Accuracy & Precision

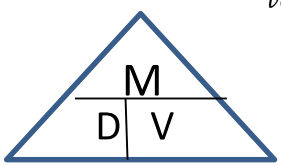
- Accuracy refers to how close a measured value is to its accepted value. (Hitting the bulls eye)
- Precision refers to how close a number of measurements are. (Always hitting the same person)

Percent Error

- Percent Error= $\frac{|experimental-actual|}{actual}$X actual

Density

- Density is the ratio that compares mass and volume on an object. *Density*= $\frac{mass}{volume}$ or remember



SI Prefixes

- 1 mL= 1 cm^3 and 1 L= 1dm^3
- Giga G 1,000,000,000 10^9
  Mega M 1,000, 000 10^6
  Kilo k 1,000 10^3
  Hecto h 100 10^2
  Deka da 10 10^1
  Deci d 1/10 10^{-1}
  Centi c 1/100 10^{-2}
  Milli m 1/1,000 10^{-3}
  Micro µ 1/1,000,000 10^{-6}
  Nano n 1/1,000,000,000 10^{-9}
  Pico p 1/1,000,000,000,000 10^{-12}

Chemical and Physical Properties and Changes

- Physical properties of matter can be observed or measured without changing the matter's composition.
- Physical properties include; density, color, odor, taste, hardness, melting point, etc.
- Chemical properties tell how a substance reacts with other substances.
- Extensive properties depend on the amount of the substances you are looking at.
  - EX. The *length* of a piece of lead depends on how much you are given so *length* is extensive.
- Intensive properties do not depend on the amount you are given.
  - EX. The *color* of silver will always be the same no matter how much you are given so *color* is intensive.
- Physical change- alters a substance without changing its composition.
  - EX. Crumbling up paper
- Chemical Change- relates to how a substance combines with or changes into one or more substances. (Reacts= chemical change)
  - EX. Na and Cl combining to form table salt

Significant Figures, Conversions, and Scientific Notation

- Sig Figs: Numbers from 1-9 always count, zeroes between nonzero numbers always count, zeroes to the right of a decimal place after nonzero numbers count, zeroes to the right of nonzero numbers without a decimal do not count, zeroes used for spacing the decimal do not count and one should not start counting sig figs until the first nonzero.
  - EX. 87600 3 Sig Figs
    87600. 5 Sig Figs
    87600.0 6 Sig Figs
    0.000675 3 Sig Figs
    1.000298 7 Sig Figs
- When adding and subtracting Sig Figs the answer can only have the same number of decimal places as the number with the least decimal places.
  - EX. 2.4|5 (2 decimal places)
    8.1| (1 decimal place)
    +3.7|2 (2 decimal places)
    14.2|7 but the final answer can only have 1 decimal place so the final answer is **14.3** (round up)
- When multiplying and dividing the answer has to have the same number of sig figs as the number with the smallest number of sig figs in the original problem.
  - EX. 2.56 x 6.2= 15.872 but it can only have two sig figs since 6.2 only has two so the final answer is **16** (round up)
- In scientific notation there is only one number to the left of the decimal.
- When writing scientific notation you increase by a power of ten every time you move the decimal to the left and you decrease by a power of ten every time you more the decimal to the right.
  - EX. 12,380 1.238 x 10^4 (you moved the decimal to the left 4 times)
    0.0097 9.7 x 10^{-3} (you moved the decimal to the right 3 times)
- When adding and subtracting scientific notation convert numbers to the same power of 10, then use the rule for adding and subtracting Sig Figs, last put the number back into correct scientific notation form.
  - EX. (3.95 x 10^5) + (7.8 x 10^3)
    1) Move the decimal in 7.8 two places to the left so both numbers are 10^5 = 0.078
    2) Add using Sig Fig rule
    3) Put back in correct scientific notation
       4.0 x 10^5
- When multiplying scientific notation multiply coefficients, add exponents, and finally put back in scientific notation with the same number of sig figs as the lowest factor.
- When dividing scientific notation divide coefficients, subtract exponents, and put back into correct scientific notation with the same number of sig figs as the lowest factor.

