

Fall 2020 Honors Chemistry Final Review

Scientific Method – define the following

- Observation
- Hypothesis
- Experiment
- Data/Analysis
- Theory or Scientific Law

1. Honeybees play an important role in agriculture, by pollinating at least 90 different species of food plants, worth about 15 billion dollars in the U.S. alone. As a result, beekeeping or apiculture, is a multi-million dollar business. However, over the last year, honeybee populations are crashing (Colony Collapse Disorder) with hives dying off in less than a week, without an obvious cause. Ian Lipkin, a senior researcher in the Epidemiology (infectious disease) Department of Columbia University's Mailman School of Public Health believed the cause to be a pathogen (virus, bacteria or fungus) of some sort. Use the information given to answer the following questions.

- What is the initial observation?
- What is the hypothesis?
- Design an experiment to test your hypothesis.
- If you did your experiment and your hypothesis was supported, what would your data indicate?
- If you did your experiment and your hypothesis was NOT supported, what would your data indicate?

Scientific Notation – is a way of writing very large or very small numbers.

- To put a number into scientific notation, move the decimal point between the 1st two numbers so that it makes a number between 1 -10.
- Then, as $\times 10$ raised to a power of how many times you moved that decimal point. The power will be negative if the number is a decimal and positive if it is a large number.
- Examples,
 - – Mass of a proton =
0.000 000 000 000 000 000 000 00 000 16726178g
 - Move the decimal point between the 1 and the 6 so you get ...
1.6726178
 - Add $\times 10$ to that and count how many times you moved your decimal point AND you should get...
 - $1.6726178 \times 10^{-30}$ g
 - Mass of the Eiffel Tower = 7300000000 g
 - 7.3×10^9 g

2. Covert these measurements into or out of scientific notation.

- 0.00000678s
- 19100000J
- 0.000546g
- 12110000000°C
- 8.76×10^4 atm
- 3.42×10^{-4} mol

Nuclear Chemistry

3. Define the following (nuclear fission and nuclear fusion) and which one
- produces the most elements?
 - is used in nuclear reactors?
 - is for fuel for stars where all the elements heavier than hydrogen are created?
 - requires extremely high temperatures?

States of Matter

4. Match each property to its correct state of matter.

- solid
- liquid
- gas
- plasma

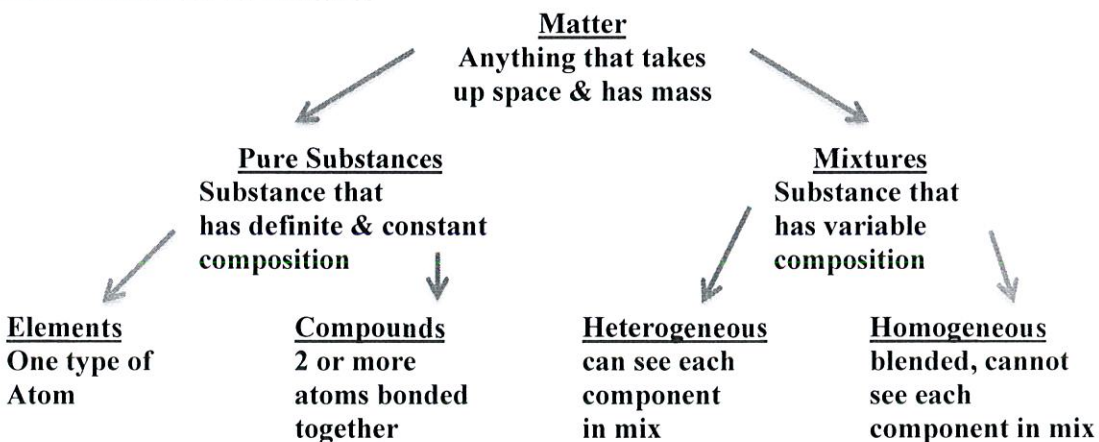
- close-particles that can slide past each other
- free-moving (random) particles that are relatively far apart from each other
- closed-packed particles that cannot do much more than vibrate in position
- free-moving particles in which the electrons have been ripped off the atom, creating charged particles, occurs at high temperatures.
- Definite shape and definite volume
- No definite shape or volume
- Definite volume but not shape

- **Physical Properties/Changes – can be observed without changing the substance's composition. (It's still the same substance!)**
- **Chemical Properties/Changes – can be observed when a substance changes composition. (Something new formed!)**

5. Determine if the following properties/changes are physical or chemical.

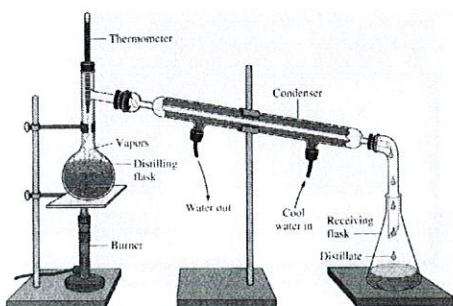
- Oxygen gas is odorless and colorless.
- Copper turns green when exposed to the environment.
- The density of water is 1.00g/cm^3 .
- Copper conducts electricity.
- Gold is nonflammable.
- Baking soda reacts with vinegar.
- The melting point of gallium is 85.6°F (30°C).
- The silver spoons tarnished and turned black.

