Notes #1 Unit 1: Foundations
Ch. Chemical Foundations

I. Significant Figures
-are digits in a number that have been measured.

Rule:
Every digit is significant, except:
-leading zeros in small numbers.
-trailing zeros in numbers without a decimal point.
In 0.0023 only the 2 & 3 count, in 5000 only the 5 counts, in 5000, all digits count.

Ex. 1) State the number of significant figures.
4.23 3
5.20 3 (has a decimal)
0.000059 2 (leading zeros)
5.09 3
20.090 5
320.00 5
10200 3 (trailing zeros, no decimal point)

II. Exponential Notation
-is a way of writing numbers using only significant figures.

#.## X 10^power (1 digit, then decimal, then rest of number)

Ex. 1) 256 = 2.56 X 10^2
     43.9 = 4.39 X10^1
     0.02 = 2 . X10^-2         Big #, positive power!
     0.000490 = 4.90 X10^-4    Small #, negative power!
     10200 = 1.02 X10^4        *non significant figures are not included

Ex. 2) Round 124973 to 3 significant figures = 1.25 X10^5
     to 2 significant figures = 1.2 X10^5
#2 Notes: III. Rules for Rounding with Significant Figures

**a) Addition/Subtraction**
The answer is rounded to the decimal place of the least accurate digit in the problem.

Ex. 1) 429.3 \( \text{tenths} \) \*tenths is least accurate (fewest decimal places)
37.45 \( \text{hundredths} \)
+ 1.93 \( \text{hundredths} \)
468.68 round it to tenths
468.7 = \( 4.687 \times 10^2 \)

Ex. 2) 0.049 \( \text{thousandths} \)
1.20 \( \text{hundredths} \) \*hundredths is least accurate
+3.02 \( \text{hundredths} \)
4.269 round to hundredths
4.27 = \( 4.27 \times 10^0 \)

Ex. 3) \( 1.956 \times 10^2 - 2.3 \times 10^1 \)
195.6 \( \text{tenths} \)
-23. \( \text{ones} \) \*ones is least accurate
172.6 round to ones
173 = \( 1.73 \times 10^2 \)

**Take numbers out of exponential notation.**

**b) Multiplication/Division**
The answer should have the same number of significant figures as the term with the least number of significant figures in the problem.

Ex. 1) \( 4.2 \times 10^4 \times 2.43 \times 10^2 = (4.2 \times 10^4) \times (2.43 \times 10^2) = 10206000 = 1.0206 \times 10^7 \)
1\(^{st}\) number has 2 sig fig, 2\(^{nd}\) number has 3 sig fig, 2 is least

Ex. 2) \( 4.9 \times 10^{-2} \times 3.11 \times 10^2 / 3.97 \times 10^3 = (4.9 \times 10^{-2})(3.11 \times 10^2) / (3.97 \times 10^3) = 0.003838539 = 3.838539 \times 10^{-3} = 3.8 \times 10^{-3} \)
1\(^{st}\) number has 2 sig fig, 2\(^{nd}\) number has 3 sig fig, 3\(^{rd}\) number has 3 sig fig, 2 is least

\( X_{10^n} \quad (-) \)
On calculator: 4.9 \( \text{EXP or EE} \quad +/- \quad 2 \quad \times \quad 3.11 \quad \text{EXP or EE} \quad 2 \quad \div \quad 3.97 \quad \text{EXP or EE} \quad 3 \quad = \)

**Some new calculators have a \( X_{10^n} \) button, instead of \( \text{EXP or EE} \). (Do not confuse this with the \( 10^x \) button, which is for inverse log! Never type in \( X \ 1 \ 0 \) separately!)**

**Some new calculators have a \( (-) \) or a \( +/- \) button, instead of a \( +/- \) button. On the calculator, these buttons are usually on top of the 7 or 8 or by the equals (=) sign.**
Ex. 3) \(3.729 \times 10^2 \times 2.36 \times 10^2 / 4.56 \times 10^2 = (3.729 \times 10^2) (2.36 \times 10^2) / (4.56 \times 10^2)\) = 192.9921053 = \(1.93 \times 10^2\)

1st number has 4 sig fig, 2nd number has 3 sig fig, 3rd number has 3 sig fig, 3 is least

Exact Numbers
- have an infinite number of significant figures.