- $(6.8 \times 10^{3}) \times (4.54 \times 10^{6})$
  1) Multiply 6.8 and 4.54 =30.872
  2) Add exponents 3+6=9
  3) Put back into form 30.872×10^{9} but must move decimal over one so 3.0872×10^{10} now put into correct sig figs $\mathbf{3.1 \times 10^{10}}$
- $(9.2 \times 10^{-3}) / (6.31 \times 10^{6})$
  1) Divide 9.2 by 6.31= 1.458
  2) Subtract exponents -3-6= -9
  3) Put back into form 1.458×10^{-9} put into correct sig figs $\mathbf{1.5 \times 10^{-9}}$

- Conversion or dimensional analysis

Quantity (with unit given)	Conversion Factor
	Unit given

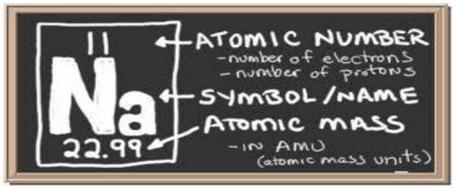
- EX. 178 mm to m

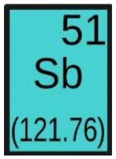
178mm	1m
	1000 mm

.178 m

Protons, Neutrons, electrons

- The center of the atom, or nucleus, is where the protons (positive charge) and neutrons (no charge) are located. Together they make up most of the atom's weight.
- The electrons (negatively charged) are located outside the atom in the electron cloud or orbitals. They have very little weight.
- Atomic number = the number of protons and if the atom is neutral or has no charge it also equals the number of electrons.

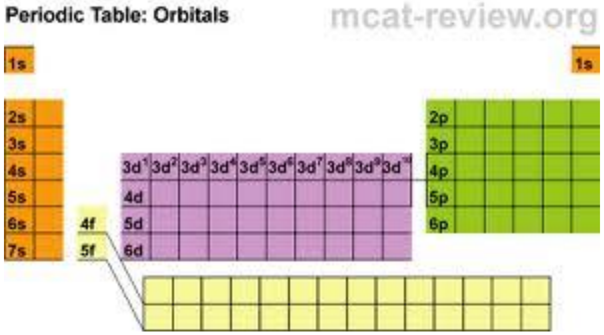


- Mass number – atomic number= number of neutrons
  - EX. 

    Protons= **51**
    Electrons= **51**
    Neutrons= 122-51= **71**

S,P,D,F Sublevels

- S orbitals are spherical while P orbitals are dumbbell



- You use sublevels to write electron configuration

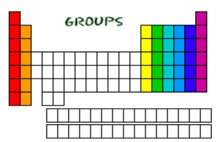
- EX. He $1s^2$
  Ti $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^2$

Quantum Numbers (n, l, m, d)

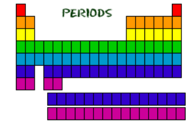
- Quantum numbers (n) indicates the relative sizes and energies of atomic orbitals. As n increases, the orbital becomes larger and the atom's energy level increases.
- n= energy level (row)
  l= sublevel (block)
  m= orbital (position of electron)
  d= direction in which electron spins

Periodic Table

- The periodic table is organized by increasing atomic number.
- Metals are on the left side or the table. Nonmetals are on the right side. Metalloids are on the "stairs" between metals and nonmetals.
- Groups (or) families are the 18 vertical columns found on the periodic table. All elements in a group/family have the same number of valence electrons.



- Valence electrons are the electrons in the outermost orbital of an atom.
- Periods are the 7 horizontal rows on the periodic table. All elements in a period have the same number of orbitals.



- Alkali metals- metals in group 1, most reactive metals (want to combine), have 1 electron in the s orbital, +1 charge
- Alkaline-earth metals- group 2, second most reactive metals, 2 electron in s orbital, +2 charge
- Transition elements- groups 3-12, charges vary
- Halogens- groups 17, most reactive nonmetals because they only need one election to stabilize, -1
- Noble gages- group 18, all of the elements are gases, all nonmetals, least reactive because they have a full outer shell of electrons
- MERIS
  M= metallic reactivity
  E= electron negativity
  R= radius size of atoms
  I= ionization energy
  S= Shielding effect