Pure Substances vs. Mixtures

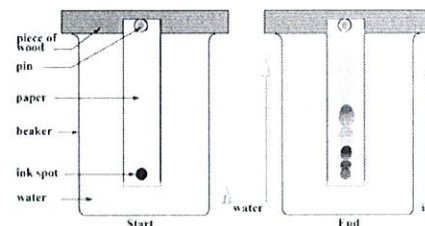


Examples

- element = iron, tungsten, chlorine
- compound = water, carbon dioxide, iron (III) chloride
- heterogeneous mix = Chex-mix, dirt
- homogeneous mix = air, soda, salt water
- Physical vs. Chemical
 - o Properties (nouns) & Changes (verbs)
 - Physical properties and changes can be observed **WITHOUT** changing the substance's composition
 - Examples – tearing, boiling, freezing, density, melting point, solubility
 - Can be used to separate mixtures by
 - o Magnets
 - o Size
 - o Solubility (ability to dissolve)
 - o Density
 - o Filtration
 - o Chromatography
 - o Distillation



Simple chromatography



- Chemical properties and changes can **ONLY** be observed by changing the substance into a completely **NEW** substance.
 - Examples – reactivity, flammability, rusting, digesting, decomposing, oxidizing

- **Can be used to change pure substances into new pure substances through chemical reactions**

6. Determine if the following substances are elements, compounds, homogeneous mixtures or heterogeneous mixtures:

- | | |
|--------------------|------------------------|
| a. copper | b. Lucky Charms cereal |
| c. air | d. oxygen |
| e. carbon monoxide | e. hot coffee |

7. Determine if the following are chemical properties/changes or physical properties/changes:

- A burning tire
- Freezing ice cubes
- The density of water is 1.0g/cm^3 .
- Rubidium will catch on fire if dropped in water.
- Platinum is a silvery-white color.
- Copper (II) chloride dissolves in water.

8. If gallium melts at 29.78°C and boils at 2403.0°C , what is its state of matter at it is in your hand? Keep in mind that your normal body temperature is 37.0°C .

Nomenclature (Names and Formulas!)

See nomenclature flow charts on the blog or the included graphic organizer and polyatomic ion chart!

~ next 2 pages

9. Write either the name or the formula for each substance. (This is NOT a matching exercise.)

They can be any type: type I ionic, type II ionic, type III covalent, or acids.

- | | |
|------------------------------|--|
| a. sodium sulfide | b. HCl |
| c. gold (III) fluoride | d. N_2O_5 |
| e. carbonic acid | f. $\text{Ca}(\text{NO}_3)_2$ |
| g. dihydrogen monoxide | h. H_3PO_3 |
| i. lead (IV) sulfate | j. $\text{Fe}_2(\text{C}_2\text{O}_4)_3$ |
| k. tetracarbon decahydride | l. $\text{Mg}(\text{CN})_2$ |
| m. nitric acid | n. KMnO_4 |
| o. chromium (III) dichromate | p. PBr_5 |

Chemical Reactions

- **Evidence a chemical reaction has occurred. One of these pieces of evidence must be present to understand that a chemical reaction has occurred.**
 - **A precipitate occurs (that means a solid forms)**
 - **Bubbles form**
 - **Water forms**
 - **There is a change in energy (the substance gets hotter or colder)**
- **Writing skeleton equations- change the reactants and products into their formulas. Separate reactants from products by a yield sign (\rightarrow), and don't forget the states of matter: solid (s), liquid (l), gas (g), and aqueous (aq).**
 - **Don't forget the 7 diatomic elements: H_2 , O_2 , N_2 , Cl_2 , Br_2 , I_2 , F_2**

FORMULAS

Type I - Binary Ionic Compds How do I know?

metal + nonmetal

Formula:

- 1) Write symbol & charge of metal.
- 2) Do same for nonmetal
- 3) criss-cross charges. Simplify if needed.

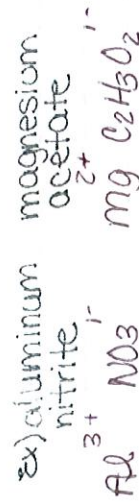


Type I - Tertiary Ionic Compds How do I know?

metal + polyatomic ion

Formula:

- 1) Write symbol & charge of metal
- 2) Do same for polyatomic ion
- 3) Criss-cross charges, use parentheses around polyatomic ion. Simplify.

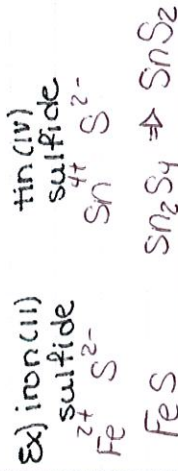


Type II - Binary Ionic Compds How do I know?

transition metal + nonmetal

Formula

- 1) Write symbol & charge of metal. (charge is Roman numeral)
- 2) Do same for nonmetal
- 3) Criss-cross charges. Simplify

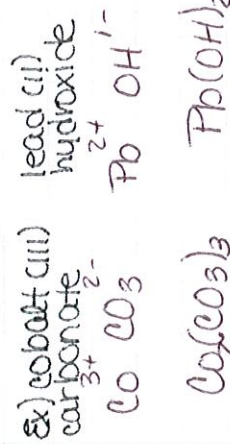


Type II - Tertiary Ionic Compds How do I know?

transition metal + polyatomic ion

Formula:

- 1) Write symbol & charge of metal (Roman numeral)
- 2) Do same for polyatomic ion.
- 3) Criss-cross charges use parentheses around polyatomic ion. Simplify

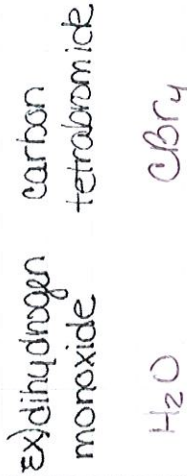


Type III - Binary Covalent Compds How do I know?