4 eggs = 4.00000000
Averaging 3 numbers \(\frac{\# + \# + \#}{3}\)
the 3 is exact, 3.00000000

Exact numbers will not limit you, look at the other numbers in the problem for rounding.
Notes #3 IV. Combined Operations

Steps:
1) Work out the problem on the calculator.
2) Then go back and find what the numbers should be rounded to.
3) Round off the original answer.

Ex. 1) \((4.23 + 5.6) \cdot (3.13 + 4.937) = (9.83) \cdot (8.067) = 79.2986\)

\[
\begin{array}{c}
4.23 \\
\downarrow +5.6 \\
9.8 \text{ rounded to tenths}
\end{array}
\quad \begin{array}{c}
3.13 \\
\downarrow +4.937 \\
8.07 \text{ rounded to hundredths}
\end{array}
\]

\[
(9.8) \cdot (8.07) = 79 = 7.9 \times 10^1 \quad \text{rounded to 2 sig fig}
\]

2 sig fig is the least amount

Ex. 2) \((1.53 + 2.961 + 37.0) = 41.491 = 8.7441517\)

\[
\begin{array}{c}
1.53 \\
\downarrow +2.961 \\
\downarrow +37.0 \\
41.5 \text{ rounded to tenths}
\end{array}
\quad \begin{array}{c}
42.3 \\
\downarrow -29.345 \\
\downarrow -8.21 \\
4.7 \text{ rounded to tenths}
\end{array}
\]

\[
(41.5) / (4.7) = 8.7 = 8.7 \times 10^0 \quad \text{rounded to 2 sig fig}
\]

2 sig fig is the least amount

Ex. 3) \((79.12 - 16.007 + 0.1) = 63.213 = 0.8511242\)

\[
\begin{array}{c}
79.12 \\
\downarrow -16.007 \\
\downarrow +0.1 \\
63.2 \text{ rounded to tenths}
\end{array}
\quad \begin{array}{c}
49.3 \quad \leftarrow \text{hundredths least accurate} \\
\downarrow +24.970 \\
74.2 \text{ rounded to hundredths}
\end{array}
\]

\[
(63.2) / (74.27) = 0.851 = 8.51 \times 10^{-1} \quad \text{rounded to 3 sig fig}
\]

3 sig fig is the least amount
V. SI Units

Mass = grams (really kilograms)
Length = meters
Time = seconds
Volume = liters

<table>
<thead>
<tr>
<th>Multipliers</th>
<th>Example with meters: (**Same for grams, seconds &amp; liters!)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mega (M)</td>
<td>1 Mm = 1 X10^6 m</td>
</tr>
<tr>
<td>**kilo (k)</td>
<td>1 km = 1 X10^3 m</td>
</tr>
<tr>
<td>**deci (d)</td>
<td>1dm = 1 X10^-1 m</td>
</tr>
<tr>
<td>**centi (c)</td>
<td>1cm = 1 X10^-2 m</td>
</tr>
<tr>
<td>**milli (m)</td>
<td>1 mm = 1 X10^-3 m</td>
</tr>
<tr>
<td>micro (µ)</td>
<td>1 µm = 1 X10^-6 m</td>
</tr>
<tr>
<td>nano (n)</td>
<td>1 nm = 1 X10^-9 m</td>
</tr>
<tr>
<td>pico (p)</td>
<td>1 pm = 1 X10^-12 m</td>
</tr>
</tbody>
</table>

length

1 m = 39.37 in
2.54 cm = 1 in
1 km = 0.621 mile
1 mile = 5280 ft

Mass

1kg = 2.205 lb
1 lb = 16 oz

Volume

1 L = 1.06 qt        ** 1 L = 1dm^3   and   1 ml = 1cm^3
1 gal = 3.773 L
1 gal = 4 qt
#4 Notes  VI. Conversions

Ex. 1) Convert 2.00 hr to min

\[
\frac{2.00 \text{ hr}}{1 \text{ hr}} \times \frac{60 \text{ min}}{1 \text{ hr}} = 120 \text{ min} = 1.20 \times 10^2 \text{ min}
\]

