Block 3 9/9 Fall 16

Chemistry Unit 1

Primary reference: Chemistry, Addison-Wesley

Topic	Essential Knowledge	Study Support
Scientific Investigation **1.1** **SOL 1a, 1b,1c, 1e, 1g**	Use chemicals and equipment safely. **Scientific notation** is used to express very small or very large measurements in powers of ten. Example: $3.2 \times 10^4 = 32{,}000$ **Accuracy** is how close a measurement is to the true value. An accurate measurement has very little error. **Precision** is measure exactness and repeatability. When making measurements, the measurement can only include 1 estimated value. All digits that are known precisely and the 1 estimated value are called **significant figures**. $\boxed{\textbf{Percent Error} = 100 \times \lvert\text{accepted value-exper. value}\rvert/\text{accepted value}}$ **Significant figures** are all the digits that can be known precisely in a measurement plus a last estimated digit. **Significant figure calculation rules** are used to round calculations with lab data. In **addition and subtraction** round the answer to the least number of decimal places as contained by the numbers used in the calculation. In **multiplication and division** round the answer to the least number of significant figures as contained by the numbers used in the calculation Common **metric unit** prefixes are **kilo** (1000), **centi**(1/100), **milli** (1/1000). **The Unit Cancelation Method (Dimensional Analysis)** is used to in calculations involving unit conversions.	Study your Safety Contract carefully and read pp. 18-19. **Ch 3:** Read pp. 52-53 on scientific notation. Read pp. 54-55 on Percent Error Read pp. 56-62 on sig. figs. Read pp. 63-67 on the metric system. **Ch 4:** Read pp. 89-99 on Unit Canceling Method.
Atomic Structure and Periodic Relationships **2.1** **SOL 2h, 2i**	All matter is made from different chemical **elements**. The **Periodic Table of the Elements** shows the known elements, arranged by increasing atomic number. The symbol for many of the elements is one capital letter. In two-letter symbols for elements, the first letter is always an upper case letter, the second one a lower case. The smallest particle of an **element** is an **atom**. Some common elements are composed of **molecules** containing two atoms of the same element, also known as the diatomic elements. Example: hydrogen H_2(g) and oxygen O_2(g). BrINClHOF or go to 7, make a 7, don't forget H. A **chemical reaction** (chemical change) is required to change one substance into another by rearranging its atoms. In a chemical change, a new substance is formed. A **physical change** occurs when the chemical makeup of a substance stays the same but some physical properties of the substance may change. $\boxed{\text{Density} = \text{mass/volume}}$ always show units **Mixtures** are a physical blend of 2 or more substances. A **substance** can be a **compound** or an **element**. In a **heterogeneous mixture**, the different parts can be easily seen (like salt and pepper mixed together). In a **homogeneous mixture** the particles are mixed so well that the separate parts cannot be seen (like salt dissolved in water.)	**Ch 2:** Read pp. 36-39 on elements & compounds. Read pp. 41-43 on chemical rxns. Read pp. 32-38 on mixtures. Look carefully at Figures 2.3 – 2.8.
Nomenclature, Formulas, and Reactions **3.1** **SOL 3c**	Atoms of different elements can join together by chemical bonds to form a **compound.** A compound has different properties from its elements. **Chemical formulas** show the ratio or number of atoms of each element in a compound. Example: 2 hydrogen atoms bonded to one oxygen atom make a water molecule (H_2O).	**Ch 2:** Read pp. 36-40 on elements and compounds.
Molar Relationships **4.1** **SOL 4a**	Atoms and molecules are too small to count. **Mole** is the unit used to count atoms and molecules, similar to using dozens to count eggs. $\boxed{1 \text{ mole} = 6.02\ 10^{23} \text{ (atoms or molecules)}}$	**Chapter 7:** Read pp. 171-176 on the mole.
Phases of Matter and Kinetic Molecular Theory **5.1** **SOL 5a, 5d**	Atoms and molecules are in constant motion. For a given substance, **solid** particles move slowest, **liquid** particles mover faster, and **gas** particles move the fastest. **Plasma** is the 4th phase of matter. Plasmas form when gases is heated to a point where electrons dissociate from the nuclei. There is a **direct relationship** between temperature in Kelvins and speed of the particles. When the temperature increases, particles move faster. $\boxed{K = {}^\circ C + 273}$	**Ch 2:** Look carefully at Table 2.2 and Figure 2.1 on p. 30. Read pp. 30-31 and p. 267.