2 nonmetals

Formula:

- 1) Write symbol for 1st nonmetal, change prefix to subscript.
- 2) Do same for 2nd nonmetal.



Type III - Hydrocarbon Alkanes How do I know?

methane-
ethane-
propane-
butane-
pentane-

hexane-
heptane-
octane-
nonane-
decane

$C_x H_{2x+2}$

Alcohols

How do I know?

methanol-
ethanol-
propanol-
butanol-
pentanol-

hexanol-
heptanol-
octanol-
nonanol-
decanol-

$C_x H_{2x+1} OH$

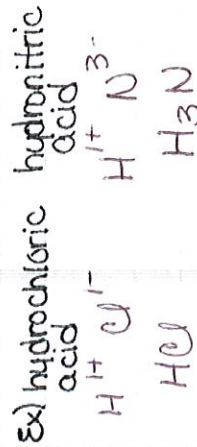
Binary Acids

How do I know?

hydro + element + ic acid

Formula:

- 1) Write H^{1+}
- 2) Write symbol & charge on nonmetal
- 3) Criss-cross charges



Oxyacid

How do I know?

poly- + ion + ic or poly- + ion + ous acid

Formula:

-ic acid

1) Write H^{+1}

2) Write symbol + charge of -ate ion.

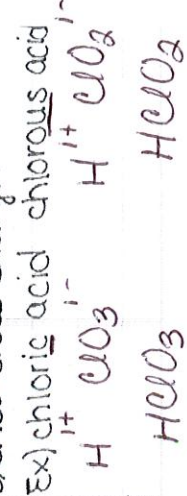
3) Criss-cross charges

-ous acid

1) Write H^{+1}

2) Write symbol + charge of -ite ion.

3) Criss-cross charges



NAMES

Type I - Binary Ionic compd
How do I know?
metal + nonmetal
Name:
1) Write name of metal
2) Write name of nonmetal, change end to -ide
Ex) Na_2S aluminum chloride
 AlCl_3

Type II - Binary Ionic compd
How do I know?
transition metal + nonmetal
Name:
1) Named mostly the same as Type I but you have to determine the original charge of the transition metal.
 $3+ \quad 3- \quad 3+ \quad 2-$
Ex) TiP Fe_2O_3
titanium(III) iron(III) phosphide oxide

Type III - Binary Covalent compds
How do I know?
Nonmetals
Name:
1) Write prefix, then name of 1st nonmetal
2) Do same for 2nd nonmetal, change ending to -ide
Ex) SO_2 PF_3
sulfur phosphorus dichloride trifluoride

Binary Acids
How do I know?
H is 1st + nonmetal
Name:
write hydro + element + ic acid
Ex) HCl hydrochloric acid
 H_3N hydronitric acid
 H_2S hydrosulfuric acid

Type I - Tertiary Ionic Compd
How do I know?
metal + polyatomic ion
Name:
1) Write name of metal
2) Write name of polyatomic ion
Ex) $\text{Ca}(\text{NO}_3)_2$ calcium nitrate

Type II - Tertiary Ionic compd
How do I know?
transition metal + polyatomic ion
Name:
1) Same as type I but you have to determine the original charge of the transition metal
 $4+ \quad -2$
Ex) $\text{Pb}(\text{CO}_3)_2$ lead(IV) carbonate
 $3+ \quad 1-$
 $\text{Cr}(\text{ClO}_2)_3$ chromium(III) chlorite

Type III - Hydrocarbons
Alkanes
How do you know? $\text{C}_x\text{H}_{2x+2}$
 C_1H_4 C_2H_6 C_3H_8 C_4H_{10} C_5H_{12}
Alcohols
How do you know? $\text{C}_x\text{H}_{2x+1}\text{OH}$
 $\text{C}_1\text{H}_5\text{OH}$ $\text{C}_2\text{H}_5\text{OH}$ $\text{C}_3\text{H}_7\text{OH}$ $\text{C}_4\text{H}_9\text{OH}$ $\text{C}_5\text{H}_{11}\text{OH}$
 $\text{C}_6\text{H}_{13}\text{OH}$ $\text{C}_7\text{H}_{15}\text{OH}$ $\text{C}_8\text{H}_{17}\text{OH}$ $\text{C}_9\text{H}_{19}\text{OH}$ $\text{C}_{10}\text{H}_{21}\text{OH}$

Oxyacids
How do I know?
H is 1st, then a polyatomic ion
Name:
1) Look at polyatomic ion...
end in ate, make -ic acid
end in -ite, make -ous acid
Ex) HClO_3 chloric acid
 HClO_2 chlorous acid
 H_2CO_3 carbonic acid
 H_2SO_3 sulfurous acid

