Ch. 1 Notes (part 2) – THE CENTRAL SCIENCE
MEASUREMENT

NOTE: Vocabulary terms are in **<u>boldface and underlined</u>**. Supporting details are in *italics.*

I. Units and Measurement - Metrics

A. Système Inernationale d'Unités – *The International System of Units (SI)*

1) *modern metric system of measurement* established in 1960
2) U.S. Metric Conversion Act of 1975—commitment to encourage metrics
3) 1978—formal metric use decreed in Europe

B. seven **<u>base units</u>** * = important

SI BASE UNIT	NAME	SYMBOL
length *	meter	m
mass *	kilogram	kg
time *	second	s
temperature *	Kelvin	K
amount of substance *	mole	mol
electric current	ampere	A
luminous (light) intensity	candela	cd

Main metric prefixes (* = most common)

Factor	Name	Symbol
10^{12}	tera	T
10^{9}	giga	G
10^{6}	mega	M *
10^{3}	kilo	k *
10^{2}	hecto	h *
10^{1}	deka	da *

Factor	Name	Symbol
10^{-1}	deci	d *
10^{-2}	centi	c *
10^{-3}	milli	m *
10^{-6}	micro	μ *
10^{-9}	nano	n
10^{-12}	pico	p

You must show all work in calculations, unless specified in the directions.
THIS IS A SHORTCUT YOU CAN USE TO CHECK YOUR WORK:

"King	Henry	Doesn't	Usually	Drink	Chocolate	Milk"
kilo	hecto	deka	(base or plain unit)*	deci	centi	milli
k	h	da	(base or plain unit)*	d	c	m
1000x larger	100x larger	10x larger	1	10x smaller	100x smaller	1000x smaller

(although kilogram is the SI base unit for mass, we use gram here for the plain mass unit)

1) *time*—**<u>second</u>** is the SI base unit because many reactions happen so quickly
2) *length—linear distance* between two points
   a) SI base unit: **<u>meter</u>**, the base unit of linear measure
   b) common metric units: km, m, cm, mm, μm, nm

3) **mass**—*the amount of matter in an object*

a) **kilogram**—SI base unit (1 kg = 2.2 lbs.)

b) common metric units: kilogram (kg), gram (g)

4) **temperature**—*the measure of the average K.E. (kinetic energy)*

a) *"cold" is actually an absence of heat*

b) heat transfer—heat energy flow from a higher-temp object to a lower-temp object; occurs when there is contact between the areas/objects

c) temperature scales

i. Celsius (°C)— 0° f.p. H_2O, 100° b.p. H_2O

ii. Fahrenheit (°F)— 32° f.p. H_2O, 212° b.p. H_2O

iii. Kelvin (K)— 273 f.p. H_2O, 373 b.p. H_2O

1. no ° sign used with Kelvin
2. **absolute zero**—*0 Kelvin; all molecular motion stops*

iv. conversions

$\mathbf{K = C + 273.15}$ $\quad$ $°C = 5/9 \times (°F - 32)$ $\quad$ $°F = (9/5 \times °C) + 32$

d) examples

EXAMPLE 1: Convert 498.00 K to Celsius.

$K = C + 273.15 \quad C = K - 273.15 \quad C = 498.00 - 273.15 = \boxed{224.85\ °C}$

EXAMPLE 2: Convert 37.00° C to Kelvin.

$K = C + 273.15 \quad K = 37.00 + 273.15 = \boxed{310.15\ K}$

5) *amount of substance* = **mole** (preview of second semester) – see sec. IV

a) a *mole (mol) is a unit used in science to count very small particles like atoms*

b) 6.02×10^{23} = 602,000,000,000,000,000,000,000

C. **derived units**—complex units formed from the combination of multiple units

Common SI DERIVED UNITS	EXAMPLE NAME	EXAMPLE UNITS
*area	square meter	m^2
*volume	cubic meter	m^3
	$1L = 1\ dm^3$	& $1\ mL = 1\ cm^3 = 1\ cc$
*density	kilogram per cubic meter	kg/m^3 (mass/volume)
speed	meter per second	m/s
velocity	meter per second, with direction	m/s NE
acceleration	meter per second squared	m/s^2

1) **volume**—*the 3-D space an object occupies*; a derived unit

a) common unit: *liter*—standard volume unit for liquid volume

b) common units: cubic meter (m^3), cubic cm (cm^3 or cc), milliliter (mL)

