Chemistry 1 Fall Exam Review

Unit 1: Scientific Measurement:

*****REVIEW LAB SAFETY AND LAB EQUIPMENT*****

An example of a quantitative measurement is 5cm (number + unit) and qualitative is red, large (description)

Explain the difference between accuracy and precision:

accuracy - closeness to true value   precision - repeatability / more sig figs.

Significant figures rules:

1) Leading zeros never count.
2) Trailing zeros only count if decimal present.
3) Sandwiched zeros always count.

How many sig figs? 110 → 2
111.00 → 5

When multiplying or dividing, use the least number of sigfigs.
When adding or subtracting, use the least number of decimal place past.

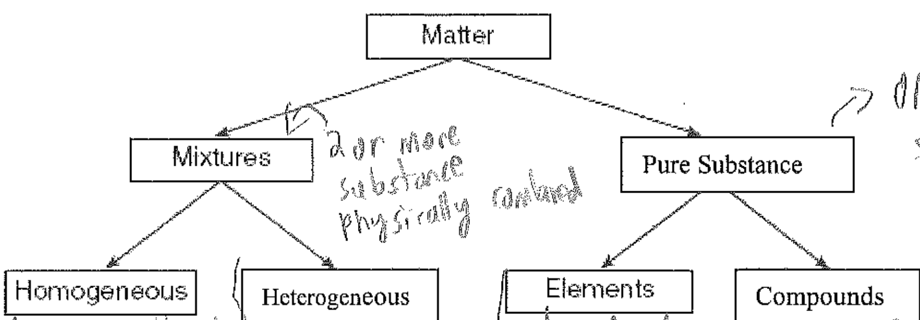
****Be able to solve problems using dimensional analysis****

****Know how to express a number in scientific notation!****

Unit 2: Describing Matter:

Define each of the following categories in the boxes AND give an example:

Classification of Matter



Mixtures: 2 or more substance physically combined

Pure Substance: ONLY ONE substance throughout

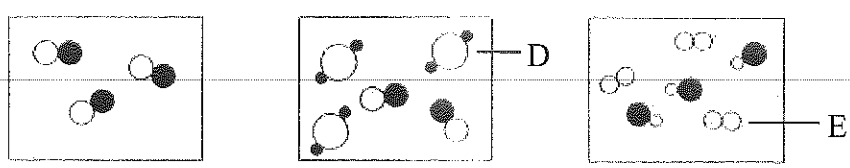
Homogeneous: Kool-Aid, Tea, salt water, Air, Brass, Sugar solution

Heterogeneous: salad, cereal, sand, water/oil

Elements: Na, O_2, Ag, Cl_2, Fe

Compounds: $NaCl$, H_2O, Fe_2O_3, MgO

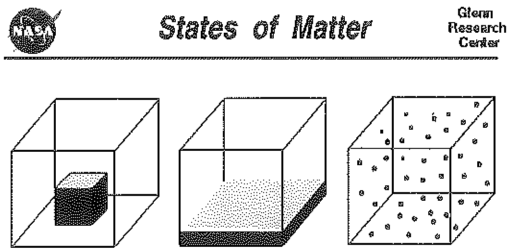
Classify these examples →



A B C

A Pure Sub → compound B Homogeneous C Homogeneous

D Compound E element

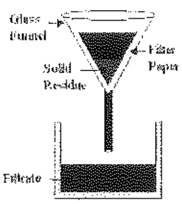


Comparing mixtures and compounds:

1. A __mixture__ retains the original properties of its components.
2. The properties of a __compound__ differ from the properties of its components.

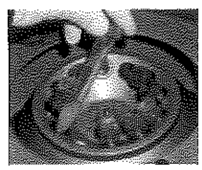
	Solid	Liquid	Gas
Shape	Definite	Indefinite	Indefinite
Volume	Definite	Definite	Indefinite
Particle attraction	High	middle	low

For each of the following separation techniques below, describe the process AND whether it's a physical or chemical separation:



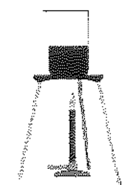
Filtration

-physical
-Liquid + undissolved solid
-Liquid passed through filter



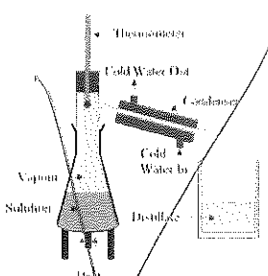
Centrifugation

-physical
-Liquid + solid
- High speed spinning
-Heavier density settles @ bottom

