a volume of 44.5 cm^3. Calculate the density of osmium in g/cm^3.

6. A student finds a piece of metal she thinks is aluminum. In the lab, she determines the metal sample has a mass of 612 g and a volume of 245 mL. The density of aluminum is 2.7 g/mL. Is the sample aluminum or not? Explain.

7. Would the density of a person be the same on the surface of the moon as it is on earth? Explain.

8. Why doesn't specific gravity have a unit?

9. If ice were more dense than water, it would certainly be easier to pout water from a pitcher of ice cubes and water. Can you imagine situations where this density switch would be a greater problem?

10. The mass of a cube of iron is 355 g. Iron has a density of 7.87 g/cm^3. What is the mass of a cube of lead that has the same dimensions as the iron? Lead's density is 11.4 g/cm^3.

Unit 1: Introduction to Chemistry's Measurement and Problem Solving
Part 9 – Notes: Accuracy and Precision

Objectives: Define, explain, and distinguish between accuracy and precision.
Calculate the absolute and relative error and the absolute and relative deviation of a set of measurements.
Text Reference: Section 3.2 – pages 54-55

Two important factors to consider in measurement are ***precision*** and ***accuracy***. **They are NOT the same thing.**

Precision indicates the ***reproducibility*** of a measurement.
Example: Imagine you took three readings on a thermometer to try to determine the boiling point of pure water at sea level. The results of your trials were 96.8^{o}C, 96.9^{o}C, and 97.2^{o}C. Note how close the measurements are to one another. The measurements show a high degree of reproducibility, they are ***precise***.

Accuracy indicates ***how close*** a measurement is to its ***accepted value***.
Consider the boiling point of pure water at sea level in the example above. It is an accepted fact that pure water at sea level boils at exactly 100^{o}C. The three measurements above, although precise, are not close to the accepted value. They are not accurate.

Determining Accuracy – Calculating Percent Error

Recall, accuracy deals with how close measurements are to the accepted or actual value, so you will be comparing the measurements to the accepted value.

Formula for Absolute Error:
(also called experimental error)

Formula for Percent Error:
(also called relative error)

From the example above we have three different values: 96.8^{o}C, 96.9^{o}C, and 97.2^{o}C.

Calculate the percent error for *each* individual measurement:

Then calculate the average of those measurements:

You can also calculate the percent error for the mean of the measurements:
What is the mean of the measurements?

What is the percent error for the mean of the measurements?

Determining Precision – Calculating Deviation

Calculating how precise measurements are is a multi-step process. You will again use the three temperature readings for the boiling water example at the beginning of the notes to complete the steps that follow. Recall, precision deals with how close measurements are to one another, so you will be looking at comparing the measurements to the average.

STEP 1: Determine the **Mean** of the measurements.

STEP 2: Calculate the **Absolute Deviation** for **EACH** measurement. In this step we are determining how much each individual measurement deviates from the average.

STEP 3: Calculate the **Average Absolute Deviation.**

STEP 4: Express the measurement in terms of Mean ± Average Absolute Deviation.

STEP 5: Calculate the **Relative Deviation.** (Also referred to as **percent deviation**.)

STEP 6: Express the measurement in terms of **Mean ± Relative Deviation**.

Unit 1: Introduction to Chemistry's Measurement and Problem Solving
Part 9 – Assignment: Accuracy and Precision

On a **separate sheet of paper**, answer the following questions. Be sure to show all work, set-ups, steps, answers, etc. Be sure to keep strict attention to significant figures.

1. In your own words, describe the difference between accuracy and precision.

2. Why is the percent error of a measurement always positive?

3. Comment on the accuracy and precision of the these basketball free-throwers:
   a. 99 of 100 shots are made.
   b. 99 of 100 shots hit the front of the rim and bounce off.
   c. 33 of 100 shots are made; the rest miss.

4. Three students made multiple massings of a copper cylinder, each using a different balance. The correct mass of the cylinder had been previously determined to be 47.32 g. Describe the accuracy and precision of each student's measurements.

Mass of Copper Cylinder

	Lisa	Lamont	LeighAnn
Mass 1	47.13 g	47.45 g	47.95 g
Mass 2	47.94 g	47.39 g	47.91 g
Mass 3	46.83 g	47.42 g	47.89 g
Mass 4	47.47 g	47.41 g	47.93 g

5. Thinking there is a problem with one of the minting machines, a US mint quality control inspector took the masses of 10 new pennies. The recorded masses were:

3.112 g	3.129 g	3.093 g	3.089 g	3.109 g
3.089 g	3.094 g	3.085 g	3.131 g	3.092 g

   a. Calculate the mean.
   b. Calculate the average absolute deviation. (First find the absolute deviation for each mass.)
   c. Calculate the relative deviation.

   According to the US mint, the mass of a new penny should be 3.160 g.

   d. Calculate the absolute error of the mean of the massing of the pennies. Use the mean value.
   e. Calculate the percent error of the mean of the massing of the pennies. Use the mean value.

6. Calculate the percent error for the measurements in each of the following conditions.
   a. The density of an aluminum block was determined in an experiment as 2.64 g/mL. The actual density of aluminum is 2.70 g/mL.
   b. The experimental determination of iron in an iron ore was 16.48%. The actual value of iron present in the ore is 16.12%.
   c. A balance measures a 1.0000 g mass as 0.9981 g.