Unit 1: Summer Assignment & Matter and its Properties

Part One: Sig Figs, Scientific Notation, SI Units, & Measurement

1. *Significant Figures:* know the rules of significant figures for:
   a. Addition/subtraction: the answer can be no more significant than the weakest number
   b. Multiplication/Division: the answer should have the same number of significant digits as the number with the smallest number of sig figs.
2. *Precision vs. Accuracy*: Repeated measurements that are in agreement are precise. A measurement that matches the actual number is accurate.

This is precise but *not* accurate	
This is accurate but *not* precise.	
This is accurate and precise.	

3. *SI Unit Conversions with Factor Labeling:*
   a. Know how to convert the SI Units for length, mass, & volume based on the following prefixes:

S.I. PREFIXES		
Prefix	Symbol	Amount
Mega	M	1,000,000 (10^{6})
Kilo	K	1,000 (10^{3})
Hector	H	100 (10^{2})
Deca	Da	10 (10^{1})
Deci	D	0.1 (10^{-1})
Centi	C	0.01 (10^{-2})
Milli	M	0.001 (10^{-3})
Micro	U	0.000001 (10^{-6})
Nano	N	0.000000001 (10^{-9})

Part Two: Matter and Its Changes

Chemistry: It is the study of Matter and its Composition, Properties, and the Changes that it undergoes.

1. **Matter:** it is the stuff that occupies space, and has mass, and which can be changed by energy.

2. **Composition:** What is it matter made of?

A. PURE SUBSTANCE: contains only one kind of matter.

Element: the fundamental substances of chemistry.

COMPOUND: made up from the combination of 2 or more elements in a precise, well-defined ratio.

B. MIXTURES: is a physical blend of two or more substances, where the composition of a mixture may vary.

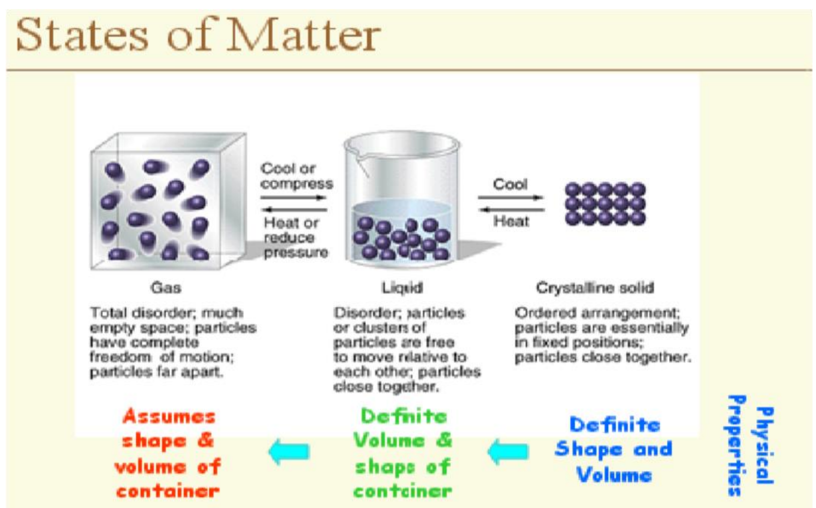
Mixtures can be divided into Two Types:

(1) **A Heterogeneous Mixture:** is one that is not uniform in composition. For example, a dinner salad, or oil in water consists of several components that are not evenly distributed. CAN BE PHYSICALLY SEPARATED into substances.

(2) **A Homogeneous Mixture:** is one that has completely uniform composition; its components are evenly distributed throughout the sample (i.e. sugar in water)
Homogeneous mixtures are VERY important in chemistry. In the liquid phase, these homogeneous mixtures are called SOLUTIONS.

3. **Suspension:** blood. Milk, orange juice. LOOKS CLOUDY, uniform. Particles will separate out (precipitate out) of mixture is left to settle. CAN BE PHYSICALLY SEPARATED into substances.

4. **Colloid:** gelatin, interior of a living cell. Particles remain in solution indefinitely. Looks like a **GEL**. CAN BE PHYSICALLY SEPARATED into substances.