Common Polyatomic Ions				
1- charge		2- charge		3- charge
Formula	Name	Formula	Name	Formula Name
H_2PO_4^-	Dihydrogen phosphate	HPO_4^{2-}	Hydrogen phosphate	PO_3^{3-} Phosphite
$\text{C}_2\text{H}_3\text{O}_2^-$	Acetate	$\text{C}_2\text{O}_4^{2-}$	Oxalate	PO_4^{3-} Phosphate
HSO_3^-	Hydrogen sulfite	SO_3^{2-}	Sulfite	
HSO_4^-	Hydrogen sulfate	SO_4^{2-}	Sulfate	
HCO_3^-	Hydrogen carbonate	CO_3^{2-}	Carbonate	
NO_2^-	Nitrite	CrO_4^{2-}	Chromate	1+ charge
NO_3^-	Nitrate	$\text{Cr}_2\text{O}_7^{2-}$	Dichromate	Formula Name
CN^-	Cyanide	SiO_3^{2-}	Silicate	NH_4^+ Ammonium
OH^-	Hydroxide			Hg_2^{2+} Mercury (I)
MnO_4^-	Permanganate	O_2^{2-}	Peroxide	
ClO^-	Hypochlorite			
ClO_2^-	Chlorite			
ClO_3^-	Chlorate			
ClO_4^-	Perchlorate			

- **Balancing chemical equations – using subscripts to make sure you have the same number of atoms of each element on both the reactants and products sides of the equation.**

10. Balance the equations below:

- $\text{N}_2 (\text{g}) + \text{H}_2 (\text{g}) \rightarrow \text{NH}_3 (\text{g})$
- $\text{AlBr}_3 (\text{aq}) + \text{K}_2\text{SO}_4 (\text{aq}) \rightarrow \text{Al}_2(\text{SO}_4)_3 (\text{aq}) + \text{KBr} (\text{aq})$
- $\text{C}_5\text{H}_{12} (\text{l}) + \text{O}_2 (\text{g}) \rightarrow \text{CO}_2 (\text{g}) + \text{H}_2\text{O} (\text{g})$
- $\text{Ag}_2\text{O} (\text{s}) \rightarrow \text{Ag} (\text{s}) + \text{O}_2 (\text{g})$
- $\text{Na} (\text{s}) + \text{Fe}(\text{NO}_3)_3 (\text{aq}) \rightarrow \text{Fe} (\text{s}) + \text{NaNO}_3 (\text{aq})$

- **5 Types of Reactions**

- o **Synthesis – 2 or more reactants combine to make 1 product**
- o **Decomposition – 1 reactant breaks down into 2 or more products**
- o **Single-replacement – an element replaces another element in a compound**
- o **Double-replacement – 2 compounds switch ions**
- o **Combustion – a hydrocarbon reacts with oxygen gas to produce carbon dioxide and water**

11. Identify the type of reaction for each equation in #10.

12. Predict the products for the following reactions:

- decomposition $\text{AlF}_3 \rightarrow$
- single – replacement $\text{Na} + \text{Ti}(\text{NO}_3)_3 \rightarrow$
- combustion $\text{C}_2\text{H}_6 + \text{O}_2 \rightarrow$
- double – replacement $\text{H}_2\text{SO}_4 + \text{KOH} \rightarrow$
- synthesis $\text{Ba} + \text{Cl}_2 \rightarrow$

Moles

- **Molar Mass – the sum of the atomic masses for each atom of each element in the compound**

- o **Example- $\text{Ca}(\text{NO}_2)_2$**

$$\text{Ca: } 1 \times 40.08\text{g} = 40.08\text{g}$$

$$\text{N: } 2 \times 14.01\text{g} = 28.02\text{g}$$

$$\text{O: } 4 \times 16.00\text{g} = \underline{64.00\text{g}}$$

$$132.10\text{g}$$

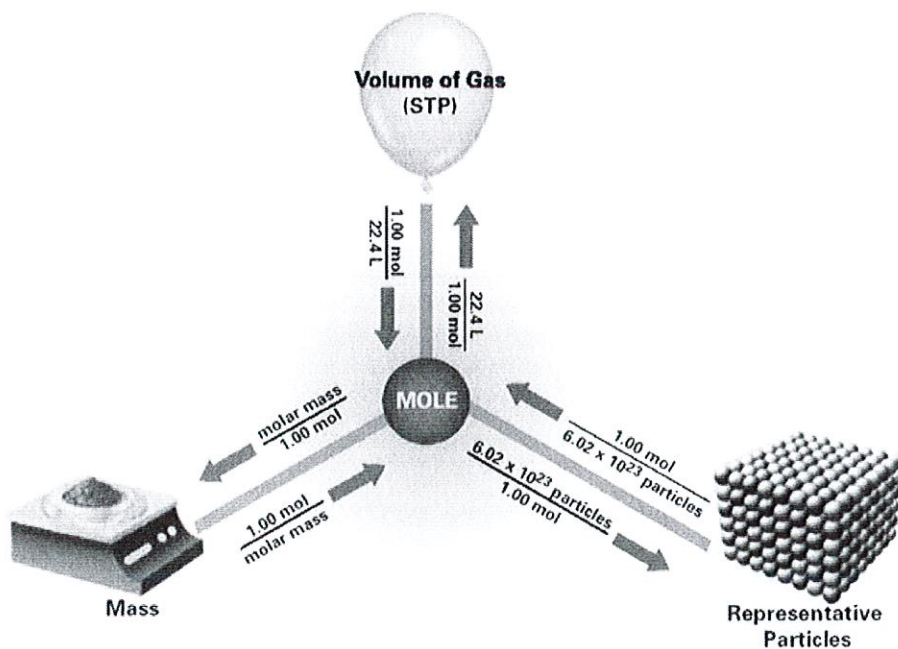
- **% Composition** – determine the % of each element by mass in a compound
 - o **% element** = $\frac{\text{mass of element}}{\text{molar mass}} \times 100$
 - o **Example** – Find the % of each element in $\text{Ca}(\text{NO}_2)_2$. (I will use the molar mass example from above for this.) You can check your answers by adding your final %'s up, they should be very close to 100%.
 - $\% \text{Ca} = \frac{40.08\text{g}}{132.10\text{g}} \times 100 = 30.34\%$
 - $\% \text{N} = \frac{28.02\text{g}}{132.10\text{g}} \times 100 = 21.21\%$
 - $\% \text{O} = \frac{64.00\text{g}}{132.10\text{g}} \times 100 = 48.45\%$