1 L = 1 dm³ → $1\ L = 1\ dm^3$ $\qquad$ $1\ mL = 1\ cm^3 = 1\ cc$

c) liquid volumes can be measured in beakers, graduated cylinders, pipets, syringes, burets, volumetric flasks

d) *solid volume = Length x Width x Height (L x W x H)*

2) **density**

a) **density = mass / volume**

b) **D = M / V** $\qquad$ M = D x V $\qquad$ V = M / D

(be comfortable with algebraic manipulation of D = M/V)

c) density usually decreases as temp. increases, due to increased volume

d) examples…

EXAMPLE 3:
A metal bar has a mass of 35.50 g and a volume of 262 cm^3. What is its density?

$$\mathbf{D} = \frac{\mathbf{M}}{\mathbf{V}} = \frac{35.50\ g}{262\ cm^3} = \boxed{0.135\ \frac{g}{cm^3} \text{ or } g/cm^3}$$

EXAMPLE 4:
500.0 mL of a liquid has a density of 0.447 g/mL. What is its mass?

$$D = \frac{M}{V} \qquad \mathbf{M = DV} \qquad M = 0.447\ \frac{g}{\cancel{mL}} \times 500.0\ \cancel{mL} = \boxed{224\ g}$$

EXAMPLE 5:
4.2 g of a substance has a density of 0.89 g/m^3. How much space does it occupy?

$$D = \frac{M}{V} \qquad \mathbf{V} = \frac{\mathbf{M}}{\mathbf{D}} \qquad V = \frac{4.2\ \cancel{g}}{0.89\ \cancel{g}/m3} = \boxed{4.7\ m^3}$$

EXAMPLE 6:
A 3.888 g sample of Bauckium is placed in 35.7 mL of water. The water level rises to 50.2 mL. What is the density of Bauckium?

water displacement volume: 50.2 mL – 35.7 mL = 14.5 mL

$$\mathbf{D} = \frac{\mathbf{M}}{\mathbf{V}} = \frac{3.888\ g}{14.5\ mL} = \boxed{0.268\ g/mL}$$

II. **SCIENTIFIC NOTATION**: *$M \times 10^n$ format*

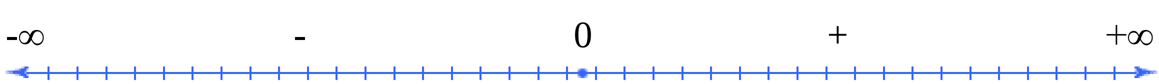
A. exponent examples

$10^0 = 1$	$10^1 = 10$	$10^2 = 100$	$10^3 = 1000$
$10^{-1} = 0.1$ (1/10)	$10^{-2} = 0.01$ (1/100)	$10^{-3} = 0.001$ (1/1000)	

A. scientific notation rules for review

(You should be comfortable with scientific notation already from previous math classes.)

1) *POSITIVE EXPONENTS INDICATE MAKING A NUMBER LARGER. MOVE DECIMAL POINT TO THE RIGHT TO EXPAND.*
2) *NEGATIVE EXPONENTS INDICATE MAKING A NUMBER SMALLER. MOVE DECIMAL POINT TO THE LEFT TO EXPAND.*
3) *ADDITION AND SUBTRACTION OF EXPONENTS CAN ONLY BE DONE IF THE EXPONENTS ARE THE SAME.*
4) *FOR MULTIPLICATION OF EXPONENTIAL NUMBERS, ADD THEIR EXPONENTS.*
5) *FOR DIVISION OF EXPONENTIAL NUMBERS, SUBTRACT THEIR EXPONENTS.*



B. Examples of scientific notation rules (see previous page)

1) Rule 1: $3.11 \times 10^{2} = 311$

2) Rule 2: $3.11 \times 10^{-2} = 0.0311$

3) Rule 3: $(6.426 \times 10^{15}) + (3.09 \times 10^{13}) =$

$(6.426 \times 10^{15}) + (0.0309 \times 10^{15}) =$

~ remember, making the exponent (n) larger makes the initial number (M) smaller

(6.4569×10^{15})