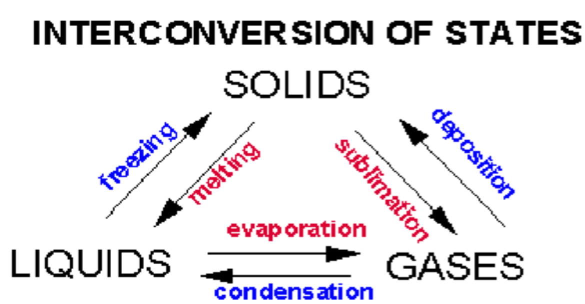


The States of Matter & The Kinetic Molecular Theory

Properties and Changes that Matter undergoes:

A. **Physical Change:** change in the physical appearance without change in the chemical composition. Changes of the states of matter, from solid, liquid, to gas (shown to the right) are PHYSICAL CHANGES.

B. **Physical Properties:** density, Boiling Point, Melting Point, Freezing Point, Conductivity. Physical properties are characteristics that can be observed without a chemical reaction



These physical properties are *INTENSIVE,*
which means that the property does not depend on the amount of material present. *Physical/Intensive Properties include Boiling Point, Melting Point, Freezing Point, Color, Odor, State, malleability, Ductility (metal can be drawn into a wire), Luster, Solubility, Viscosity, Luminosity, Hardness, Density, & Conductivity*

By contrast, *EXTENSIVE* Properties depend on the amount of material; i.e. mass, volume, energy.

C. **Chemical Properties:** describes how one substance reacts with another.
D. **Chemical Change:** the substance undergoes a change in chemical composition, in a Chemical Reaction.
A chemical reaction is given by a **Chemical Equation:**

A + B	→	C + D
Reactants	"Yields"	Products

Some example chemical reactions:

a. $2\ H_2O \rightarrow 2\ H_2 + O_2$
b. *Combustion* of a hydrocarbon: rapid combination of oxygen with other materials
i.e. $2\ C_8H_{18} + 25\ O_2 \rightarrow 16\ CO_2 + 18\ H_2O$

The Ten signs of a Chemical Reaction:

1. **Bubbles Appear**
2. **A Precipitate Forms**
3. **A color change occurs 4. The temperature changes.**
5. **Light is emitted.**
6. **A Change in Volume occurs.**
7. **A change in electrical conductivity occurs.**
8. **A change in melting point or boiling point occurs.**
9. **A change in smell or taste occurs.**
10. **A change in any distinctive chemical or physical property occurs.**

Thermal Change: the change that occurs when matter absorbs or releases heat.
Remember:
1. *Temperature* is a measure of the average kinetic energy of particles in matter.
2. *Heat* is the total energy of all of the particles in the sample; it is the form of energy that flows between two bodies when the bodies are at different temperatures; heat will flow from the hotter body to the cooler body.
   a. *Endothermic Change:* occurs when matter ABSORBS energy (in=endo)
   b. *Exothermic Change:* occurs when matter RELEASES energy (exit=exo)

Separation of Mixtures: *how can we separate mixtures into their components?*

a. Physical methods: use of a magnet to separate a mixture of sulfur and iron? Why can these be separated by a magnet?
   i. Separation based on **physical properties**; i.e. magnetic

ii. **Physical separation with tweezers**
iii. **Filtration** of a solid that does not dissolve in a liquid, but is suspended, or precipitated at the bottom of the flask.
iv. **Solubilities:** we can separate mixtures based on solubilities.

b. Distillation: separation of a liquid mixture based on the boiling points (physical, intensive property) of each component.
c. Chromatography: separation based on solubility, and competition between attraction to paper and eluting solvent (TLC)

Separation of Compounds into Elements: *separation of compounds back to elements is accomplished only via Chemical Change.*

Unit 2: Atomic History and Structure (includes nuclear chem.)

1. Describe the various models in the historical development of modern atomic theory:
   - **a.** Aristotle: Matter is made of Air, Fire, Earth, & Water.
   - **b.** Democritus: The first to say that matter is composed of atom, or "atomos."
   - **c.** Dalton: Had five basic principles in his model of the atom
   - **d.** Thomson: discovered the charge of the electron by deflecting the flow of electrons through his Cathode Ray Tube with magnetic and electrical fields, and theorized that the atom was a "plum pudding" of electrons and positive charges.
   - **e.** Rutherford: Used the Gold Foil experiment to prove the Plum Pudding Model. He shot Alpha particles (positively charged He nuclei) at Gold foil only to discover that the atom has a nucleus, filled with protons, and other significant mass. Since most of the alpha particles passed through, he concluded that the atom is mainly empty space. The small number of particles that bounced back proved that the nucleus had mass, and positively charged protons.
   - **f.** Bohr: Electrons move about the nucleus in "orbits" in Quantum Energy levels. The maximum number of electrons per energy level is $2n_2$, where n = the energy level number.
   - **g.** Quantum Mechanical Model: electrons spin around the nucleus in clouds of probabilities.