1 hr = 60 min

Ex. 2) Convert 5.0 m to km

\[
\frac{5.0 \text{ m}}{1 \text{ km}} \times \frac{1 \text{ km}}{1 \times 10^3 \text{ m}} = 0.005 = 5.0 \times 10^{-3} \text{ km}
\]

1 km = 1 \times 10^3 m

Ex. 3) Convert 2.5 g to mg

\[
\frac{2.5 \text{ g}}{1 \text{ mg}} \times \frac{1 \text{ mg}}{1 \times 10^{-3} \text{ g}} = 2.5 \times 10^3 \text{ mg}
\]

1 mg = 1 \times 10^{-3} g

Ex. 4) Convert 49.3 cl to l

\[
\frac{49.3 \text{ cl}}{1 \text{ cl}} \times \frac{1 \times 10^{-2} \text{ l}}{1 \text{ cl}} = 4.93 \times 10^{-1} \text{ l}
\]

1 cl = 1 \times 10^{-2} l
Ex. 5) Convert 49.6 in to miles

12 in = 1 ft  
1 mile = 5280 ft

\[
\begin{array}{c|c|c}
\text{in} & \text{ft} & \text{mi} \\
12 & 1 & 5280 \\
\end{array}
\]

1 m = 39.37 in  
1 km = 1 X10^3 m  
1 km = 0.621 mile

\[
\begin{array}{c|c|c|c|c|c|c}
\text{in} & \text{m} & \text{km} & \text{mi} \\
\text{m} & 1 & 10^3 & 0.621 \\
\end{array}
\]

2.54 cm = 1 in  
1 cm = 1 X10^-2 m  
1 km = 1 X10^3 m  
1 km = 0.621 mile

\[
\begin{array}{c|c|c|c|c|c|c}
\text{in} & \text{cm} & \text{m} & \text{km} & \text{mi} \\
\text{cm} & 1 & 10^{-2} & 10^3 & 0.621 \\
\end{array}
\]

\[
\left( \frac{49.6 \text{ in}}{12 \text{ in}} \right) \left( \frac{1 \text{ ft}}{1 \text{ in}} \right) \left( \frac{1 \text{ mi}}{5280 \text{ ft}} \right) = 7.83 \times 10^{-4} \text{ mi}
\]

Ex. 6) Convert 6.2 hr to sec

hr → min → sec

\[
\left( \frac{6.2 \text{ hr}}{1 \text{ hr}} \right) \left( \frac{60 \text{ min}}{1 \text{ hr}} \right) \left( \frac{60 \text{ sec}}{1 \text{ min}} \right) = 2.2 \times 10^4 \text{ sec}
\]

hr → min → sec

1 hr = 60 min, 1 min = 60 sec
Ex. 7) Convert 49.6 in to km

\[
\begin{array}{c|c|c}
49.6 \text{ in} & 1 \text{m} & 1 \text{km} \\
39.37 \text{ in} & 1 \times 10^3 \text{ m} & = 1.26 \times 10^{-3} \text{ km}
\end{array}
\]

\[1 \text{ m} = 39.37 \text{ in}, \quad 1 \text{km} = 1 \times 10^3 \text{ m}\]
Ex. 1) Convert $4.3 \text{ dm}^3$ to $\text{cm}^3$

\[
\begin{array}{c|c|c|c}
4.3 \text{ dm}^3 & 1 \text{ L} & 1 \text{ ml} & 1 \text{ cm}^3 \\
\hline
1 \text{ dm}^3 & 1 \times 10^{-3} \text{ L} & 1 \times 10^{-3} \text{ ml} & 1 \text{ cm}^3
\end{array}
\]