14. Find the molar mass and % composition for all the elements in the following compounds:

- CuBr_2
- $\text{CO}_2(\text{CO}_3)_3$
- $\text{Cr}_3(\text{PO}_4)_2$
- SF_6

- **Mole Calculations**

- o Use your mole map to guide you in solving mole problems. Remember to always figure out your given and unknown first! That way you know where to start and where to end up on your mole map.



- How many moles are in 34.5 g of water?
- How many molecules are in 0.456 liters of water as steam?
- How many moles are in 21.1g of calcium chloride?
- How many moles are in 6.789×10^{22} formula units of calcium chloride?

19. How many atoms are in 99.99g of nitric acid, HNO_3 ?

20. What is the volume, in liters, of 7.87×10^{25} molecules of sulfur trioxide?

– **Empirical and Molecular Formulas**

○ **What's the difference?**

- A molecular formula is the true formula for a substance, for example, tetraphosphorous decoxide is P_4O_{10} .
- An empirical formula has the coefficients reduced down to their lowest possible ratio, for example, that P_4O_{10} can be reduced to P_2O_5 .

21. Write the empirical formula for the following substances. If it already is an empirical formula, let it be!

a. C_6H_{14}

b. C_3H_8

c. $\text{C}_6\text{H}_{12}\text{O}_6$

d. $\text{Co}_2(\text{CO}_3)_3$

22. Determine the empirical formula for a coconut oil, a compound that is 66.611% carbon, 11.204% hydrogen, and 22.185% oxygen.

23. Determine the empirical formula for isoamyl acetate, the scent of bananas. It is composed of 64.56% carbon, 11.06% hydrogen, and 24.58% oxygen.

Stoichiometry – determining how much product can be made from a specific amount of reactants. It is based on the law of conservation of mass and uses **MOLE RATIOS** to convert between substances!

– **Mole ratio = $\frac{\text{moles of unknown}}{\text{moles of given}}$**

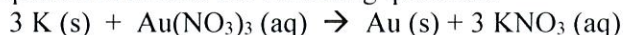
– If you want to see a step by step procedure for doing stoichiometry, please go to your notes or the following website-

<http://misterguch.brinkster.net/stoichiometryexplained.pdf>

– **Limiting and Excess reactants**

- Limiting reactants control how much product can be made during a chemical reaction because you run out of that reactant first!
- An excess reactant is one in which you have more than enough for the reaction to run to completion.

24. Use the balanced equation to answer the following questions:



- a. How many moles of potassium nitrate are produced from 21.0 moles of potassium?
- b. How many grams of gold are produced from the reaction of 1.50 moles of potassium with excess gold (III) nitrate?
- c. What is the mass, in grams of potassium nitrate produced from the reaction of 50.00g of gold (III) nitrate with excess potassium?
- d. Determine the limiting reactant if 30.00 moles of potassium reacts with 9.00 moles of gold (III) nitrate.

- e. Determine the limiting reactant if 1173g of potassium reacts with 3447g of gold (III) nitrate.

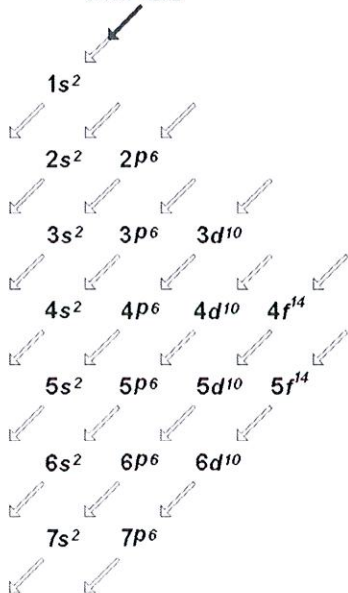
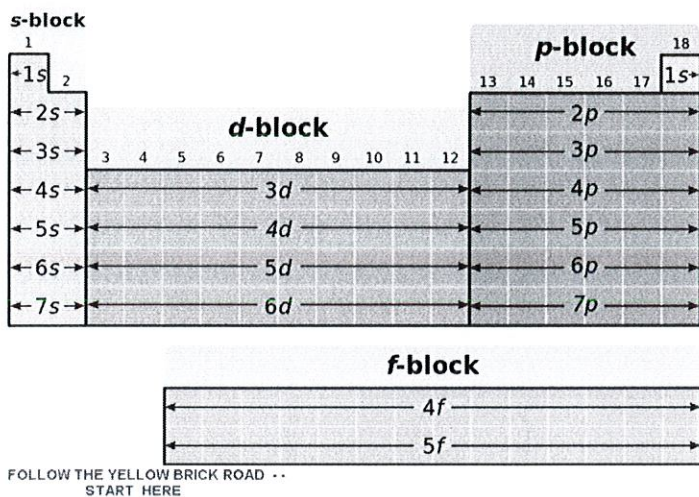
Atomic Structure