2. Distinguish among protons, neutrons, and electrons in terms of their relative masses, charges, and location with respect to the nucleus

3. Infer the number of protons, neutrons, and electrons using the atomic number and mass number of an element from the periodic table and symbol notation
   i.e. ${}_{9}^{19}F$ has 9 protons, 9 electrons (both determined by the Atomic Number), and 10 neutrons (Mass Number – Atomic Number)
4. Explain how isotopes of an element differ: Isotopes are atoms of the same element with the same atomic number but different Mass Numbers. Thus, *isotopes have the same # of protons, but different numbers of neutrons*.

5. Explain why the atomic masses of elements are not whole numbers: Atomic Masses are "Weighted Averages" of the naturally occurring isotopes.

6. BE ABLE to calculate the average atomic mass of an element from isotope data. Predict which of the naturally occurring isotopes is most abundant from average atomic mass.

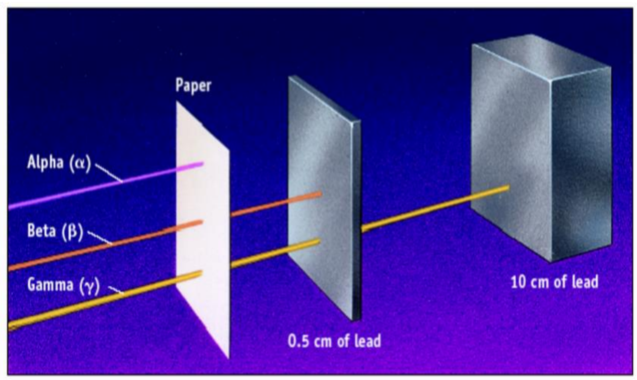
7. Differentiate between an atom and an ion. MAKE SURE that you understand that a "+" ion has LOST electrons, while a "-" ion has gained electrons … seems redundant … oh well?

8. Determine ion charge given proton, neutron, and electron data

9. The Periodic Table: What is the Periodic Law? **The Periodic Law states that when the elements are arranged in order of increasing atomic number, *there is a periodic repetition of their chemical and physical properties*.**
   a. The horizontal rows are called the periods. There are seven periods. Going across a period from left to right, elements are filling that energy level's "s & p" orbitals, eventually getting to a full octet at the noble gas.
   b. The vertical columns are called groups or families. Elements within the same group have the same number of valence electrons, and thus have similar properties.

10. Where are the metalloids? Where are the Transition Metals? The Lanthaniides? The Actinides? The Halogens? The Nobel Gases? The Alkali Metals? The Alkaline Earth Metals?

11. Contrast the characteristics of alpha, beta, and gamma radiation emitted during radioactive decay. Which is the strongest/most penetrating? The least?



12. Use symbol notation for subatomic particles and particle radiation

13. Write balanced nuclear equations for each of the following:

   a. alpha decay process
   b. beta decay processes

Alpha Decay

$$^{226}_{88}\text{Ra} \longrightarrow {}^{4}_{2}\alpha + {}^{222}_{86}\text{Rn}$$

radium-226 ⟶ α particle + radon-222

Mass number: (protons + neutrons)	226	=	4	+	222
Atomic number: (protons)	88	=	2	+	86

Beta Decay

$$^{7}_{4}\text{Be} + {}^{0}_{-1}\text{e} \longrightarrow {}^{7}_{3}\text{Li}$$

beryllium-7 + electron ⟶ lithium-7

Mass number: (protons + neutrons)	7	+	0	=	7
Atomic number: (protons)	4	+	−1	=	3

14. Compute the amount of radioisotope remaining at a given time using the half-life method.

15. What makes a Nuclei Unstable? Elements above an Atomic Number of 83 are all radioactive. The ratio of Neutrons to Protons are too high.