Objectives for Unit One (Chapters 2, 3 & 4)
Chemistry, Addison-Wesley, 2002

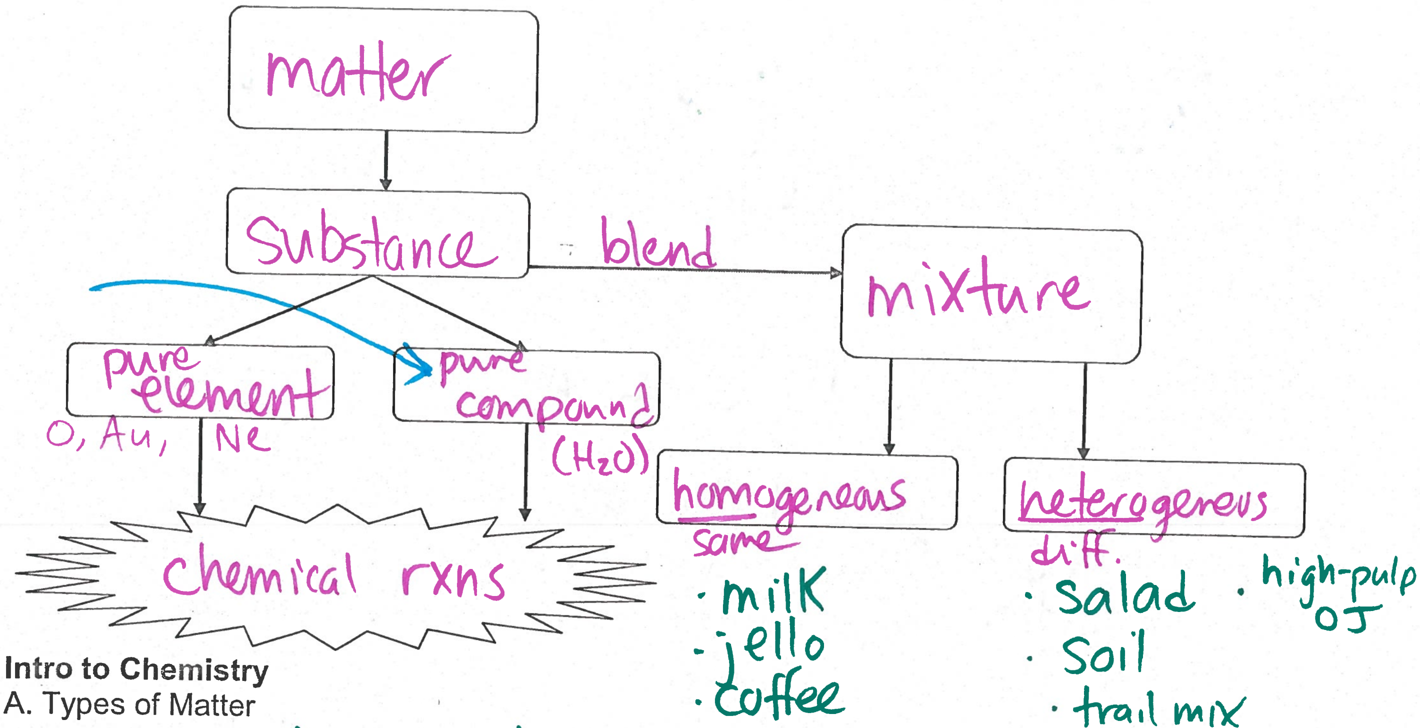
Topic Outline

I) Laboratory Safety
II) Introduction to Chemistry
   A) Types of matter (definitions)
   B) Phases of matter and kinetic theory
      1) Kinetic Theory
      2) Phases of Matter
      3) Converting between °C and K.
   C) Physical vs. chemical properties and changes
   D) Basics of chemical reactions
III) Scientific Measurements and Math
   A) Measurement uncertainty
      1) Accuracy and precision
      2) % Error Calculations
   B) Scientific Calculation Basics
      1) Scientific notation
      2) Significant figures
      3) Conversion factors and the unit cancellation method(a.k.a. dimensional analysis)
      4) Metric System units and the mole
      5) Calculating density