M E R I S

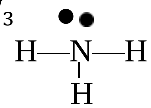
Electron Dot Diagrams

- Lewis dot structure- Write the element's symbol in the middle and put a dot for every electron on each side of the symbol starting from right, top, left, bottom

- EX. Sodium Na •

  Sulfur 

- When writing molecules anytime two dots of different elements are together a line forms.
  - EX. NH_3

- Remember all atoms want to be stable and have a full outer shell of 8 electrons.
- All elements can have 4 pairs of electrons when combining except H (1 pair), Be (2 pairs), B and Al (3 pairs)

Homogeneous and Heterogeneous materials, solutions, mixtures

- A mixture is a combination of two or more substances where each pure substance retains it individual chemical properties. Substances are not chemically bond.
- Mixtures can be heterogeneous or homogenous. In a *heterogeneous* mixture substances do not blend throughout and each substance remains distinct. In a *homogenous* mixture the composition is constant and the substances are the same throughout.
  - EX. Salt water- Homogenous mixture
    Raisin Bran- heterogeneous
- Alloy- homogenous mixture of metals. (Bronze, brass, pewter)
- Suspension- settles over time (muddy water, liquid medicine)
- Solution- a homogenous mixture in where the solute is dissolved in a solvent.
- Ways to separate mixtures:
  Sorting- by texture, color, shape
  Filter- particle size is different
  Magnet- one substance contain iron
  Chromatography- the substances move or are drawn across the surface or another material
  Density- one substances sink while some float
  Distillation- different boiling points

Elements/Compounds, Bonds, and Naming

- Elements are only one type of atom (Fe, N_2)
- Compounds contain a combination of elements with different proportions then the individual elements themselves. (NaCl, salt water, HNO_3)
- Ionic bond- between a metal and nonmetal, electrons are transferred
- Covalent bond- two nonmetals, also called a molecule, atoms share electrons

Greek Prefixes Used in Naming Molecular Compounds	
Prefix	**Meaning**
Mono-	1
Di-	2
Tri-	3
Tetra-	4
Penta-	5
Hexa-	6
Hepta-	7
Octa-	8
Nona-	9
Deca-	10

- When naming a covalent bond use prefixes. Do not use mono on first element. End the last element in ide.
    - EX. SbCl_3 Antimony Trichloride

  P_2O_5 Diphosphorus pentaoxide

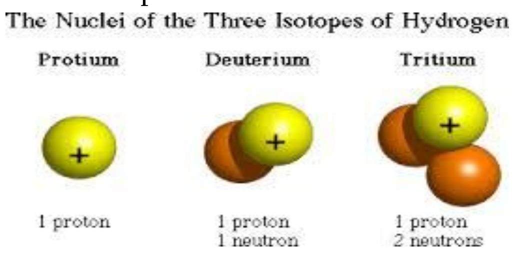
- When naming a ionic bond and the charge is known you name the first element then the second element ending in ide.
    - EX. $BaCl_2$ Barium Chloride
- When naming a metal and a polyatomic and charge is known you name the metal and then the polyatomic. (No changes)
    - EX. $LiClO_4$ Lithium perchlorate
- When naming an ionic bond and the charge is unknown you must reverse criss cross or solve for the oxidation number (charge). Then use the stock system where you write the roman numeral after the element to tell what it's charge is.
    - EX. CuCl write the oxidation number $Cu^{+1}Cl^{-1}$, criss cross, write the name with the stock system **Copper (I) Chloride**
      Pb$(C_2H_3O_2)_2$ Lead (II) acetate
- For hydrates use ionic naming for first part then use prefixes for hydrate
    - EX. $CaSO_4 \cdot 2H_2O$ Calcium sulfate dehydrate
- For acids if it's binary (H+ element) hydro____ic acid
    - EX. HI hydroiodic acid
- If it's acid + polyatomic (oxyacid) then change endings ate→ic ite→ous and add the word acid
    - EX. HNO_2 nitrous acid
      HNO_3 nitric acid

Oxidation Numbers

- All free, uncombined elements have a number of zero (F_2 = 0)
- Hydrogen is +1 except metal hydrides number is -1
- Oxygen has a charge of -2 unless it's peroxide where it's -1, or OF_2 is +2
- Alkali metals are +1
- Alkaline earth metals are +2
- Aluminum is always +3
- Halogens are always -1 unless bonding with oxygen or another halogen
    - EX. Na_2O since O is -2, the two Na must be -1

Isotopes

- In an isotope the number of neutrons vary.