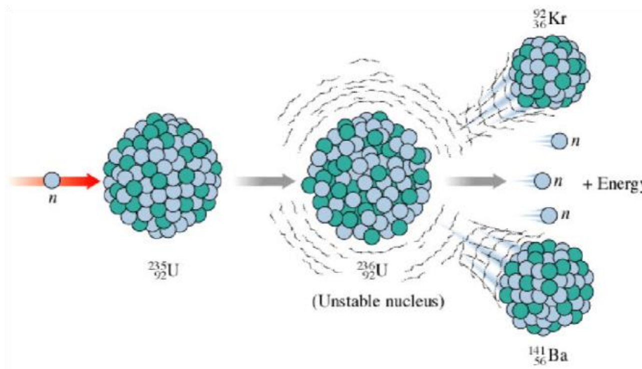
16. Distinguish between natural transmutations listed in #13a-b, and artificial transmutations of fission and fusion.

17. Write balanced nuclear equations for artificial transmutation

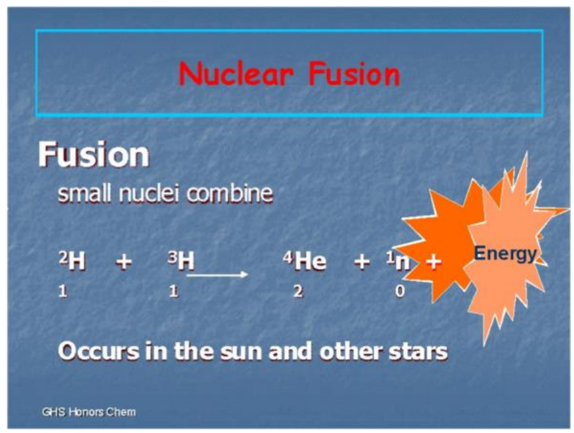
i.e. $^{9}_{4}Be + ^{4}_{2}He \rightarrow ^{12}_{6}C$ + ____

18. Compare and contrast nuclear fusion and nuclear fission

For **Fission**, a Critical Mass is needed to Start and maintain a Chain Reaction. The only fuels for Fission are Uranium-235 and Plutonium239.

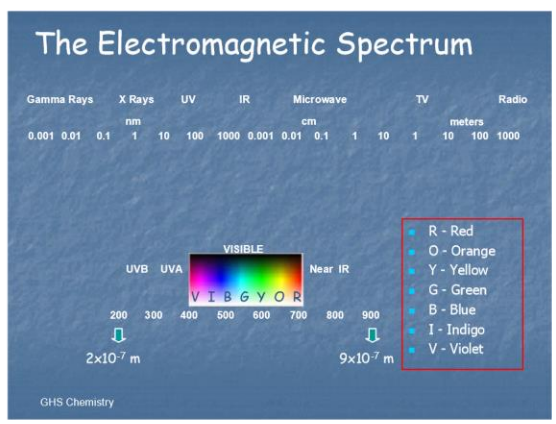


Fusion is the combination of small nuclei, and occurs in the Sun and other stars. Fusion is not practical as a source of energy since it occurs at EXTREMELY high temperatures.



Unit 3: Electrons and the Periodic Table (Chapters 4 & 5)

1. Explain the atomic emission spectrum of an element using Bohr's model of the atom. Electrons go from the ground state to the excited state when energy is absorbed. Then the electron moves from the excited state back to the ground state, *light is emitted*!



2. Calculate the Energy, frequency and wavelength of electromagnetic radiation (light) using the equations:

E = hυ, and c = λυ

Where h = Planck's constant (6.63×10^{-34} J-s)

&

c = 3×10^{8} m/s.

Unit 3: Periodic Table and Bonding

3. What are the parts of the Wave? Predict the relationships between wavelength, frequency, and energy for a wave.

4. Describe the properties of the different types of electromagnetic radiation.

5. Describe the Quantum mechanical Model of the atom.

6. Apply the Aufbau principle (filling of electron levels from low to high energy), the Pauli Exclusion Principle (two electrons occupying the same orbital must have opposite spins), and Hund's rule (1 electron must be given to each of equivalent suborbitals; ie. 3 p and 5 d suborbitals) to write electron configurations and orbital diagrams ($1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, …) of elements. Do not forget the exceptions to the Aufbau principle, specifically the d^4 and d^9 exceptions.