$4.3 \times 10^3 \text{ cm}^3$

$1 \text{ dm}^3 = 1 \text{ L}, \quad 1 \text{ ml} = 1 \times 10^{-3} \text{ L}, \quad 1 \text{ cm}^3 = 1 \text{ ml}$

OR

\[
\begin{array}{c|c|c|c}
4.3 \text{ dm}^3 & 1 \times 10^{-3} \text{ m}^3 & 1 \text{ cm}^3 \\
\hline
1 \text{ dm}^3 & 1 \times 10^{-6} \text{ m}^3 & 1 \text{ cm}^3
\end{array}
\]

$(1 \text{ dm})^3 = (1 \times 10^{-1} \text{ m})^3, \quad (1 \text{ cm})^3 = (1 \times 10^{-2} \text{ m})^3$

$1 \text{ dm}^3 = 1 \times 10^{-3} \text{ m}^3, \quad 1 \text{ cm}^3 = 1 \times 10^{-6} \text{ m}^3$

Ex. 2) Convert $9.0 \text{ m/sec}$ to $\text{km/hr}$

\[
\begin{array}{c|c|c|c}
9.0 \text{ m} & 1 \text{ km} & 60 \text{ sec} & 60 \text{ min} \\
\hline
\text{sec} & 1 \times 10^3 \text{ m} & 1 \text{ min} & 1 \text{ hr}
\end{array}
\]

$3.2 \times 10^1 \text{ km/hr}$

$1 \text{ km} = 1 \times 10^3 \text{ m}, \quad 1 \text{ min} = 60 \text{ sec}, \quad 1 \text{ hr} = 60 \text{ min}$

Ex. 3) Convert $15.6 \text{ kg/m}^3$ to $\text{g/cm}^3$

\[
\begin{array}{c|c|c|c}
15.6 \text{ kg} & 1 \times 10^3 \text{ g} & 1 \times 10^{-6} \text{ m}^3 \\
\hline
\text{m}^3 & 1 \text{ kg} & 1 \text{ cm}^3
\end{array}
\]

$1.56 \times 10^{-2} \text{ g/cm}^3$

$1 \text{ kg} = 1 \times 10^3 \text{ g}, \quad (1 \text{ cm})^3 = (1 \times 10^{-2} \text{ m})^3$

$1 \text{ cm}^3 = 1 \times 10^{-6} \text{ m}^3$

Ex. 4) Convert $1.34 \times 10^6 \text{ cm/sec}$ to $\text{mi/day}$

\[
\begin{array}{c|c|c|c|c|c|c|c|c}
1.34 \times 10^6 \text{ cm} & 1 \times 10^{-2} \text{ m} & 1 \text{ km} & 0.621 \text{ mi} & 60 \text{ sec} & 60 \text{ min} & 24 \text{ hr} \\
\hline
\text{sec} & 1 \text{ cm} & 1 \times 10^3 \text{ m} & 1 \text{ km} & 1 \text{ min} & 1 \text{ hr} & 1 \text{ day}
\end{array}
\]

$7.19 \times 10^5 \text{ mi/day}$
VIII. Temperature & Density

A) Temperature Conversions

°C = 5/9 (°F – 32)  K = °C + 273

°C = Celsius  °F = Fahrenheit  K = Kelvin ( 0 K = Absolute Zero)

-lowest possible temperature, all motion stops, no Kinetic Energy

Ex. 1) What is 154 K in °F?

K = °C + 273
154 K = °C + 273
-119 = °C

°C = 5/9 (°F – 32)
-119 °C = 5/9 (°F – 32)
9/5 (-119 °C) = °F – 32
-214.2 = °F – 32
-182 = °F

Ex. 2) What is 72 °F in Kelvin?

