Unit 1: Scientific Measurement

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Honors Chem

Unit 1 Learning Objectives:

By the end of the unit students will be able to…

- Convert between scientific notation and standard form
- Measure and convert between metric units and temperature units
- Determine how many significant figures a measure has, and use this knowledge to report the most accurate answer.
- Solve conversion problems using units
- Determine and calculate density

Monday	Tuesday	Wednesday	Thursday	Friday
August 20	**21**	**22**	**23** Introduction Syllabus & Safety Contract Mini Desalination Plant Activity	**24** Mini Desalination Plant Activity
27 Notes: Scientific Notation & Metric conversions	**28** Notes: Conversions continued	**29** Notes: How to read an apparatus, Temperature, and taking measurements	**30** Notes: Significant Figures	**31** Notes: Sig Figs & Density **Quiz: Lab Safety**
September 3 **No School! Happy Labor Day!**	**4** Lab: Measurements of mass and volume	**5** Lab: Measurements of mass and volume	**6** Review of Unit 1	**7** **Unit 1 Test** **HW packet Due**

Scientific Notation:

Scientists often measure vey large and very small numbers. For example, scientists can measure time periods as short as 0.000000000000001 seconds. But writing all of the zeros in this number can take a long time, and if written incorrectly can results in error. Therefore, scientists use ***scientific notation*** *to write these numbers more compactly and more accurately.*

In scientific notation the above example (0.000000000000001) would be written as 1×10^{-15}.

Scientific Notation consists of the following three parts:

$$1\ x\ 10^{-15} \longleftarrow \text{Exponent}$$

Decimal part   Exponential part

To convert from ***expanded notation to scientific notation:***

- Move the decimal to the right of the first integer.
- Count te number of spaces you moved the decimal (this number is the exponent)
- Numbers larger than 1 have a positive exponent
- Numbers smaller than 1 have a negative exponent

To convert from ***scientific notation to expanded notation:***

- Move the decimal the number of spaces indicated by the exponent
- Positive exponent indicates the number is greater than 1
- Negative exponent indicates the number is less that 1

Practice: Convert from expanded notation to scientific notation

1. 0.00393
2. 45,000
3. 10
4. 45.872

Practice: Convert from scientific notation to expanded form

1. 7.4×10^{5}
2. 6.34×10^{-7}
3. 6.90×10^{3}
4. 5.983×10^{9}

The Metric System

The metric system is based on powers of 10, so it is easy to use.

Prefix	**Symbol**	**Magnitude**
Giga	G	10^9
Mega	M	10^6
Kilo	k	10^3
Hecto	h	10^2
Deca	da	10
Basic unit	-	1
Deci	d	10^{-1}
Centi	c	10^{-2}
Milli	m	10^{-3}
Micro	μ (greek letter)	10^{-6}
nano	n	10^{-9}

The Standard Units, and the ones that we will use in this class are:

Quantity	**Unit**	**Symbol**
Length	meter	m
Mass	Kilogram *gram	kg *g
Time	Second	s
Temperature	Kelvin	K

Practice: Fill out the following chart based on the metric system

Unit name	**Symbol**	**magnitude**
Kilometer	*km*	*10^3 meters*
	μL	
centigram		
	nm	
		10^9 grams
nanometer		
	dm	

Converting Between Units:
<u>*Units should always be written in calculations*</u>. Using the metric system you can convert between metric units without changing the value of the measurement. Units can also be converted from the metric systems to other common units.

Converstion factors are consturcted from any two quantities known to be equivalent. For example 1 m is known to equal 100 cm. This known fraction or ration can be writen in <u>any</u> of the following ways.

$$1\ m = 100\ cm \qquad or \qquad \frac{1\ m}{100\ cm} \qquad or \qquad \frac{100\ cm}{1\ m}$$

Useful converstions to remember

1 inch = 2.54 cm
1 mile = 1.61 kilomters

1 gallon = 4.55 Liters
1 mL = 1 cm^3

1 kg = 2.204 lbs
1 oz. =28.35 g

1 hectare = 1000 m^2 = 2.47 acres

Practice: Convert each of the following
One Step Problems:
1. 548 mL into L
2. 10,000 kg into g
3. 8.5467 Mm into nm
4. 13.85 in into cm

Multi-Step Problems:
5. 0.33 cm into dm
6. 800 L into gallons
7. 176.8 ft into cm

Conversions with squared and cubed units:

In order for the units to cancel, the conversion factor <u>must</u> be squred or cubed in order for the units to cancel.