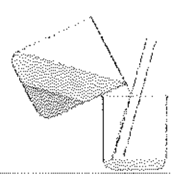


Evaporating

-physical
-Liquid + dissolved solid
- Heating solution evaporates liquid & leaves solid



Distillation



Decanting

-Physical
-Layered mixtures
-Pouring off top layer

Heating

- Chemical !
- Applying heat to break compounds

	Chemical	Physical
Examples of properties:	reactive, combustability, flammability, oxidation	color, density, hardness, conductivity, M.P., B.P, luster
Examples of changes:	oxidizing, digesting, heating (cooking)	DISSOLVING, phase changes

Explain the difference between an intensive and extensive property:

-Intensive property does not depend on the amount of substance. (color, luster, density)

-Extensive property does depend on the amount. (mass & volume)

Properties of metals include:

High melting point, good electrical conductivity and heat conductivity, luster, malleable, ductile.

Define malleable and ductile:

ductile ← ability to be pulled in thin wires.

malleable ← how easily it can be bent or shaped

4 Signs of a chemical change:

1. unexpected color change
2. heat or light produced
3. solid percipitate formation
4. gas formation

Unit 3: Atomic Structure:

Subatomic particle	Location in atom	Charge	Relative mass (amu)
PROTON p^+	NUCLEUS	+	1
NEUTRON n^0	NUCLEUS	0	1
ELECTRON e^-	e^- CLOUD	−	0

Fill in the blank:

The smallest particle of an element is a(n) atom. In an atom, the nucleus accounts for most of the mass and the e^- cloud occupy most of the volume. The atomic number is the total number of protons and the mass number is the total number of protons plus neutrons. The element is identified by the number of protons. In an isotope, the number of neutrons can change. The atomic mass on the periodic table is determined by a weighted average of all existing isotopes of the element. An ion has an unequal number of protons and electrons.

Label X, A, and Z to illustrate isotope notation:

$$^{A}_{Z}\mathrm{X}$$

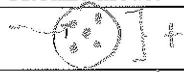
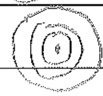
X element symbol A Mass $= p^+ + n^0$ Z atomic # $= p^+$

Explain how to calculate average atomic mass:

Avg. AM = Σ(abundance)(mass)

1) % → decimal form 2) Multiply % by mass of isotope 3) add up all the products for avg.

Complete the table:

	Atomic model	Experiment	Discovery
Thomson		Cathode Ray Tube	discovered e^-
Rutherford	 dense nucleus	Gold Foil	discovered p^+
Bohr		2-D Model	postulated the planetary model

Light behaves as both particles and waves. Light moves at a specific wavelength with a specific frequency which shows it behaves like a wave. Since light can be absorbed and released by electrons, this shows that light can behave as a particle and has a specific energy (or quantum) associated with it.

What is the base unit for wavelength? ___M___ For frequency? ___Hz or $\frac{1}{s}$___ For energy? ___J___

Solve the following problems (speed of light = 3.00×10^8 m/s, Plank's constant = 6.63×10^{-34} Js):

1. Light has a wavelength of 3.28×10^3 m. What is the frequency of light?

$\nu = ?\,Hz$ $\quad \lambda = 3.28 \times 10^3\,m$ $\quad c = \lambda\nu \rightarrow \underset{c}{3.00 \times 10^8\,m/s} = \underset{\lambda}{(3.28 \times 10^3\,m)}\,\nu \rightarrow \boxed{\nu = 9.15 \times 10^4\,Hz}$

2. Light is released by an electron with an energy of 5.34×10^{-14} J. What is the frequency of light?

$\nu = ?\,Hz$ $\quad E = 5.34 \times 10^{-14}\,J$ $\quad E = h\nu \rightarrow \underset{E}{5.34 \times 10^{-14}\,J} = \underset{h}{(6.63 \times 10^{-34}\,J \cdot s)}\,\nu \rightarrow \boxed{\nu = 8.05 \times 10^{14}\,Hz}$

ELECTROMAGNETIC SPECTRUM

Circle the correct response for the direction of the arrow:

Energy increases/decreases **Wavelength** increases/decrease **frequency** increases/decreases →

Radio/TV waves microwaves infrared visible light ultraviolet x-rays gamma rays

r o y g b i v

Unit 4: Electrons in Atoms

Explain what each of these rules state:

1. Aufbau's Rule: electrons fill orbitals from lowest energy state.
2. Pauli Exclusion Principle: electrons that fill same orbital must have opposite spin ↑↓.
3. Hund's Rule: electrons will spread out evenly in same orbitals before pairing up to minimize repulsion.