Diatomic Atoms

- H, N, F, O, I, Cl, Br

Trends

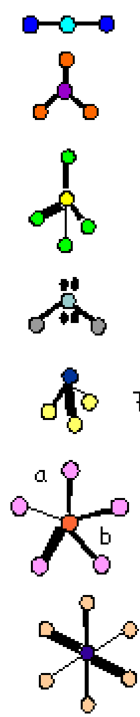
- Atomic radii- distance between the nuclei of atoms that are bond together.
- Ionization energy- the energy required to move one electron from a neutral atom of an element.
- Electronegativity- the ability for an atom to attract electrons
- Electron affinity- the change of energy when an electron is added to a neutral ion

Ionic, polar, and nonpolar bonds

- If ≥ 2.0 it's ionic
  If < 2.0 it's polar covalent
  If ≥ 0.5 it's polar covalent
  If < 0.5 it's nonpolar

- A polar molecule has a partial positive and partial negative side
- Polar covalent bonds form when not all atoms that share electrons attract them equally.

Molecule Shape, Bond angles, and VSEPR Theory



Shape.	Bond Angle.	Examples.
Linear	180	$HgCl_2$, CO_2
Trigonal Planar	120	BF_3, C_2H_4, CO_3^{2-}, NO_3^-
Tetrahedral	109.5	SiF_4, NH_4^+, PO_4^{3-}, CH_4
~~Tetrahedral~~ (bent)	104.5	H_2S
~~Tetrahedral~~ (pyramidal)	106.7	PF_3, H_3O
Trigonal Bipyramidal	a 120 b 90	PCl_5
Octahedral.	90	$[PF_6]^-$

- Water H_2O is always 104.5°

- VSEPR- Valence Shell Electron Pair Repulsion Model, an arrangement that minimizes the repulsion of shared and unshared pairs of electrons around a central atom.

Polyatomic ions and Ions

- Anions are ions with a negative charge
- Cations have a positive charge
- Polyatomic ions are a group of charged atoms that are covalently bonded (all nonmetals)

Table 8.4

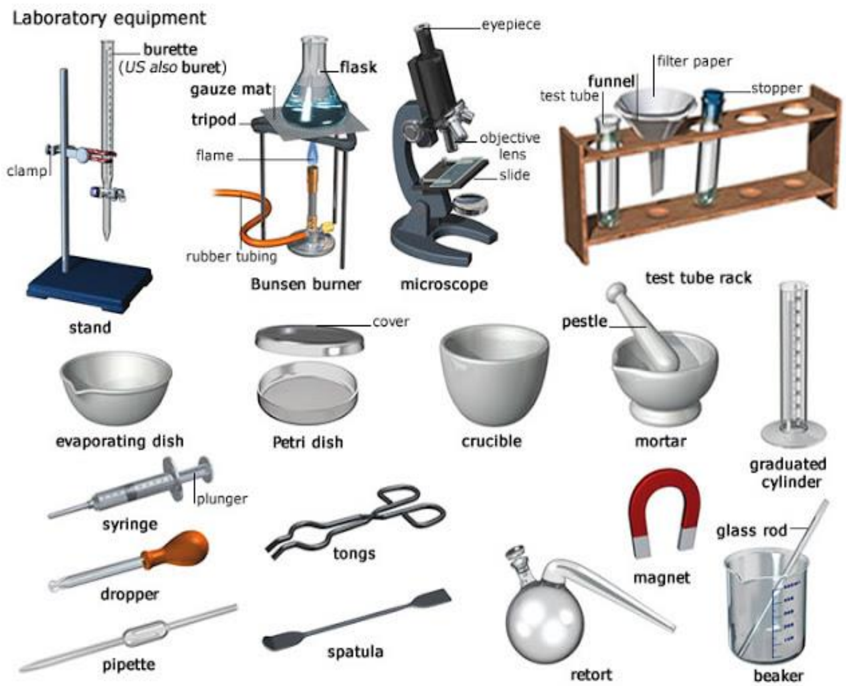
Common Polyatomic Ions					
1– charge		2– charge		3– charge	
Formula	Name	Formula	Name	Formula	Name
$H_2PO_4^-$	Dihydrogen phosphate	HPO_4^{2-}	Hydrogen phosphate	PO_3^{3-}	Phosphite
$C_2H_3O_2^-$	Acetate	$C_2O_4^{2-}$	Oxalate	PO_4^{3-}	Phosphate
HSO_3^-	Hydrogen sulfite	SO_3^{2-}	Sulfite		
HSO_4^-	Hydrogen sulfate	SO_4^{2-}	Sulfate		
HCO_3^-	Hydrogen carbonate	CO_3^{2-}	Carbonate	1+ charge	
NO_2^-	Nitrite	CrO_4^{2-}	Chromate	Formula	Name
NO_3^-	Nitrate	$Cr_2O_7^{2-}$	Dichromate	NH_4^+	Ammonium
CN^-	Cyanide	SiO_3^{2-}	Silicate		
OH^-	Hydroxide				
MnO_4^-	Permanganate				
ClO^-	Hypochlorite				
ClO_2^-	Chlorite				
ClO_3^-	Chlorate				
ClO_4^-	Perchlorate				