Example: Convert 200 cm^2 into m^2

Wrong:

Correct:

Practice: Covert the following

1. 3.4×10^8 cm^3 into m^3
2. 8.325 in^2 into cm^2

Real World Word Problems:

ONE STEP Conversion practice

1. Your refrigerator probably contains 8.5×10^{24} molecules of ozone destroying CFCs. If each CFC molecule eventually results in the destruction of 100,000 Ozone molecules, how many Ozones will be destroyed when your fridge releases all of its CFC from a landfill?

Multi STEP Conversion practice

1. If a bulb is rated 100 Watts, this means that it uses 100 Jouls of energy per second (100 J/s). An incandescent bulb typically can burn for 10,000 hours. How much energy will a 100 watt bulb use in its lifetime?
2. A 23 watt compact flourescent lamp (CFL) will burn for about 15,000 hrs. How much energy will it use in its lifetime?

3. Use the following information to answer the next questions.
   A kilogram of coal can create 2.4×10^4 kg of energy (electricity)
   A kilogram of caol also releases 2.93 kg of CO_2 greenhouse gas.

   a. An average incandescent bulb will use 3.6×10^6 kJ before it burns out. How many kg of CO_2 will be produced during the bulb's lifetime?

   b. An avrage compact fluorescent bulb (CFL) will use 2.4×10^6 kJ in its lifetime. How many kg of CO_2 will be produced?

Temperature

There are three scales for measuring temperature.

(1) **Fahrenheit (F),** Gabriel Fahrenheit (1686-1736)
   a. 0 F was set as the coldest temperature that could be obtained in a liquid in the lab, a mixture of salt, ice, and water; 32 F was set as the temperature of a mixture of ice and water.

(2) (2) **Celsius (C),** Anders Celsius (1701-1744)
   a. 0 ºC was set as the freezing point of water and 100 ºC was set as the boiling point of water.

(3) (3) **Kelvin (K),** Lord Kelvin, William Thomson (1824 -1907)
   a. 0 K (also called "absolute zero") is the coldest temperature that is possible. At 0 K all molecular motion stops.

Temperature Conversions

The following equations can be used to convert between the different temperature scales.

$$F = \frac{9}{5}(C) + 32 \qquad C = \frac{5}{9}(F - 32) \qquad K = C + 273$$

Practice: Convert the following temperatures. Round to one decimal place.

1. 80°F = ____________ °C = ________________ K

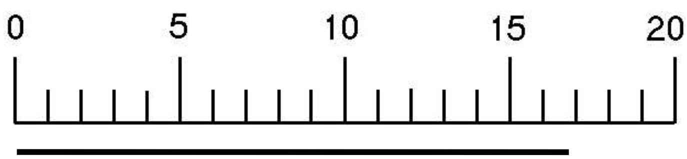
2. 310 K = ____________ °C________________ °F

Taking Measurements

Increment:

The accuracy of the apparatus (thermometer, ruler, graduated cylinder, etc.) is determined by the size of the smallest increment.

Note: Always record data to one decimal place past the increment!

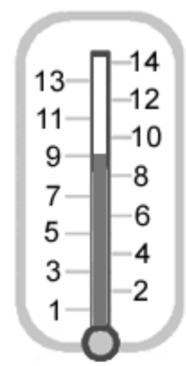


Increment:_____________

Length:_________________

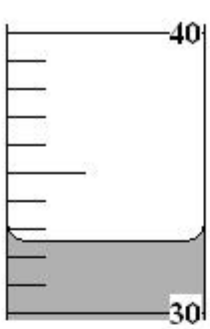
Increment:_____________

Temperature:__________



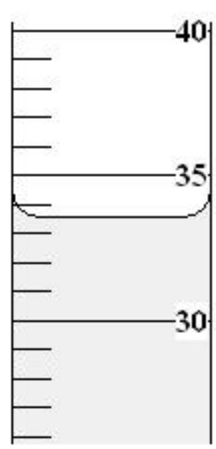
How to read a graduated cylinder

- Be eye level to the meniscus
  Meniscus: **The curved surface that a liquid forms in a narrow tube**
- Read the level to the bottom of the meniscus
- Holding a white paper behind the glassware
- Record measurement to one order of magnitude greater than the marks on the graduated cylinder



Increment: ________

Volume:___________



Increment: ________

Volume: _________

Significant Figures

In the measuremnts above the digits reported in the readings are called **significant figure**s. When we use measured quantities in calculations, the results of the calculation must reflect the precision of the measured quantities. We should not lose or gain precision during mathematical operations. The following are the rules that allow this precision not to be lost.