Objectives *(text problems follow in italics)*

1. Identify the chemical symbol for elements 1-38 plus Ag, Cd, Sn, I, Xe, Cs, Ba, Pt, Au, Hg, Pb, Rn, Fr from the elements name and *visa versa* (3a) Flashcards required for these 51 elements!
2. Know the basic laboratory safety rules
3. Differentiate between elements, substances, compounds, and heterogeneous/homogeneous mixtures p.*47#30,32*
4. Memorize the seven diatomic elements (BrINClHOF)
5. Differentiate between chemical and physical properties and changes *p47#34,35,40*
6. Understand the basic differences between a gas, liquid, and solid in terms of kinetic theory p47#29
7. Understand the direct relationship between temperature and speed of particles.
8. Understand the inverse relationship between pressure and volume of a gas.
9. Use scientific notation properly including multiplying and dividing using scientific notation *p.53#3,4*
10. Determine the number of significant figures in any number p78#41
11. Use significant figures correctly in multiplication, and division problems *p78#44,#49*
12. Memorize and use (SI) metric base units correctly (mass, length, volume, temperature, mole)
13. Memorize and use the conversion equation between °C and K temperature scale. *p.75#30,3*
14. Memorize and convert between metric unit prefixes (kilo, centi, milli) *p94#11,12;p95#16,17*
15. Memorize that 1 mole = 6.02×10^{23} particles
16. Explain the difference between precision and accuracy *p78#39*
17. Calculate percent error from word problems *p.78#48*
18. Memorize and use the density equation (D=m/v) to calculate density, mass, or volume from word problems.*p71#24, p72#26,28.*
19. Use the unit cancellation method to convert between units and measurements in word problems *p100#29-31*

Recommended Text: Chapter 2 pp. 29-43, Chapter 3: pp. 51-71, Chapter 4: pp 83-101 (but depend on lecture)



Intro to Chemistry

A. Types of Matter

Matter: anything w/ mass

Mass: amt of material in something

Substance: any matter w/ uniform composition.

Examples: CO_2, O_2, H_2O, blood, Fe_2O_3 (rust), O_3

Element: Au, O_2 ✓, O_3 ✓ ← ozone ← uv

2 Diatomic elements Br_2 I_2 N_2 Cl_2 H_2 O_2 F_2

("Brinclhof")

N–N H–H

H_2O → H_2 & O_2

Compound: ≥2 different elements chemically combined (e- sharing / stealing)

Mixtures: ≥2 diff. substances

Homogeneous: uniform composition

Examples: gatorade, air (N_2, O_2), alloy ← 2+ metals

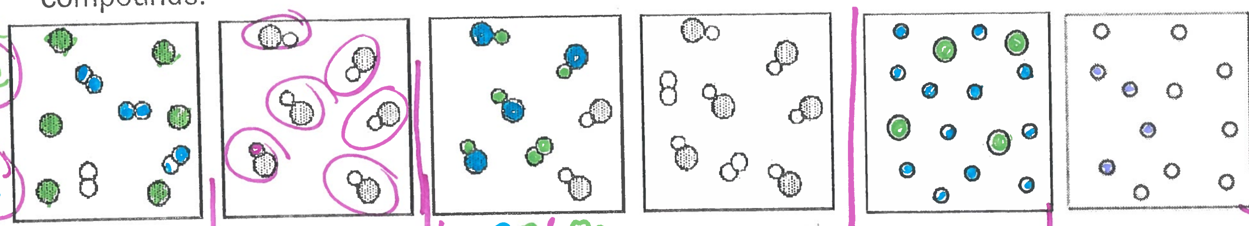
Heterogeneous: not uniform.