- **Subatomic Particles**
 - o **Inside the nucleus of an atom**
 - **Protons (p^+)**
 - Positively charged particles, charge = +1
 - # of protons in the nucleus identifies an element, is the atomic number
 - **Neutrons (n^0)**
 - Particles with no charge, charge = 0
 - They maintain the stability of the atom, in a nuclear/radioactive sense.
 - **Mass number = $\#p^+ + \#n^0$**
 - Is **ONLY** the same as a round atomic mass if you are describing the most common isotope of that element!
 - o Isotopes= are atoms of an element that have the same atomic number but different mass numbers (they have different #'s of neutrons!)
 - o **Outside the nucleus**
 - **Electrons (e^-)**
 - Negatively charged particles, charge = -1
 - Equals the # p^+ in a neutral atom (not an ion)
 - They are involved in bonding atoms to form compounds

25. Fill in the chart below. You may use your periodic table also.

Element	atomic #	mass #	# protons	# electrons	# neutrons
sulfur		32			
	48				64
		181	73		
				99	153

- **Electron Configurations** – tell you the placement of the electrons within the energy levels (orbitals) and sublevels within the atom
 - o The ground state is the lowest energy levels for the electrons and how we write electron configurations.
 - o Sublevels- each type of sublevel can hold a different amount of electrons
 - $s = 2 e^-$, $p = 6 e^-$, $d = 10 e^-$, $f = 14 e^-$

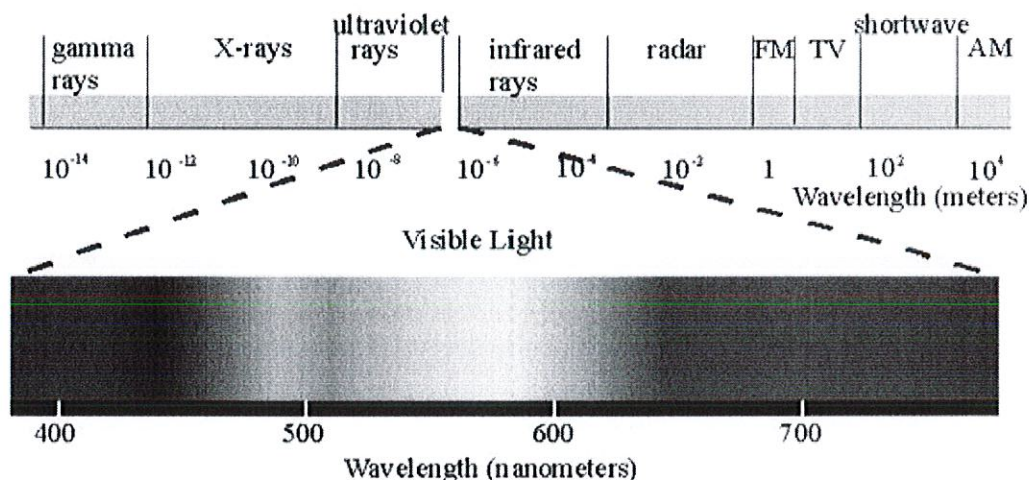


26. Write ground state electron configurations for the following elements:

- Ti
- Ni
- Br
- Sr

– Electrons and Light

- Electrons can give off light when they are excited. This means that the electron has absorbed energy and moved to a higher energy level where it is unstable. In order to become stable again and move back down to the ground state, the electron releases a photon of energy. This photon has a wavelength and a frequency in the visible light spectrum that we can see. We use this to identify elements (Remember the Flame Test Lab?) and to determine the age of stars. We also use this property to make colorful fireworks!
 - The lower the wavelength, the higher the frequency and the more energy the photon has!



$$c = \lambda \times \nu$$

$$c = 3.00 \times 10^8 \text{ m/s}$$

$$E = h \times \nu$$

$$h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s}$$

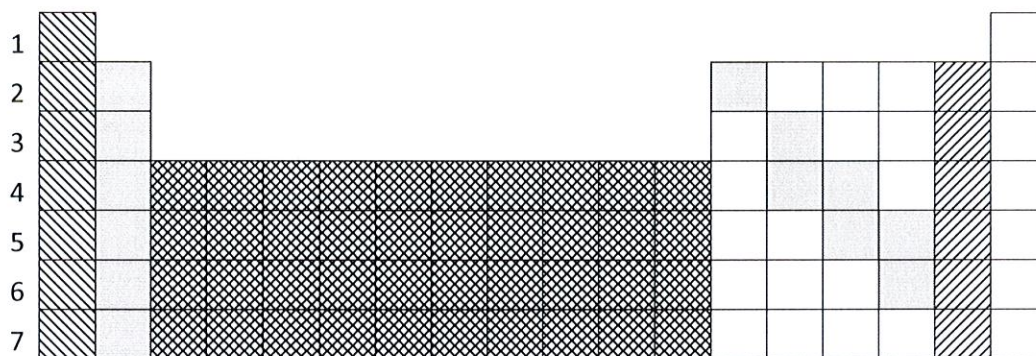
27. Which color of visible light has the highest frequency? Which color has the lowest frequency? How do you know?
28. Which color of visible light has the highest energy? Which color has the lowest frequency? How do you know?

