Chemistry Unit 1
Primary reference: *Chemistry: Matter and Change* [Glencoe, 2017]

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**. **Percent Error** = 100 x |accepted value-exper. value| /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, and page 976. **Ch 2:** Read pp. 40-41 (and pp. 946-948) on scientific notation. Read pp. 47-49 on accuracy, precision, and percent error. Read pp. 50-54 (and pp. 949-953) on sig. figs. Read pp. 32-38 on the units and metric system. Read pp. 44-46 (and pp. 956-958) on Unit Canceling / Dimensional Analysis / Unit Conversion.
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. Density = 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 3:** Read pp. 84-90 on elements & compounds. Read pp. 76-77 on chemical and physical changes. Read pp. 80-82 on mixtures.
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 3:** Read pp. 84-90 on elements & 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. 1 mole = $6.02\ 10^{23}$ (atoms or molecules)	**Ch 10:** Read pp. 320-324 on the introduction of 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. K = °C + 273	**Ch 3:** Read pp. 70-73 on introduction to states of matter. **Ch 2:** Read pp. 34-35 on temperature and Kelvin

Objectives for Unit One

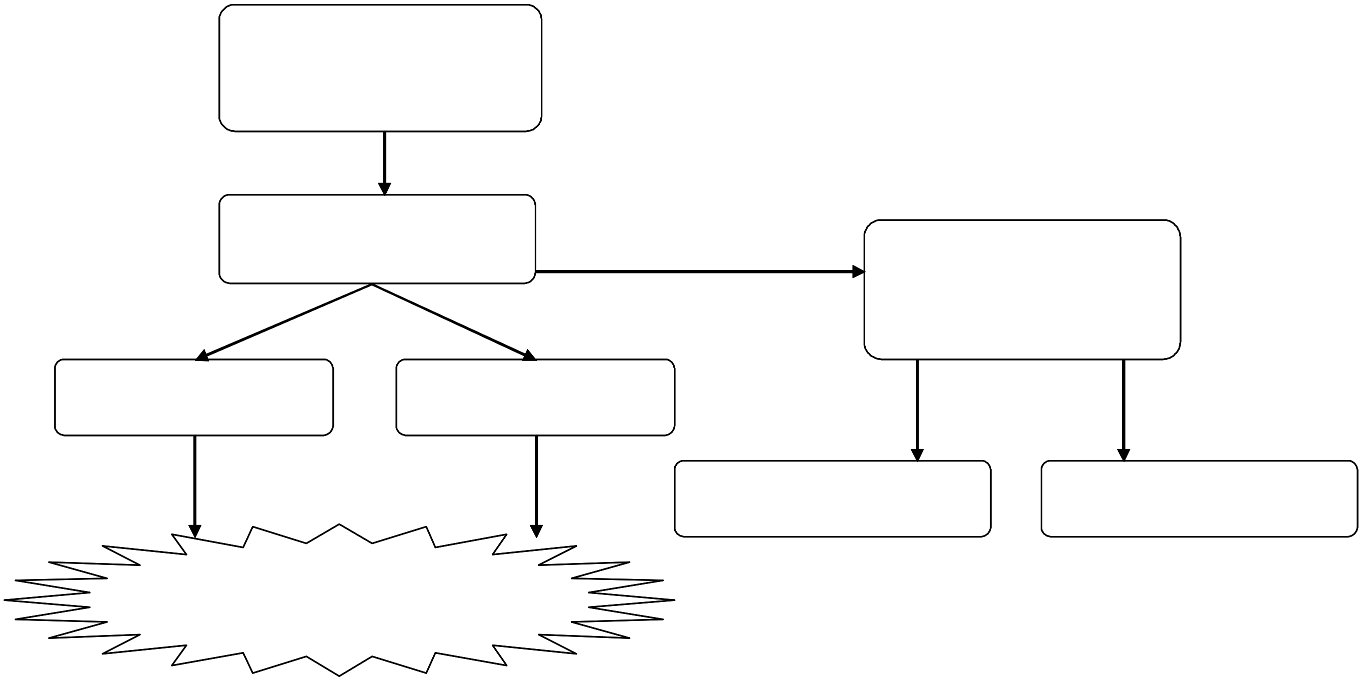
Chemistry: Matter and Change (Glencoe, 2017)

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
4. Memorize the seven diatomic elements (BrINClHOF)
5. Differentiate between chemical and physical properties and changes
6. Understand the basic differences between a gas, liquid, and solid in terms of kinetic theory
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
10. Determine the number of significant figures in any number
11. Use significant figures correctly in multiplication, and division problems
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.
14. Memorize and convert between metric unit prefixes (kilo, centi, milli)
15. Memorize that 1 mole = 6.02×10^{23} particles
16. Explain the difference between precision and accuracy
17. Calculate percent error from word problems
18. Memorize and use the density equation (D=m/v) to calculate density, mass, or volume from word problems.
19. Use the unit cancellation method to convert between units and measurements in word problems