4) Rule 4/5: $(8 \times 10^{6}) / (2 \times 10^{-4}) =$

$8/2 = 4$ $\quad$ $6-(-4) = 10$

4×10^{10}

C. Calculator differences

1) You are responsible for knowing how to use a scientific calculator. Learn how to use your own, as well as the types we have in our room.
2) Exponent keys vary: $\boxed{\text{EXP}}$ $\boxed{\text{EE}}$ $\boxed{2^{\text{nd}}}\boxed{\text{EE}}$ $\boxed{\text{x } 10^{x}}$
3) With my classroom calculators, DO NOT type in x or 10.
4) If you have the $\boxed{\text{x } 10^{x}}$ button, you must put parentheses around the entire number.
5) Example: 6.02×10^{23} could be entered as

6.02 $\boxed{\text{EXP}}$ 23 6.02 $\boxed{\text{EE}}$ 23 6.02 $\boxed{2^{\text{nd}}}\boxed{\text{EE}}$ 23 **(**6.02 $\boxed{\text{x } 10^{x}}$ 23**)**

III. Significant Figures, "Sig.Figs" (also known as Significant Digits, "Sig.Digs")

A. **significant figures**— *all numbers in a measurement that are measured accurately, plus a last estimated digit*

B. Why do sig.figs???

The answer is only as precise as the equipment used to take the measurements.

C. important sig.fig. rules

1) **ALL NONZERO DIGITS ARE ALWAYS SIGNIFICANT.**

4.2 and 27 both have two sig.figs.

2) **ZEROES BETWEEN TWO NONZERO DIGITS ARE ALWAYS SIGNIFICANT. ZEROES BETWEEN TWO SIGNIFICANT DIGITS ARE ALWAYS SIGNIFICANT.**

("Sig. Fig. Sandwich") 8.909 and 1005 both have four sig.figs.

3) **ZEROES TO THE LEFT OF NONZERO DIGITS ARE NOT SIGNIFICANT.**

0.0006 and 0.06 both have only one sig.fig.

4) **TERMINAL ZEROES AFTER THE DECIMAL POINT ARE ALWAYS SIGNIFICANT.**

1.000 and 9.820 both have four sig.figs.

5) **TERMINAL ZEROES NOT INVOLVING A DECIMAL POINT ARE NOT SIGNIFICANT... UNLESS WRITTEN IN SCIENTIFIC NOTATION FOR CLARIFICATION or UNLESS A DECIMAL POINT IS PLACED AFTER THE LAST ZERO.**

1230 written as 1.23×10^{3} has three sig.figs.

1230 written as 1230. or 1.230×10^{3} has four sig.figs.

C. rounding rules

1) round up if the number after the last sig.fig. is 5 or greater
(48.47 rounded to three sig.figs. is 48.5)
2) round down if the number after the last sig.fig. is less than 5
(140.081 rounded to five sig.figs. is 140.08)

D. sig.fig. addition, subtraction, multiplication, and division rules

1) **IN ADDITION AND SUBTRACTION, THE ANSWER MAY CONTAIN ONLY AS MANY DECIMAL PLACES AS THE LEAST ACCURATE VALUE.**

5.2208 + 0.1 = 5.3208	5.3 adjusted
121.50 + 9000 = 9121.50	9122 adjusted

2) **IN MULTIPLICATION AND DIVISION, THE ANSWER MAY CONTAIN ONLY AS MANY TOTAL DIGITS AS THE LEAST ACCURATE VALUE USED.**

5 x 10.000 = 50.000	50 adjusted
49.600 / 47.40 = 1.0464135	1.046 adjusted

IV. Dimensional analysis

A. <u>dimensional analysis (DA)</u>

1) also known as the *factor unit* and *factor label* methods
2) *using the units (dimensions) to solve problems*
3) using dimensional analysis:

"Play checkers" with the units, moving them diagonally, canceling when appropriate. All units should cancel except those of the desired answer.

B. Steps for success

1) identify unknown (read carefully)
2) identify known (read carefully)
3) plan solution
4) calculate
5) check (sig.figs., units, and math)

C. **<u>conversion factor</u>**—*a ratio of two equivalent measurements*

(SMALL #)	*(LARGE UNIT)*	=	*(LARGE #)*	*(SMALL UNIT)*
1	foot	=	12	inches
1	day	=	24	hours
1	mole	=	6.02×10^{23}	particles

Conversion factors that are exact are an infinite number of sig.figs. (do not limit the sig.figs.), such as metrics or time conversions. The 1's in conversions are exact.