Examples:

Identify the following as pure element, pure compound, mixtures of elements and/or compounds.

monatomic Ne

diatomic N_2



CO & O_2

M.O. elements | pure compound | mixtures of C & E | He, Ar M.O. elements! | pure element

B. Phases of Matter and $\boxed{\text{Kinetic}}$ Theory

motion

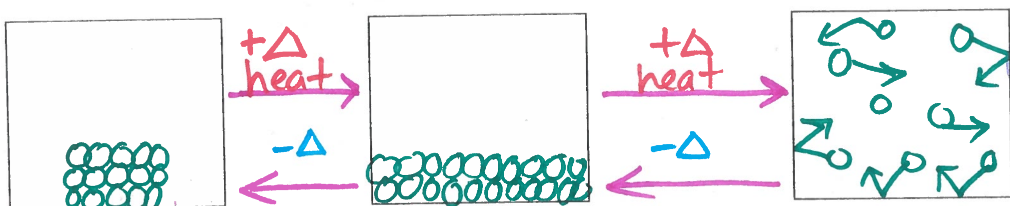
Solid: def. shape ; def. volume (slow motion, in-place)

Liquid: indef. shape ; def. vol. (slide past) each other

Gas: indef. shape ; indef vol. (fast! far apart.)

fluids "flow"

Solid Liquid Gas



+Δ heat → +Δ heat →

← −Δ ← −Δ

$P = \frac{F}{A}$

SLOW motion

FAST MOTION

~~gases~~ gasses

Which phases can you compress (decrease the volume)? gases

Plasmas: "superheated gas"

e^- & nuclei are seperated → {lightning, stars, flames}

Substances change phases as temperature increases.

motion

Kinetic Theory: hot things have faster particles.

SFN: Intermolecular Forces

→ Why do substances change phases? heat added or removed

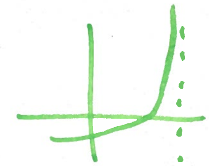
Temperature Scales

Celsius Scale

0 °C: = 32°F FP of H_2O

100 °C = 212°F BP of H_2O

$E=mc^2$



No motion of particles

Kin Energy! Kelvin Scale

0 K: "absolute zero" ~~[illegible]~~ Lowest theoretical temp.

273 K: = 0°C

Converting between Celsius and Kelvin

°F
°C
K

Equation: K=°C + 273

Kelvin	0	273	373
Celsius	-273°C	0	100

BP of H_2O

Convert the following

20°C = +273 293 K

300K = 27 °C

327°C= 600 K +273

OYO

-10C=____________K

250K = ____________ °C

9/9

9/12

C. Physical vs. Chemical Properties and Changes

Physical property: properties that ~~[illegible]~~ can be observed w/o changing the substances

Examples: weight, color, shape, BP, MP, FP, ~~melts~~, hardness, malleable, density, ductile

Chemical Property: ability of substance to form a new compound or react.

Examples: flammable, reactive (stability), shock sensitivity, oxidative, acidity.

Physical Changes: ~~[illegible]~~ keeps identity (KEEPS FORMULA)

Examples: dissolve, boil, melt, bend, break

Chemical Changes: new substance is formed (new formula)

Examples: heat, reaction, decompose, synthesize, fermentation (alcohols), burn, gas, color changes, digest.

D. Basics of Chemical Reactions:

Reactants → Products

Example: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

Indicators: heat changes, new smell, light emitted, bubbles (gas), new solid forms.

Scientific Measurement and Math

A. Measurement uncertainty for a single measurement.

avg of msmts — actual

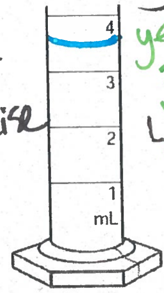
Accuracy: How close is the msmt to the real one?

Precision: How many digits can you read off of instrument

162.4 lbs

162.385 lbs precise

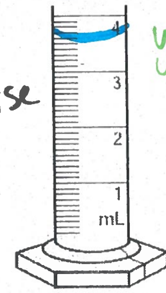
less precise



yes! guess? 3.5 — 2 sf

Cylinder A

more precise



yes! yes! Guess? 3.51 — 3 sf

Cylinder B

Measurement uncertainty for a set of measurements.

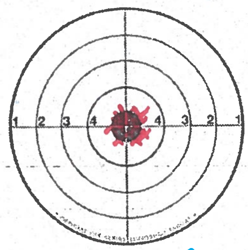
$\bar{X}$ Accuracy: How close is the avg to the {real / true / actual} value?

Δ Precision: How close are msmts to each other?