7. Identify the electron configuration for an **Excited Atom**, where an electron has moved from the Ground State to a higher energy level. For example, the electron configuration for the ground state of 20Ca is $1s^22s^22p^63s^23p^64s^2$. An possible electron configuration for an excited $_{20}$Ca atom may be $1s^22s^22p^63s^23p^54s^25s^1$

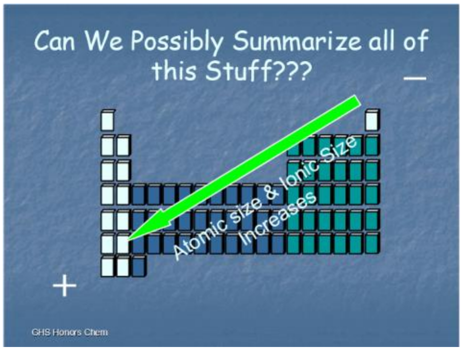
8. Complete "**Orbital Filling Diagrams**" for an element's electron configuration using the Aufbua, Pauli, and Hund's Rules.

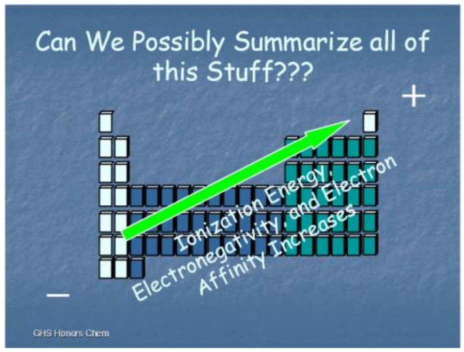
9. Describe the origins of the modern periodic table. Describe the organization of the periodic table (**periods, groups, periodic law, s block, p block, d block, & f block**), and categorize the elements as halogens, alkali metals, alkaline earth metals, noble gases, transition metals, inner transition metals, and representative elements (all of the Group A elements). *As a refresher, the Periodic law states that when elements are arranged by their Atomic Number, there is periodic repetition of properties within a group or family.*

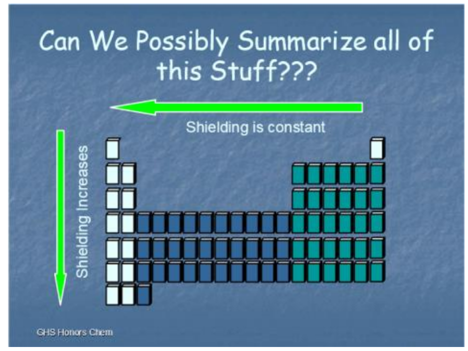
10. Contrast the physical and chemical properties of metals (LOSE ELECTRONS), nonmetals (GAIN ELECTRONS), and metalloids (DO BOTH), and locate them on the periodic table

11. Explain the relationship between the electron configuration of an element, its position on the periodic table, and its chemical properties. Simply put, be able to determine the electron configuration using the Periodic Table, and predict the charge of the ion based on the Group A number:
i.e. You would expect a Group 2A metal, which has 2 valence electrons, to lose both of those electrons and become a +2 ion.

12. State the trends of properties of elements within periods and groups of the Periodic Table. Interpret the trend shown by atomic radii, ionic radius, electronegativities, and 1st, 2nd, & 3rd ionization energies, within the periodic table.







13. Determine the number of valence electrons &/or predict the stable ion formed by a representative element, using the periodic table. Remember, the representative elements are the GROUP A elements.

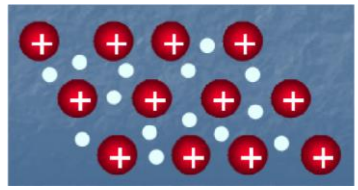
14. Predict Element location based on period and valence electrons. *For example, a valence configuration of $4s^2 4p^5$ tells me that the element is in the 4^{th} period, in Group 7A. The element is therefore Br (Atomic No. 35).*

1. Predict the formation of cations & anions from a representative atom, based on the location on the Periodic Table; more specifically, based on the atom's Group and the atom's desire to obtain an OCTET of electrons (or a noble-gas electron configuration).
   - Group 1A: loses s^1 to become a +1 cation
   - Group 2A: loses s^2 to become a +2 cation
   - Group 3A: loses s^2p^1 to become a +3 cation
   - Group 5A: gains 3 electrons to become s^2p^6, and a -3 anion
   - Group 6A: gains 2 electrons to become s^2p^6, and a -2 anion
   - Group 7A: gains 1 electron to become s^2p^6, and a -1 anion

2. Predict the formula unit of an ionic compound based on the expected charge of the ions, and write the Lewis Dot Structure for the Ionic compound.