Chemical Bonds

- **Ionic Bonds** – occur between a metal and a nonmetal
 - o An electron is transferred from the metal to the nonmetal
 - The metal becomes + charged and the nonmetal becomes – charged.
 - Shown by drawing electron dot diagrams
 - o Very strong bonds, takes a LOT of energy to break an ionic bond. Therefore, ionic compounds are hard crystals with very high melting and boiling points!
 - **Covalent Bonds**- occur between 2 nonmetals
 - o An electron is shared between the 2 nonmetals because neither one is strong enough (has a high enough electronegativity) to take the electron from the other element.
 - Shown by drawing Lewis structures
 - o Strong bonds but not as strong as an ionic bond so these compounds have lower melting and boiling points.
29. Determine if the following compounds have ionic or covalent bonds, or both ionic and covalent bonds. Draw Lewis structures for a and c.
- a. CO_2
 - b. FeF_2
 - c. NH_3
 - d. CaBr_2
 - e. MgSO_4
 - f. $\text{KC}_2\text{H}_3\text{O}_2$

Periodic Table and Periodic Trends

30. Label the periodic table legend below with the following:
Noble gases, alkali metals, transition metals, halogens, inner transition metals, alkaline earth metals, metalloids
31. What are groups on the periodic table and how many are there?
32. What are periods on the periodic table and how many are there?



legend



– Periodic Trends

- **When the periodic table is set up as it is, properties of the elements show reliable trends. So it is possible to predict how an element is going to react or act based on its location on the periodic table.**
- **In class, we discussed 3 properties: atomic radius, ionization energy (energy needed to remove an electron from the atom) and electronegativity (the attraction the nucleus of an atom has for another atom's electrons).**

- **Francium (Fr) vs. Fluorine (F)**

	Fr	F
atomic radius	largest	smallest
ionic radius	largest	smallest
ionization energy	smallest	highest
Electronegativity	smallest	highest

33. Based on the information above, rank these elements, **Sc, Br, and Ni**, in order of
- Increasing atomic radius
 - Decreasing ionic radius
 - Increasing ionization energy
 - Decreasing electronegativity
34. Based on the information above, rank these elements, **Se, Pb, and F**, in order of
- Increasing atomic radius
 - Decreasing ionic radius
 - Increasing ionization energy
 - Decreasing electronegativity

Kinetic Molecular Theory and States of Matter

- **Assumptions of the Kinetic Molecular Theory**
 - **Gases particles are very tiny and are relatively far apart from each other**
 - **Collisions between gas particles is elastic- no energy is lost or gained in the collision**
 - **Gas particles are in constant, rapid, random motion**
 - **There is no attraction between the gas particles, they ignore each other**
 - **The temperature of the gas is a measure of the average kinetic energy of the gas particles.**
- **Atmospheric pressure – the force of the atmosphere pressing down on the earth.**
- **Evaporation and Boiling**
 - **A liquid evaporates at the surface of a liquid when some of the liquid particles gain enough kinetic energy to break free from the liquid state and become a gas.**
 - **Boiling occurs when the pressure created by the gas particles that just escaped from the liquid equals the atmospheric pressure**

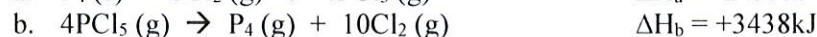
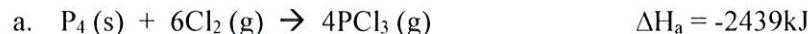
Heat

- **Flow of energy from a higher temperature to a lower temperature**
 - **Measured in joules (J) or calories (cal) or kilocalories (Cal)**
 - **Exothermic vs Endothermic Processes**
 - **Exothermic processes – occur when heat is released from the system into the surroundings.**
 - **Endothermic processes – occur when heat is added to the system from the surroundings.**
23. On a cold day outside watching a football game, you decide to get a huge hot chocolate to hold.
- If the hot chocolate is the system, is it endothermic or exothermic?
 - If your hands are the system, are they endothermic or exothermic?
24. In lab, you add some nitric acid to water and notice that the beaker gets warm, is this an endothermic or exothermic process?
25. In the summer, plants produce glucose (their energy source) by reacting carbon dioxide, water and heat and light from the sun. Is this process, endothermic or exothermic?