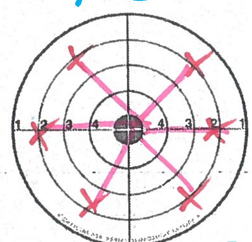
Not ACC

Not ACC

ACC



ACC.



consistent → PRECISE — NOT PRECISE — PREC. — NOT PRECISE

Example: Three students are determining the density of a sample of silver, Ag. The accepted density of silver is 10.50 g/cm³. Which student is most accurate? Which student is most precise?

	Julie	Robert	Terry
Trial 1	10.54 g/cm3	10.61	10.44
Trial 2	10.46 g/cm3	10.60	10.51
Trial ~~#~~3	10.47 g/cm3	10.62	10.55
$\bar{X}$ Average/Mean	10.49 g/cm³	10.61 g/cm³	10.50 g/cm³
Δ Range	0.08	0.02	0.11

Precise

Accurate!!!

$$\%\,err = \frac{|EV - TV|}{TV} \times 100$$

you / someone else

Percent Error: $\frac{|\text{experimental value} - \text{true value}|}{\text{true value}} \times 100$

Example: My bathroom scale indicates that I weigh 135 lbs (EV). The calibrated Doctor's scale says 142 lbs (TV). What is the percent error of my scale?

$$\%\,err = \frac{|EV - TV|}{TV} \Rightarrow \frac{|135 - 142|}{142} \Rightarrow \left(\frac{7}{142} \times 100\right) \approx 4.9\%$$

A student uses a ruler to determine a circle has a diameter of 3.8 centimeters. The true diameter is 3.7 centimeters. What is the student's percent error? *(Ans = 2.7%)*

Calibration: making sure an instrument is accurate by known standard. (like tuning a guitar string to a note)

9/12

9/13 B. Scientific Calculation Basics

1) Scientific Notation: expresses BIG or SMALL #s

only one non-zero digit before decimal point

$\boxed{1.25} \times 10^2$ NOT 12.5×10^1

mantissa

$10^1 =$ 10 $\qquad 10^2 =$ 100 $\qquad 10^3 =$ 1000

$10^{-1} = \frac{1}{10} = 0.1 \qquad 10^{-2} = \frac{1}{100} = 0.01 \qquad 10^{-3} = \frac{1}{1000} = \cancel{0.000}$ 0.001

$10^0 =$ 1

converting decimal notation to scientific notation

1. Count the number of places you move the decimal point = exponent
2. If the |number| is greater than 1: positive exponent

   If the |number| is less than 1: negative exponent

Examples:

123,000 = 1.23×10^5 (54321) $\qquad$ 0.0047 = 4.7×10^{-3} $\qquad$ -420 = -4.2×10^2

Converting scientific notation to decimal notation

1. Move the decimal point to make the number smaller if the exponent is negative
2. Move the decimal point to make the number larger is the exponent is positive

Examples: $\quad 4.5 \times 10^{-3} =$ 0.0045 $\qquad 7.4 \times 10^4 =$ 74,000.

0.0045

$3x^{2}+4$

mantissas

multiplying: multiply coefficients; add exponents.

$(Y\times10^{A})\cdot(Z\times10^{B}) = (Y\cdot Z)\times10^{A+B}$

Examples: +

$(2\times10^{2})(3\times10^{3})$ = 6×10^{5} $(3\times10^{-2})(1.5\times10^{-1})$= 4.5×10^{-3}

$(3\times10^{-10})(5\times10^{4})$ = 15×10^{-6} → must be fixed

e)dividing: divide ~~coe~~ coefficients; subtract exponents

$\frac{Y\times10^{A}}{Z\times10^{B}} \rightarrow (Y/Z)\times10^{A-B}$

$\frac{6\times10^{3}}{2\times10^{-4}} \rightarrow 3\times10^{7}$

correcting scientific notation:

only one digit in front of the decimal point is allowed.