Chemical Names and Formulas **147**

Quantum and Light

- Quantum mechanics describes the behavior of subatomic particles traveling at velocities near that of light.
- Wavelength (λ) is the distance between equivalent points on a wave. (Crest to Crest)
- Frequency (√) number of waves that pass a given point per second
- Speed of light (c) is 3.00 x 10^8
- C=λ√
  - EX. Wavelength is 633nm, what is the frequency? $\surd = \frac{3.00\times10^8}{633nm} = \frac{3.00\times10^8}{633\times10^{-9}} = 4.74 \times 10^{-4}$ Hz
- When energy is absorbed and an electron enters a higher energy level it is in a *excited state*
- The lowest allowable energy state of an atom is its *grounded state.*
- The light emitted by an electron in its excited state is called a *photon*

Lab Equipment



Endo and Exothermic

- Endothermic is when a reaction absorbs energy in the form of heat
- Exothermic is when a reaction gives off energy in the form of heat.

Scientists

- Thomson- discovered electrons
- Rutherford- gold foil experiment that discovered the atom was mostly empty space, positive particles called protons were in the center or nucleus, and that electrons orbit the nucleus,
- Millikan- determined that the electron's charge as -1 and found the mass on an electron
- Crookes- cathode ray experiment that discovered negatively charges particles
- Plank- came up with plank's constant and a relationship between frequency and energy
- Einstein- electromagnetic radiation has wavelike and particle like natures
- Heisenberg- it is impossible to know the exact velocity and position of a particle at the same time.
- Schrodinger- came up with an equation that treats electrons as waves
- Dalton- law of multiple proportions (If two elements form more than one compound between them, then the ratios of the masses of the second element which combine with a fixed mass of the first element will be ratios of small whole numbers), also developed the first atomic theory
- Democritus- first person to say matter was not infinitely divisible
- Bohr- came up with idea of grounded and excited states
- Pauli- electron in the same orbital have different spins
- De Broglie- all moving particles have characteristics
- Aufbau – electrons will fill the lowest energy level first

- Hund- each orbital in an energy level must have an electron before it starts to pair
- Mendeleev- created the first periodic table, was organized by mass of the element
- Chadwick- discovered neutron and that they have no charge and are needed to stabilize the nucleus.
- Mosley-discovered each atom has a unique positive charge in the nucleus and created the modern periodic table based off atomic number
- Becquerel- discovered radioactivity
- Lavoisier- law of conservation of mass (says that the total mass of the substance in a chemical reaction cannot change)
- Proust- law of definite proportions (every compound has a fixed proportion)

Average Atomic Mass

- Calculated by the percent abundance times the mass of that isotope added to the percent abundance times the mass of that isotope…
    - EX. Mg has three isotopes: 78.7% Mg 24, 10.1% Mg 25, and 11.2 % Mg 26

      (.787 X 24)+ (.101 X 25)+ (.112 X 26)= 24.426 but must put in corrects sig figs so the final answer is 24.4