Intro to Chemistry

A. Types of Matter

Matter:____

Mass:____

Substance:____

Examples: ____

Element: ____

Diatomic elements ____

____

Compound:____

____

Mixtures:____

____

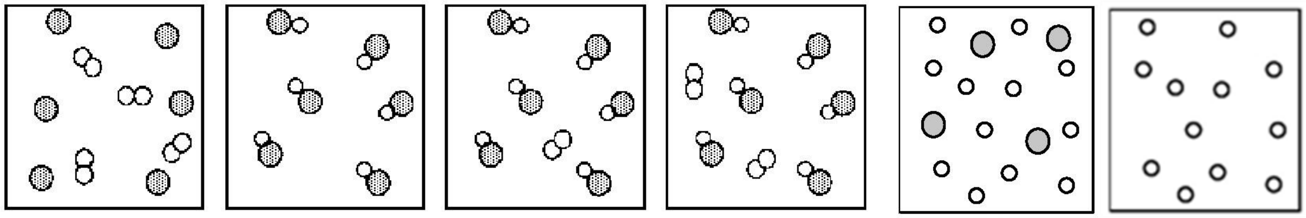
Homogeneous:____

Examples:

Heterogeneous:____

Examples:

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



B.Phases of Matter and Kinetic Theory

Solid:____

Liquid:____

Gas:____

Solid Liquid Gas

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

Plasmas: ____

____

Substances change phases as temperature increases.

Kinetic Theory:____.

Intermolecular Forces____

Why do substances change phases?____

____

Temperature Scales

Celsius Scale

0 ºC:____

100 ºC____

Kelvin Scale

0 K:____

273 K:____

Converting between Celsius and Kelvin

Equation: $K = {}^{o}C + 273$

Kelvin	0		
Celsius		0	100

Convert the following

20ºC = ____________ K

-10C=_____________K

300K = _____________ °C

250K = ___________ °C

327ºC=______________ K

C. Physical vs. Chemical Properties and Changes

Physical property:____

____

Examples:

Chemical Property:____

____

Examples:

Physical Changes:____

Examples:

Chemical Changes:____

____

Examples:

D.Basics of Chemical Reactions:

Reactants → Products

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

Indicators: ____

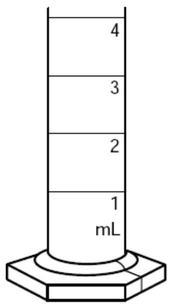
____

Scientific Measurement and Math

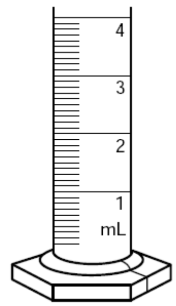
A. Measurement uncertainty for a single measurement.

Accuracy:____

Precision:____



Cylinder A

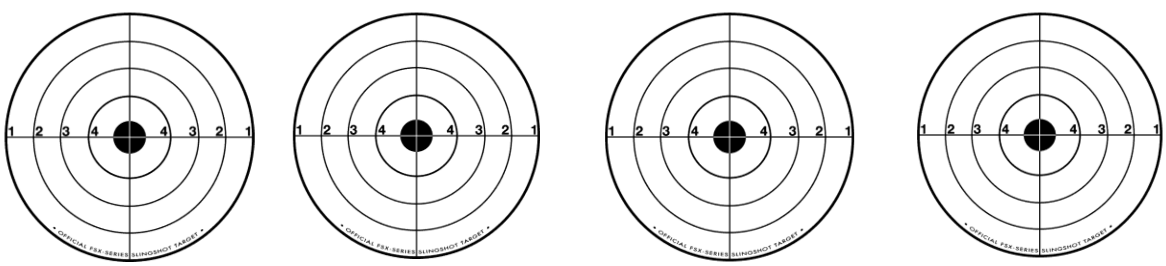


Cylinder B

Measurement uncertainty for a set of measurements.

Accuracy:____

Precision:____



Example: Three students are determining the density of a sample of silver, ____. The accepted density of silver is 10.50 g/cm^3. 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 1	10.47 g/cm3	10.62	10.55
Average/Mean			
Range			

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

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

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:____

____

B. Scientific Calculation Basics

1)Scientific Notation:____

only one non-zero digit before decimal point

1.25×10^2 NOT 12.5×10^1

10^1 = _____ 10^2=_____ 10^3=________

10^{-1} = ______ 10^{-2}=______ 10^{-3}=_________

10^0 = ______

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: ___________________ exponent

   If the │number│is less than 1: __________________ exponent

Examples:

123,000 = _____________ 0.0047 = ______________ -420 = ___________

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: 4.5×10^{-3} = ________________ 7.4×10^4 = ___________

multiplying:____

Examples:

$(2x10^{2})(3x10^{3})$ = ______________ $(3x10^{-2})(1.5x10^{-1})$=_____________

$(3x10^{-10})(5x10^{4})$ = __________________

e)dividing:____

correcting scientific notation:

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

15×10^{-6} = ____

$0.073 \text{ X } 10^{4}$ = ____

Convert to scientific notation:

235 = __________________ 0.0521=__________________ 102,400=______________

Convert to decimal notation:

1.2×10^{-4}=_______________ $4.2x10^{3}$=_______________

Solve:

$(3x10^{2})(3x10^{4})$=________________ $(8x10^{4})/(2x10^{-2})$=___________________

$(3x10^{3})(4x10^{-5})$=______________________

Using your scientific calculator

Solve $(3.0 \times 10^{4})x(7.2 \times 10^{-9})$

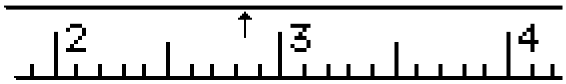
TI-30XA enter 3.0 [EE] 4 x 7.2 [EE] [(-9)] =

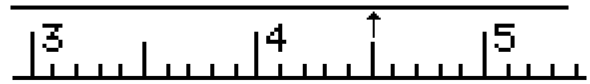
TI Graphing Calculator enter 3.0 [2nd] [EE] 4 x 7.2 [2nd] [EE] [(-)] 9 ENTER

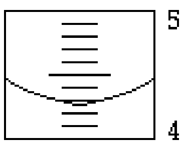
Significant Figures: digits that indicate a measurement's or calculation's precision.

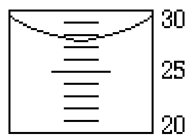
For measurement equipment, always estimate one digit beyond the last division. The estimated digit is the last significant digit. For electronic equipment, the last displayed digit is significant.

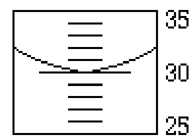
Examples: Read measurement equipment using significant digits:











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:

Examples:

23 has 2 significant digits

203 has 3 significant digits

0.0203 has 3 significant digits

2030 has 3 significant digits

2030.0 has 5 significant digits.

2.0×10^{-3} has 2 significant figures

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

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

5 x 8 = 40, but 5.0 x 8.0 = 40.

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

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			
Abbrev.			
Common chemistry units			

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:

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

Given: Find:

Convert 8.2×10^8 mg to kg

Given: Find:

More about units

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

so 1 L = ________mL = ________cm^3

Mass: 1 kilogram:____

1 gram____

The Mole:____

____

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 **mole**. One mole = 602,000,000,000,000,000,000,000 of things.

1 mole = 6.02×10^{23} representative particles or

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

8.25 dozen eggs = ___________________________ eggs

8.25 moles of eggs = _____________________________eggs

220,000 doughnuts = ___________________________ dozen doughnuts

220,000 doughnuts = _____________________________moles of doughnuts

0.04221 moles of iron atoms = ___________________________ iron atoms

4.5×10^{26} sodium atoms = ______________________moles of sodium atoms

3.01×10^{-4} moles of water molecules, H_2O, = ________________________ water molecules

8×10^{20} potassium atoms = ____________________________ moles of potassium 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.

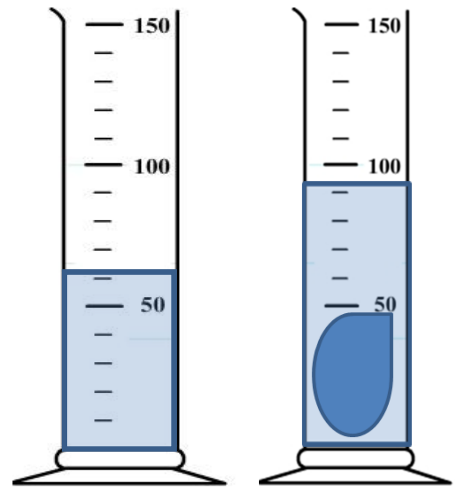
Example 3: What is the volume of a gold bar with a mass of 1.81 x 10^4 grams. Au's density = 19.3 g/cm^3

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?



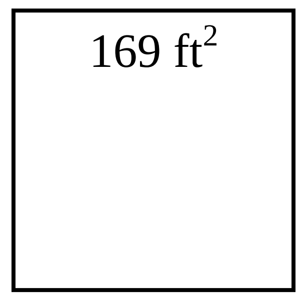
A machinist needs to identify if an unlabeled box of screws is made of aluminum or stainless steel. The machinist puts 15 screws with a mass of 28 grams into a graduated cylinder that contains 20.0 mL of water. The water level rises to 30.4 mL. Steel has a density of 8.0 g/cm^3 whereas aluminum has a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

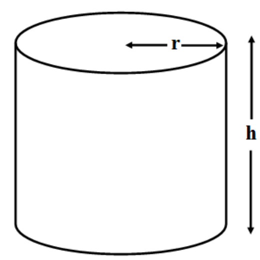
2) Earth is 1 "astronomical unit" away from the Sun. (1 AU is 150,000,000 km, by the way) Jupiter is 5.2 AU away from the Sun. How many miles is Jupiter from the Sun? (1 mile = 1.609 km)

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)