SFN

15×10^{-6} = ____________________

0.073×10^{4} = ____________________

Convert to scientific notation:

oeo 235 = ______________ 0.0521= ______________ 102,400= ______________

Convert to decimal notation:

oeo 1.2×10^{-4}= ______________ 4.2×10^{3}= ______________

Solve:

oeo $(3\times10^{2})(3\times10^{4})$= ______________ $(8\times10^{4})/(2\times10^{-2})$= ______________

oeo $(3\times10^{3})(4\times10^{-5})$= ______________

5.6×10^{-7}

$\boxed{5}\ \boxed{.}\ \boxed{6}\ \boxed{E}\ \boxed{-}\ \boxed{7}$

"times 10 to the ___"

$\times10$

Using your scientific calculator

Solve $(3.0\times10^{4})\times(7.2\times10^{-9})$

TI-30XA enter 3.0 [EE] 4 x 7.2 [EE] [(-9)] =

TI Graphing Calculator enter 3.0 [2nd] [EE] 4 x 7.2 [2nd] [EE] [(-)] 9 ENTER

4.92×10^{17}

4.92E17

160.5

→ 160.45793

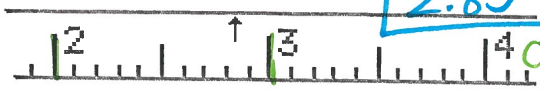
Significant Figures: digits that indicate a measurement's or calculation's precision.

For measurement equipment, always estimate one digit beyond the last division. The estimated digit is the last significant digit. For electronic equipment, the last displayed digit is significant.

Examples: Read measurement equipment using significant digits:

Y Y G

2.83

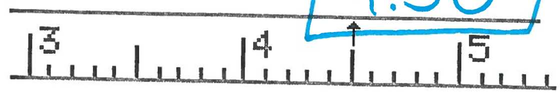


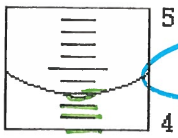
cm

~~4.332697~~

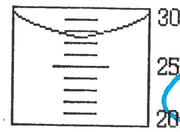
Y Y G

4.50

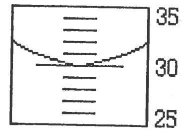




4.27



27.5



30.0

010.0 mL

Math with significant digits:

1. Leading zeros never count
2. Trailing zeros only count if there's a decimal point
3. Exact counts and conversion factors have an infinite number of significant digits:

007 mL

1 sf 700 vs 700.0 4 sf

00000000.024 ~~0.024~~

Examples:

23 has 2 significant digits

203 has 3 significant digits #CentralZeroesMatter #CZM

0.0203 has 3 significant digits

2030 has 3 significant digits

.5 mg

2030.0 has 5 significant digits.

5 mg

2.0×10^{-3} has 2 significant figures

mantissa

4 sf

4.700×10^{4}

0.5 mg

When multiplying or dividing, the answer is rounded to the same number of significant digits as the factor with the least number of significant digits. Use scientific notation if you get stuck.

Example: 3.0 x 3 = 9, but 3.0 x 3.0 = 9.0

9 x 5.0

"weakest link"

7.0 x 5.0 = 35, but 7 x 5.0 = 40 (35 rounds to 40)

45

50

$ 100,000

→ 5 x 8 = 40, but 5.0 x 8.0 = 40.

$ 100,100.72

mere

SFN → 5.0 x 80.0 = 400, correct to 4.0×10^2

When adding or subtracting, the final answer should be rounded to the least number of decimal places.

9	9	8	10.27	2200
+ 2.1	+ 2.6	+ 2.1	+ 9.4	+ 15
11.1	11.6	10.1	19.67	2215
11	12	10	19.7	2200

0 vs 1

Unit Canceling Method(A.K.A. Dimensional Analysis or Factor-Label)

Unit Canceling Method: ____

____

Some math terms:

$$\frac{4 \text{ quarts}}{1 \text{ gallon}}$$

Numerator: Denominator: Coefficents:

Units:

Parking lot problem: I have 22 quarters, but I want nickels. How many nickels should I get?

Given: Find: Know:

Side Street Problem: How many teaspoons are in 3.2 cups?