°C = 5/9 (°F – 32)
°C = 5/9 (72 °F – 32)
°C = 22

K = °C + 273
K = 22 °C + 273
K = 295
IX. **Density**

Density \(= \frac{\text{mass}}{\text{volume}}\) (grams/cm\(^3\))

- Mass is constant, weight changes (weight = mass \(\times\) gravity)
- Density of water = 1 g/cm\(^3\)
  - \(\text{O}_2\) = 1.33 \(\times\) 10\(^{-3}\) g/cm\(^3\)
  - \(\text{Au}\) = 19.32 g/cm\(^3\) L. aurum
  - \(\text{Al}\) = 2.70 g/cm\(^3\)
  - \(\text{Ag}\) = 10.5 g/cm\(^3\) L. argentums

**Ex. 1**) Bismuth has a density of 9.80 g/cm\(^3\). What is the mass of 4.32 ml of Bi?

\[
\begin{align*}
4.32 \text{ ml} & \quad \frac{1 \text{ cm}^3}{1 \text{ ml}} = 4.32 \text{ cm}^3 \\
D &= \frac{m}{v} \\
9.80 \frac{\text{g}}{\text{cm}^3} &= \frac{m}{4.32 \text{ cm}^3} \\
4.23 \times 10^1 \text{ g} &= m
\end{align*}
\]

**4.23 \times 10^1 \text{ g} = m**

**Ex. 2**) Iron has a density of 7.87 g/cm\(^3\). What volume would 2.46 \(\times\) 10\(^{-2}\) kg of Fe occupy?

\[
\begin{align*}
2.46 \times 10^{-2} \text{ kg} & \quad \frac{1 \times 10^3 \text{ g}}{1 \text{ kg}} = 24.6 \text{ g} \\
D &= \frac{m}{v} \\
7.87 \frac{\text{g}}{\text{cm}^3} &= \frac{24.6 \text{ g}}{V} \\
7.87 (V) &= 24.6 \quad **\text{When in doubt, cross multiply!} \\
V &= \frac{24.6}{7.87} \\
V &= 3.13 \times 10^0 \text{ cm}^3
\end{align*}
\]
#7 Notes

**X. States of Matter**

Matter is anything that occupies space and has mass.

1) **Solid**: particles arranged in a rigid pattern, definite shape & volume.
2) **Liquid**: particles close together in no pattern, takes the shape of its container, definite volume.
3) **Gas**: no definite shape or volume (shape and volume of container)
4) **Plasma**: hot gas-like mixture over 5000 °C, collisions cause some electrons to be knocked off, creating (+) ions. This charged mixture that conducts electricity is plasma.

**XI. Mixtures**

[Heterogeneous Mixtures] ↔ [Homogeneous Mixture = Solutions]

Parts are visible

methods parts are indistinguishable

\[\downarrow \text{physical methods}\]

Pure Substances (Homo)

\[\downarrow \text{chemical methods}\]

Elements (Homo) ↔ Compounds (Homo)

(One type of atom) ↔ (2 or more types of atoms)

Ex. 1) Heterogeneous or Homogeneous?

a) salt poured into water Hetero
b) salt stirred/mixed into water Homo
c) soda Hetero
d) flat soda Homo
e) unmixed Kool Aid in water Hetero
f) apple Hetero

Ex. 2) Element or Compound? (=Pure Substance = Homo)

a) salt compound
b) isopropyl alcohol compound *look up formula in textbook/internet
c) gold element
d) oxygen element
e) sugar compound

**XII. Physical/Chemical Characteristics/Changes**

Physical: melting, freezing, boiling, solubility, changing shape, malleability, conductivity.

Chemical: burning, exploding, reacting with acid, toxicity.
*End of Notes*  (Assignments #8-9 are Review Assignments. There are no notes for these assignments.)