3. **What are the characteristics/properties of an ionic bond?** Ionic compounds occur when very reactive metals transfer 1 or more electrons to reactive nonmetals. The resulting positive metal ion and nonmetal negative ions are highly attracted, and strongly bonded together. Ionic compounds exist as solids in a crystalline structure at room temperature, are dense, and have high & specific melting points. Ionic compounds are conductive only when melted or dissolved in water.

4. Explain the physical properties of metals using the theory of metallic bonding, where positively charged metal ions "float" within a sea of mobile valence electrons. Metals are conductive, malleable, and ductile due to this metallic bonding.



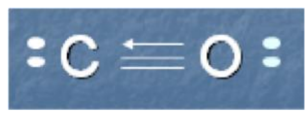
5. Describe the formation of a covalent bond between two nonmetallic elements. These bonds care either nonpolar covalent (for diatomic molecules; i.e. H_2, F_2, Cl_2), or polar covalent (all others that are not ionic, or nonpolar).

6. Identify bonds as ionic, polar covalent, nonpolar covalent using electronegativity values *__If asked to judge the polarity of a covalent bond based on the difference in electronegativity, use the following chart as a guide:__*

Electronegativity Difference	Type of Bond	Example
0 – 0.4	Nonpolar Covalent	H – H (0)
0.4 – 1.0	Moderately Polar Covalent	H – Cl (0.9)
1.0 – 2.0	Very Polar Covalent	H – F (1.9)
> 2.0	Ionic	Na^+Cl^- (2.1)

7. Describe double and triple covalent bonds and draw Lewis structures to represent covalent bond structures containing single, double, triple bonds, including exceptions to the octet rule. Those exceptions covered in class were:
   - Hydrogen (1 bond, 2 valence electrons only)
   - Boron (3 bonds, 6 valence electrons only)
   - Sulfur (can exceed the octet rule, and can have 6 bonding atoms, or 12 valence electrons)
   - Phosphorus (can exceed the octet rule, and can have 5 bonding atoms, or 10 valence electrons)

8. Explain the formation of a coordinate covalent bond (when one atom donates both electrons in a covalent bond).



9. What are the attractions between molecules? How do they relate to attractive forces between atoms? Describe the three Intermolecular forces covered in class from:

Weakest (Dispersion Forces) < Dipole-Dipole Interactions < Hydrogen Bonding < Covalent Bonds < Ionic Bonds (Strongest)

Chapter 7: Inorganic Naming & Formula Writing

1. Distinguish among atoms, a molecular formula for covalent molecules, and a formula unit for ionic compounds.
2. Distinguish between ionic compounds and covalent molecules.
3. Explain how a compound obeys the Law of Definite Proportions. *In a sample of a molecule or compound, the masses of the elements are always in the same proportions; i.e. water (H_2O) is always 2 Hydrogens to 1 Oxygen.*

4. Explain how two different compounds composed of the same elements obey the Law of Multiple Proportions. *The same atoms can combine in different ratios to form different molecules, with different properties; i.e. water (H_2O) and hydrogen peroxide (H_2O_2)*
5. Memorize the charges of common monoatomic ions.
6. Memorize those metals that have multiple charges, and use both the Stock Naming System (Roman numerals) and the Classical System for naming these metals in an ionic compound.

Those metals with Multiple charges;

The Transition Metals with Multiple Charges

Ion:	Name:	Ion:	Name:
Cr^{+2}	Chromium (II)	Ni^{+1}	Nickel (I)
Cr^{+3}	Chromium (III)	Ni^{+2}	Nickel (II)
Mn^{+2}	Manganese (II)	Cu^{+1}	Copper (I)
Mn^{+3}	Manganese (III)	Cu^{+2}	Copper (II)
Fe^{+2}	Iron (II)	Sn^{+2}	Tin (II)
Fe^{+3}	Iron (III)	Sn^{+4}	Tin (IV)
Co^{+2}	Cobalt (II)	Pb^{+2}	Lead (II)
Co^{+3}	Cobalt (III)	Pb^{+4}	Lead (IV)
		Hg_2^{+2}	Mercury (I)
		Hg^{+2}	Mercury (II)