Given: Find: Know:1 cup=16 Tbs, 1 Tbs=3 tsp

Main Street Problem: How many feet are in 0.41 meters?(Ans = 1.349 ft = ________ w/ sig figs)

Given: Find: Know: 1 inch = 2.54 cm and 1 m = 100 cm

Metric System Units for Chemistry

	Length	Volume	Mass
Base unit	meter	Liter	gram
Abbrev.	m	L	g
Common chemistry units	m, mm, cm, μm, nm	L, mL, μL	Kg, g, mg, μg

Metric System Prefixes (using meter as base system)

Number of meters, liters, or grams	prefix	Abbeviation with meter	Written as a power of 10
1000	kilo	km	1 km = ______________ m
100	hecto	hm	1 hm = ______________ m
10	deka	dkm	1 dkm = ______________ m
1	base unit (m, L, g.)		
0.1	deci	dm	1 dm = ______________ m
0.01	centi	cm	1 cm = ______________ m
0.001	milli	mm	1 mm = ______________ m

Conversions to memorize (using meters as example)

1000 m = 1 km   10 dm = 1 m   100 cm = 1 m   1000 mm = 1 m   1 cm = 10 mm

1 Liter = _________ mL   1 kg = ____________g

Metric Conversions with Unit Analysis

Convert 320 mm to _____ m.   Given: Find:

Convert 3.23 kilograms to grams
Given: Find:

Km → m → cm

OYO

A student ran 5.8 km. How many centimeters did the student run?

Given: Find:

$$\frac{5.8\ \cancel{km}}{1} \times \frac{1000\ \cancel{m}}{1\ \cancel{km}} \times \frac{100\ cm}{1\ \cancel{m}} = \_\_\_\_\_\_\_\_\ cm$$

Convert 8.2×10^8 mg to kg

Given: Find:

$$\frac{8.2E8\ \cancel{mg}}{1} \times \frac{1\ \cancel{g}}{1000\ \cancel{mg}} \times \frac{1\ kg}{1000\ \cancel{g}} = \underline{820}\ kg$$

9/19 ~~9/16~~

More about units

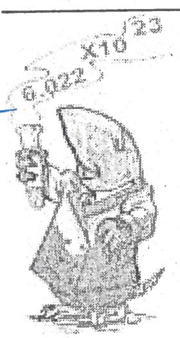
Volume: one Liter = $1 dm^3$ by definition and 1 mL = $1 cm^3$

so 1 L = ________ mL = ________ cm^3

Mass: 1 kilogram: 1 L of H_2O @ $\boxed{4°C}$

1 gram 1 mL of H_2O @ 4°C

The Mole: 6.02E23 aka 6.02×10^{23} "things" (atoms, molecules, or protons, etc.)



1 dozen = 12 things

1 trio = 3 things

1 ream = 500 thing

1 grand = 1000 things

1 score = 20 things

1 mole = $\underbrace{6.02 \times 10^{23}}_{\text{huge}}$ things

~~moles.~~

"Start w/ what you know"

Unit Cancelation and the Mole

We know a dozen equals 12 of anything. We know a trio of singers means three singers. Chemists wanted a similar convenient term to count atoms and molecules. They came up with the term <u>mole</u>. One mole = 602,000,000,000,000,000,000,000 of things.

1 mole = 6.02 x10^{23} representative particles or

$$\frac{1\ \text{mol}}{6.02 \times 10^{23}\ \text{rep. part}} = \frac{6.02 \times 10^{23}\ \text{rep. part.}}{1\ \text{mol}}$$

3

8.25 dozen eggs = 99.0 eggs

$$\frac{8.25\ \cancel{dz}}{1} \times \frac{12\ \#\text{ofeggs}}{1\ \cancel{dz}}$$

8.25 moles of eggs = 4.97×10^{24} eggs

$$\frac{8.25\ \cancel{mol}}{1} \times \frac{6.02E23\ \text{eggs}}{1\ \cancel{mol}}$$

220,000 doughnuts = ______ dozen doughnuts

$$\frac{220{,}000\ \cancel{dnts}}{1} \times \frac{1\ dz}{12\ \cancel{dnts}}$$

220,000 doughnuts = 3.7×10^{-19} moles of doughnuts

$$\frac{220{,}000\ \cancel{dnts}}{1} \times \frac{1\ mol}{6.02E23\ \cancel{dnts}} = 3.65\ E{-}19$$

4

$\boxed{0.04221}$ moles of iron atoms = ______ iron atoms

$$\frac{0.04221\ \cancel{mol}\ Fe}{1} \times \frac{6.02E23\ \text{Fe atoms}}{1\ \cancel{mol\ Fe}}$$

4.5 x 10^{26} sodium atoms = 7.5×10^{2} moles of sodium atoms

$$\frac{4.5E26\ \text{Na}\ \cancel{atoms}}{1} \times \frac{1\ \text{mol Na}}{6.02E23\ \text{Na}\ \cancel{atoms}}$$