Significant Figures: ____

Recognizing significant figures:

1. All non-zero digits are significant (1,2,3,4,5,6,7,8,9)

   Ex. 5677 # sig figs: _____________

2. All zeros between non-zero digits are significant

   Ex.5005 # sig figs: _____________

3. All zeros before a decimal place are significant.

   Ex. 120. # sig figs: _____________

4. Trailing zeros after a decimal place, after a number are significant

   Ex. 100.0 # sig figs: _____________

5. Leading zeros (both before and after a decimal place) are ***<u>NOT</u>*** significant

   Ex. 0.0567 # sig figs: _____________

6. Powers of ten in scientific notation are not significant.

   Ex. 1.87×10^3 # sig figs ____________

To determine the number of significant figures in a number, follow test rules:

- All nonzero digits are significant.
  - 1,2,3,4,5,6,7,8,9
- Interior zeros (zeros between numbers) are significant.
  - 1.**0**5**0**8
- Trailing zeros (zeros to the right of a nonzero number) that fall after a decimal point are significant.
  - 5.1**0**
- Trailing zeros that fall before a decimal place are significant
  - 5**0**.00
- Leading zeros before/after a decimal place are **<u>not</u>** significant
  - **0.0**55
- Powers of ten in scientific notation are not significant
  - $5.5 \times \mathbf{10^3}$
- **Trailing** zeros at the end of a number, but before the implied decimal point, are ambiguous and should be avoided when using scientific notation.
  - 35**0**

Practice: How many significant figures are in each value?

A. 58.31

B. 0.00250

C. 2.7 x 10^3

D. 0.01

E. 0.500

F. 2100

Rounding Significant Figures:

When rounding to a given place value the only digit that needs to be considered is the number directly following the digit being rounded. If the number following the digit being rounded is 5 or greater, the values is rounded up by one. If the number following the digit being rounded is four or less, the value is rounded down and stays as is.

Practice: Round each of the following to the value indicated in the parentheses.

- *5*.589 (2)
- *0*.012561 (3)
- *11*.213(2)
- *6*.72 x 103 (2)
- *0*.1349999 (2)
- *25601 (2)*

Significant Figures in Calculations:

- Addition/Subtraction

  The answer is rounded to the fewest number of **decimal places**

  Practice:

  - *26.46 + 4.123 =*
  - *100.5-75.46=*
  - *20.6-2.4*

- Multiplication/Division

  The answer is rounded to the fewest number of **significant figures**

  - 5.83 x 2.3 =
  - 100.5 / 20.2 =
  - 2.4 x 0.45 /12 =

- Combined Operations

    Follow order of operations at each step, round to the appropriate number of significant figures.

    - (1.54 +2.4) x 8.57 =

    - (20.5 / 4.2) – (3.2 x 1.25) =

Practice: Perform each operation and round to the appropriate number of significant figures.

1. 15.858 + 2.3017

2. 190.2 – 60

3. 2.0 x 500.6

4. 32.8 ÷ 1.052

5. 1.50 x 2.20 – 1.15

Density: The mass of a given volume of matter.
Density can be calculated according toe the following equaitons:

$$D = \frac{m}{V} \qquad V = \frac{m}{D} \qquad m = D \times V$$

Where: D = density (g/mL or g/cm^3)
m = mass (g)
V = volume (mL or cm^3)

Density Calculations:

Ex. The density of sugar is 1.54 g/mL.

a. If a sugar cube has sides of 1.20 cm, what is the volume in cm3? In mL?

b. What is the mass of a sugar cube?

Ex. Determine the mass if a sample of ethanol has a volume of 35 mL.

Ex. Calculate the volume if an iron nail has a mass of 2.00 g. Determine the volume in L.