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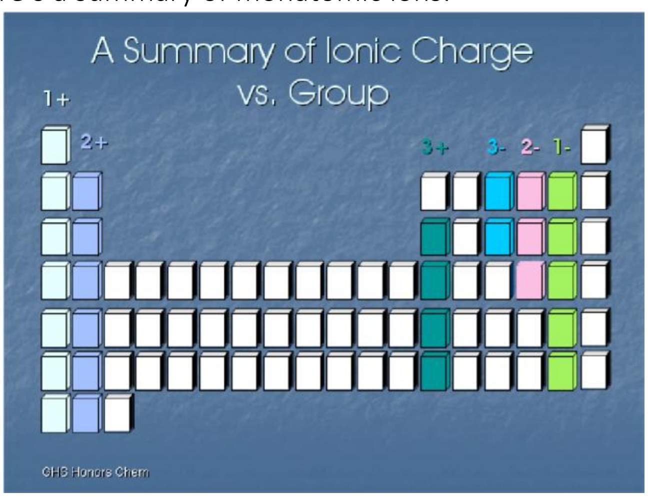
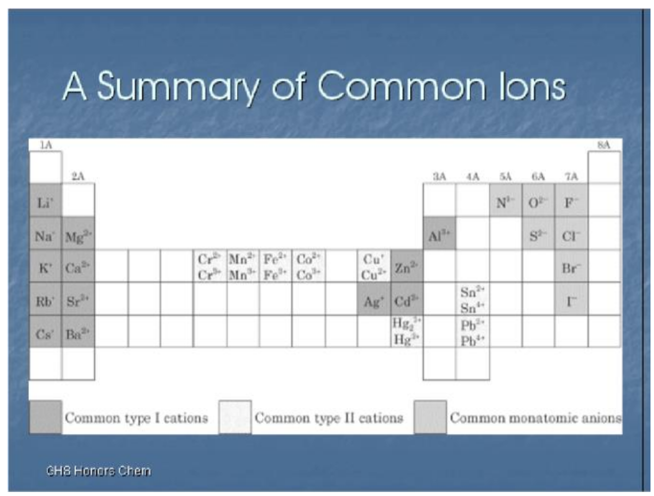
Here are the Classical names that you are responsible for:

Naming Ionic Compounds

	Stock System	Classical System
Cu^{+1}	Copper (I)	Cuprous
Cu^{+2}	Copper (II)	Cupric
Fe^{+2}	Iron (II)	Ferrous
Fe^{+3}	Iron (III)	Ferric
Sn^{+2}	Tin (II)	Stannous
Sn^{+4}	Tin (IV)	Stannic

Here's a summary of monatomic ions:





7. **Distinguish between an ion and a** polyatomic ion.

8. Memorize the names, formulas and charges of the common polyatomic ions.
9. Write chemical formulas for binary ionic compounds (only 2 elements present).

Common Polyatomic Ions

+1 Polyatomics	
NH_4^+	Ammonium ion
H_3O^+	Hydronium ion
- 1 Polyatomics	
NO_3^-	Nitrate ion
NO_2^-	Nitrite ion
ClO_3^-	Chlorate ion
ClO_2^-	Chlorite ion
$C_2H_3O_2^-$	Acetate ion
OH^-	Hydroxide ion
MnO_4^-	Permanganate ion
HCO_3^-	Bicarbonate ion
CN^-	Cyanide ion

- 2 Polyatomics	
SO_4^{2-}	Sulfate ion
SO_3^{2-}	Sulfite ion
CrO_4^{2-}	Chromate ion
$Cr_2O_7^{2-}$	dichromate ion
$C_2O_4^{2-}$	Oxalate ion
CO_3^{2-}	Carbonate ion
- 3 Polyatomics	
PO_4^{3-}	Phosphate ion
PO_3^{3-}	Phosphite ion

- ite for lesser Oxygen Atoms
- ate for more Oxygen Atoms

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10. Name binary ionic compounds when given the chemical formula.
11. Identify by name and write the chemical formulas for ternary ionic compounds (with polyatomic ions).
12. Identify the name of a covalent molecule from the formula. Write the chemical formula for a covalent molecule given the name.
13. Identify by name and write formulas for common acids.

For Simple Binary Acids:

If the anion attached to hydrogen ends in -ide, put the prefix hydro- and change -ide to -ic acid
HCl - hydrogen ion and chloride ion becomes hydrochloric acid

For Oxyacids: If the anion has oxygen in it, the polyatomic ion's name ends in -ate or -ite

change the suffix -ate to -ic acid (use no prefix)
H_3PO_4 Hydrogen and phosphate ions becomes Phosphoric acid
change the suffix -ite to -ous acid
H_3PO_3 Hydrogen and phosphite ions becomes Phosphorous acid