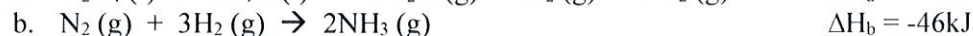
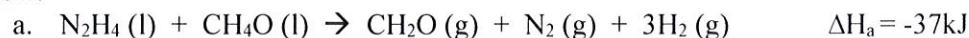
– **Hess's Law**

- **If you reverse a reaction, change the sign of ΔH**
- **If you multiply an equation, don't forget to multiply ΔH**

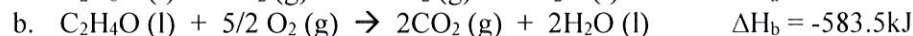
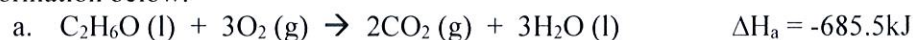
26. Calculate ΔH for the reaction, $\text{PCl}_5(\text{g}) \rightarrow \text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g})$ from the information below:



27. Calculate ΔH for the reaction, $\text{N}_2\text{H}_4(\text{l}) + \text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$ from the information below:



28. Calculate ΔH for the reaction, $2\text{C}_2\text{H}_4\text{O}(\text{l}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{C}_2\text{H}_6\text{O}(\text{l}) + \text{O}_2(\text{g})$ from the information below:



– **Calculating heat**

- **$q = mC\Delta T$**
 - **If q is negative, it's an exothermic process**
 - **If q is positive, it's an endothermic process**

29. A 15.75g piece of iron absorbs 1086.75 J of heat and its temperature changes from 25°C to 175°C. Calculate the specific heat capacity of iron.

30. How much heat is needed to raise the temperature of a 2300.g aluminum lawn chair from 2°C to 16°C? The specific heat capacity of aluminum is 0.90J/g°C.

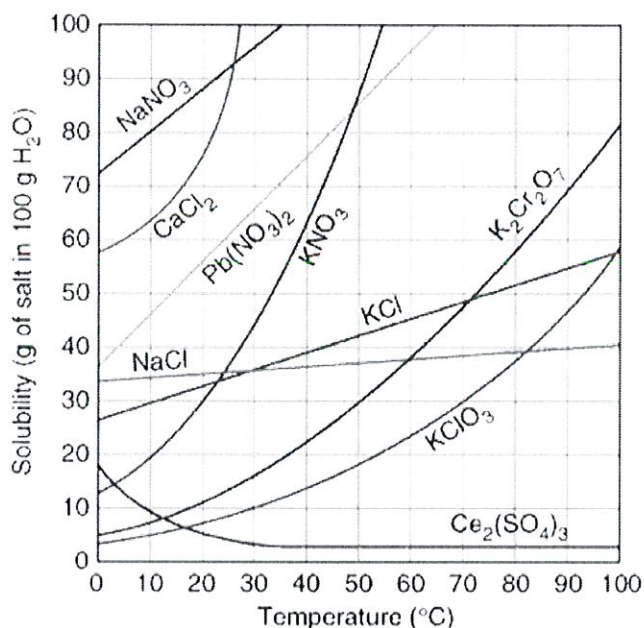
31. How much heat is released when 454g of zinc are cooled from 96.0°C to 28.0°C? The specific heat capacity of zinc is 0.386J/g°C.

Solutions

- **Factors that affect the rate at which a solute can dissolve in a solvent**

- 1. _____
- 2. _____
- 3. _____
- 4. Why do each of these factors affect how fast the solute dissolves? _____

○ **Solubility Graph**



32. How much potassium nitrate will dissolve in 100g of water at 70°C?
33. How much lead (II) nitrate will dissolve in 200g of water at 60°C?
34. If you dissolve 20g of potassium dichromate in 100g of water at 50°C, is the solution unsaturated, saturated, or super saturated?

○ **Molarity – a measure of concentration**

- $M = n/V$
 - n = moles of solute, V = volume of solution in L
- **Dilutions**
 - $M_1V_1 = M_2V_2$

35. Calculate the molarity if you dissolve 15.0g of Mg(OH)₂ in 400mL of water.
36. How many liters of water were used to create a 0.76M solution of HCl using 5.00g of HCl?
37. What mass of NH₃ was used to create a 1.23M solution that has a volume of 1.00L?
38. What is the new molarity of a 2.00L solution made by add water to 1.20L of a 0.856M solution of KCl?

- **Colligative Properties – properties of a solution that depend only on the number of particles in solution not what the solute is.**

- **Boiling Point Elevation** – adding a solute to a solution raises the boiling point of the solution because _____
- **Freezing Point Depression** – adding a solute to a solution lowers the freezing point of the solution because _____

39. If you increase the number of solute particles in a solution, what will happen...
- To its boiling point?
 - To its freezing point?

- **Acids and Bases**

40. According to Arrhenius, what is the definition of an acid? A base?
41. According to Bronsted- Lowry, what is an acid? A base? A conjugate acid? A conjugate base?
42. Write which of the substances is a B.L. Acid, B.L. Base, Conjugate Acid, and Conjugate Base:
- $\text{H}_2\text{O} + \text{H}_2\text{SO}_4 \rightarrow \text{H}_3\text{O}^+ + \text{HSO}_4^-$
 - $\text{H}_2\text{O} + \text{NH}_3 \rightarrow \text{OH}^- + \text{NH}_4^+$
43. What is the difference between a strong acid and a weak acid? Name 7 strong acids. Name 8 strong bases.
44. Are the following solutions acidic or basic?
- A solution of NaOH with a pH = 10.5
 - A solution of NaOH with a pH = 6.7
 - A solution of $\text{HC}_2\text{H}_3\text{O}_2$ with a pH = 4.6
 - A solution of $\text{HC}_2\text{H}_3\text{O}_2$ with a pH = 7.2
45. Titrations (Neutralization), $\mathbf{M_A V_A = M_B V_B}$
- For the titration between KOH and HCl, determine the volume of 0.75M KOH you would need to titrate 0.0250L of 0.85M HCl.
 - How much 1.23M HCl would you need to neutralize 45.0 mL of a 0.990M LiOH solution?