Circle the correct choice:

Electrons fill orbitals with (lowest/highest) energy first. They enter equal energy orbitals (singly/doubly). When they pair, it is with the (same/opposite) spin. When an electron moves from a lower to a higher energy level, it (absorbs/releases) energy.

Electrons in the outermost energy level are called ___valence___ electrons. The group number of the representative elements on the periodic table corresponds to the number of ___valence___ electrons. The dots on a Lewis dot diagram indicate the number of ___valence___ electrons. These electrons are located in ___s___ and ___p___ orbitals. Noble gases (except ___He___) have full valence orbitals with ___8___ electrons. An atom will gain or lose electrons to become ___Isoelectronic___ with the nearest noble gas.

Write the electron configurations for:

K $= 19e^-$ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$

K^{+1} $= 18e^-$ $1s^2 2s^2 2p^6 3s^2 3p^6$

Cl^{-1} $: 18e^-$ $1s^2 2s^2 2p^6 3s^2 3p^6$

isoelectronic

Determine the Lewis Dot for oxygen: $8e^- = 1s^2 2s^2 2p^4$ valence electrons $= 6e^-$

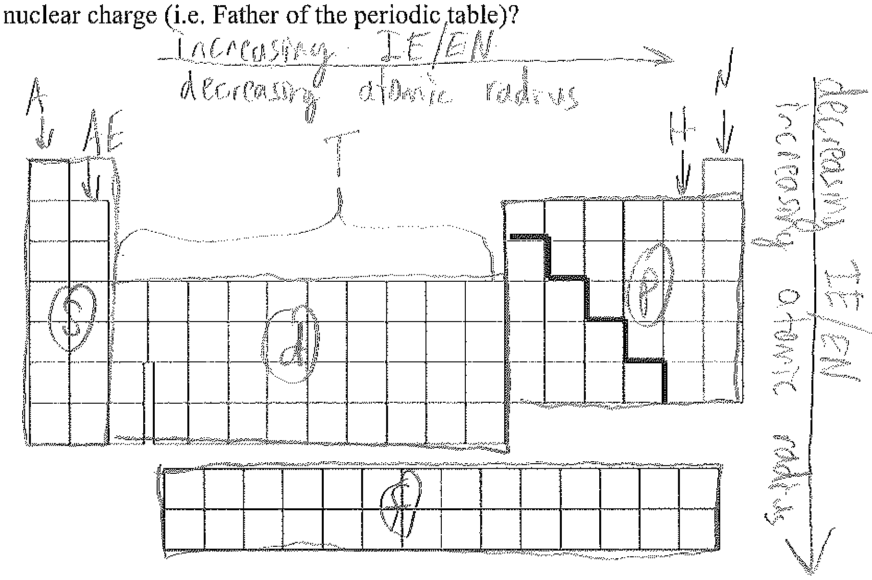
Ö

-Who first arranged elements in a table according to their properties and atomic mass (i.e. Grandfather of the periodic table)? Mendeleev

-Who arranged the modern periodic table based on nuclear charge (i.e. Father of the periodic table)? Moseley

Label the periodic table:

1. Color the metals red, the nonmetals blue, and the metalloids yellow
2. Label the alkali metals (**A**), alkaline earth metals (**AE**), transition metals (**T**), halogens (**H**), and noble gases (**N**)
3. Label a period and a group
4. Label the s, p, d, and f blocks
5. Label the trends for:
   - Ionization energy (IE)
   - Atomic radii
   - Electronegativity (EN)



Define effective nuclear charge:

-charge or pull felt by valence e- by p+

- increases across a period.

Explain electron shielding:

- the "pull" being blocked by valence e- due to inner electrons shielding.