D. mole conversions (preview of second semester)

1) particle –mole conversions:

1 mole = 6.02×10^{23} particles and *6.02×10^{23} particles = 1 mole*

2) liter –mole conversions:

1 mole = 22.4 L and *22.4 L = 1 mole*

- Under STP conditions (standard temperature and pressure), 1 mole of any gas will occupy 22.4 L of space.
- This is called the molar volume of a gas.

3) gram-mole conversions:

1 mole = (atomic mass) g	and	*(atomic mass) g = 1 mole*

- The decimal number of an element on the periodic table (atomic mass) is equal to the number of grams in one mole of that element.
- This is called the molar mass of an element.
- We round the atomic masses from the periodic table to 0.01 g (hundredths) in this class.

E. DA Examples

EXAMPLE 7: How many seconds are in exactly one century?

$$1\ \cancel{\text{century}} \times \frac{100\ \cancel{\text{yrs.}}}{1\ \cancel{\text{century}}} \times \frac{365.25\ \cancel{\text{days}}}{1\ \cancel{\text{yr.}}} \times \frac{24\ \cancel{\text{hrs.}}}{1\ \cancel{\text{day}}} \times \frac{60\ \cancel{\text{min.}}}{1\ \cancel{\text{hour}}} \times \frac{60\ \text{sec.}}{1\ \cancel{\text{min.}}} = \boxed{3{,}155{,}760{,}000\ \text{sec.}}\ (\infty\ \text{s.f.})$$

EXAMPLE 8: A car can travel 28.0 miles on one gallon of gas. Convert this value to m/L. (1 gal = 4 quarts 1.61 km = 1 mi. 1 L = 1.06 quarts)

(remember, m = meter and mi = mile)

$$\frac{28.0\ \cancel{\text{mi}}}{\cancel{\text{gal}}} \times \frac{1.61\ \cancel{\text{km}}}{1\ \cancel{\text{mi}}} \times \frac{1000\ \text{m}}{1\ \cancel{\text{km}}} \times \frac{1\ \cancel{\text{gal}}}{4\ \cancel{\text{qu}}} \times \frac{1.06\ \text{qu}}{1\ \text{L}} = \boxed{11900\ \text{m/L}}$$

EXAMPLE 9: How many atoms are found in 0.2332 mol of helium?

$$0.2332\ \text{mol He} \times \frac{6.02 \times 10^{23}\ \text{atoms He}}{1\ \text{mol He}} = \boxed{1.40 \times 10^{23}\ \text{atoms He}}$$

EXAMPLE 10: Find the number of moles in 67 g of helium.

$$67\ \cancel{\text{g He}} \times \frac{1\ \text{mol He}}{4.00\ \cancel{\text{g He}}} = \boxed{17\ \text{mol He}}$$

EXAMPLE 11: Calculate the volume of helium gas in L if there are 0.8154 mol in the container. Assume STP conditions.

$$0.8154\ \text{mol He} \times \frac{22.4\ \text{L He}}{1\ \text{mol He}} = \boxed{18.3\ \text{L He}}$$

V. Uncertainty in Data

A. sig.figs (see earlier notes)

B. important terms

1) **accuracy**—*how close a measurement comes to the true (theoretical) value*

2) **precision**—*how a measurement can be reproduced (consistent; repeatable)*

C. *graph – visual data display*

1) types of graphs: circle (pie chart), bar, line

2) graph <u>**slope**</u> = $\Delta Y / \Delta X = (y_2 - y_1) / (x_2 - x_1)$ = RISE / RUN

3) <u>**independent variable**</u> (x axis)

4) <u>**dependent variable**</u> (y axis)

C. interpolation and extrapolation

1) *interpolation*—"the insertion of an intermediate value into a series by estimating or calculating it from surrounding known values" (from Oxford)

2) *extrapolation*—estimating a projected path of a line beyond its calculated values

E. error in measurement

1) **error** = | accepted value – experimental value| = |(A – E)|

2) **percent error** = $\frac{|A - E|}{A} \times 100$

EXAMPLE 12:

Dewayne estimated that he had 150 mL of acetic acid in his beaker. When he poured it into a graduated cylinder which gives more accurate measurements, the volume of liquid was 157.6 mL. What was Dewayne's percent error?

$$\frac{|A - E|}{A} \times 100 = \frac{|157.6 - 150|}{157.6} \times 100 = \frac{7.6}{157.6} \times 100 = \boxed{4.8\% \text{ error}}$$