OYO

3.01 x 10^{-4} moles of water molecules, H_2O, = ______ water molecules

$$\frac{3.01E{-}4\ \cancel{mol\ H_2O}}{1} \times \frac{6.02E23\ \text{molecules}\ H_2O}{1\ \cancel{mol\ H_2O}}$$

8 x 10^{20} potassium atoms = ______ moles of potassium atoms

$$\frac{8E20\ \text{K atoms}}{1} \times \frac{1\ \text{mol K}}{6.02E23\ \text{K atoms}}$$

Calculating Density

Density is an intrinsic physical property of a substance.

Example: Au, ________ , density = 19.3 g/cm^3 and Al, ________________, density = 2.7 g/cm^3

Equation: unit =

Example 1: A 4.8 gram sample of grey metal has a volume of 3.9 cm^3. What is the metal's density?

Example 2: What is the mass of a pine block measuring 2.0 x 3.0 x 6.0 cm with a density of 0.50 g/cm^3.

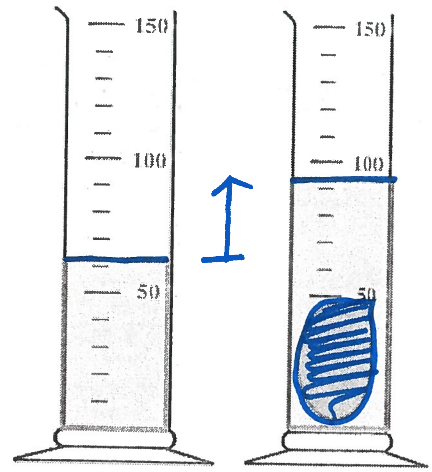
Example 3: What is the volume of a gold bar with a mass of 1.81 x 10^4 grams. Au's density = 19.3 g/cm^3

Approach 1—use equation.

Approach 2—use unit cancelation and density as a conversion factor.

Density by Displacement.

A ingot of unknown metal with a mass of 241 grams is dropped into a graduated cylinder containing _________ mL of water. The water level rises to _________ mL. What is the density of the unknown metal?



A machinist needs to identify if an unlabeled box of screws is made of aluminum or stainless steel. The machinist puts 15 screws with a mass of 28 grams into a graduated cylinder that contains 20.0 mL of water. The water level rises to 30.4 mL. Steel has a density of 8.0 g/cm^3 whereas aluminum has a density of 2.7 g/cm^3. What are the screws made of? Justify your answer using a calculation.

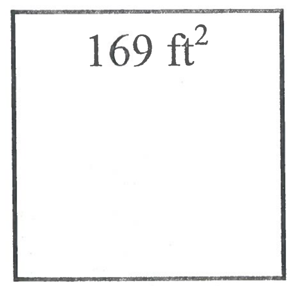
More Dimensional Analysis Practice

1) Determine how many milligrams (abbreviation: mg) are in 3.21 lbs of lead. (1 lb = about 2.204 kg)

$$\frac{3.21 \text{ lbs Pb}}{1} \times \frac{\text{kg Pb}}{\text{lbs Pb}}$$

2) Earth is 1 "astronomical unit" away from the Sun. (1 AU is 150,000,000 km, by the way) Jupiter is 5.2 AU away from the Sun. How many miles is Jupiter from the Sun? (1 mile = 1.609 km)

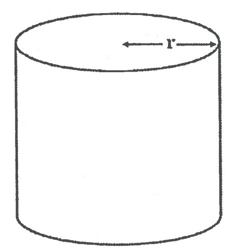
3) The area of this square garden is 169 cubic feet. What is the area in cubic meters? (1 foot = 12 inches. 1 inch = 2.54 cm)



4) 1 mL is a volume unit that is equivalent to 1 cubic centimeter (cm^3). This cylinder has a radius of 3.84 cm, and a height of 12.57 cm.

a. Determine the volume of the cylinder in cubic centimeters.

$$V = \pi r^2 h$$



b. Determine the volume of the cylinder in milliliters.

c. Determine the volume of the cylinder in liters. Use scientific notation.

d. Determine the volume of the cylinder in ounces. (1 oz = about 29.57 mL)