Metals tend to __lose__ valence e^- and become __+__ ions called __cations.__
Nonmetals tend to __gain__ valence e^- and become __−__ ions called __anions.__

Ionic bonds form between __metal (+)__and __nonmetal (−)__ Electrons are __transferred.__ The bonded atoms are held together by the attraction of oppositely __charged__ ions. An ionic compound is also called a __formula unit__

Covalent bonds form between __nonmetal__. ~~Electrons are ______. When electrons are shared unequally the bond is ______. When they are shared equally, as in diatomic elements, the bond is ______.~~

Name the 7 diatomic elements: H_2 N_2 O_2 F_2 Cl_2 Br_2 I_2

Unit 5: Nomenclature:

Covalent compounds:

Name the first nonmetal with a __prefixes__ (except mono) followed by the second nonmetal with a __prefix__ and an __ide__ ending. DO NOT CRISS-CROSS OR REDUCE!

Give the prefixes:

1: mono 2: di 3: tri 4: tetra 5: penta 6: hexa

7: hepta 8: octa 9: nona 10: deca

Ionic compounds without a polyatomic:

Name the cation followed by the anion with an -ide ending. Criss-cross charges and REDUCE if possible!

What metals DO NOT need roman numerals?

groups 1+2, Cd^{+2} Zn^{+2} Ag^{+1} Al^{+3} (univalent metals)

Which metals need roman numerals?

multivalent → ALL other metals!

Ionic compounds with a polyatomic:

Name the metal followed by the name of the polyatomic ion. Do not change the ending – usually -ate or -ite. If writing the formula criss-cross charges and REDUCE if possible. The only polyatomic with a positive charge that will substitute for the metal is NH_4^+.

Unit 6: Mole Calculations

Explain how to calculate the following:

1. Molar mass: Multiply average atomic mass by # of each atom then add together ex: magnesium phosphate = $Mg_3(PO_4)_2$ { Mg 3(24.31) P: 2(30.97) O: 8(16.00) } $\boxed{262.85 \text{ g/mol}}$

2. % composition from a chemical formula:

$\% = \frac{part}{whole} \cdot 100$

3. Empirical formula from % composition:

1) convert g → mol
2) Divide by smallest # to get whole #

4. Molecular formula from the empirical formula:

$n = \frac{\text{molecular formula mass}}{\text{empirical formula mass}}$ → multiply empirical formula by "n"

What is the difference between an empirical and a molecular formula?

Most reduced   doesn't have to be reduced

Calculate the following mole conversions using dimensional analysis:

5. 6.8 g of CO_2 to L $6.8 g\ CO_2 \cdot \frac{1 mol\ CO_2}{44.01 g\ CO_2} \cdot \frac{22.4 L\ CO_2}{1 mol\ CO_2} = \boxed{3.5 L\ CO_2}$

6. 0.32 moles of $Ba(NO_3)_2$ to grams

$0.32 mol\ Ba(NO_3)_2 \cdot \frac{261.34 g\ Ba(NO_3)_2}{1 mol\ Ba(NO_3)_2} = \boxed{83.63 g\ Ba(NO_3)_2}$

7. 5.6 g of H_2O to molecules

$5.6 g\ H_2O \cdot \frac{1 mol\ H_2O}{18.02 g\ H_2O} \cdot \frac{6.02 \times 10^{23} \text{ molecule } H_2O}{1 mol\ H_2O} = \boxed{1.9 \times 10^{23} \text{ molecule } H_2O}$

Unit 7: Chemical Reactions:

Define the law of conservation of mass: matter cannot be created or destroyed

Be able to balance chemical equations.

Name the types of reactions depicted below:

$A + B \rightarrow AB$ synthesis

$AB \rightarrow A + B$ decomposition

$AB + C \rightarrow AC + B$ single replacement

$AB + CD \rightarrow AD + CB$ double replacement

$C_xH_y + O_2 \rightarrow CO_2 + H_2O$ combustion

Which type of reaction always has products CO_2 and H_2O? combustion

Identify the state of matter of each of the following compounds: AgCl(s) O_2(g) H_2O(l) $CaNO_3$( ) aq

Write the balanced equation, predict, and state the phase notation for the following:

Calcium hydroxide reacts with magnesium chloride:

$Ca(OH)_{2(aq)} + MgCl_{2(aq)} \rightarrow Mg(OH)_{2(s)} + CaCl_{2(aq)}$