Atomic Concepts

1. Dalton – basic unit of matter is the atom; cannon ball or hard sphere model
2. JJ Thomson – used cathode ray tube; found electrons; "plum pudding model"
3. Rutherford – gold foil experiment; small, dense, positively charged nucleus; mostly empty space
4. Bohr – nucleus surrounded by electrons in circular orbits; planetary model or electron shell model
5. Atoms have a positive charged nucleus, surrounded by negatively charged electrons
6. Protons and neutrons are found in the nucleus
7. Protons are positive, neutrons have no charge, electrons are negative (table O)
8. In an atom, # protons = # electrons (no charge = neutral atom)
9. Mass of a proton and neutron are 1 amu; electron < 1 amu
10. Orbitals – regions of the most probable electron location
11. Each electron in an atom has its own specific amount of energy
12. The first principle energy level has the least energy. Energy increases as you go from PEL 1-7.
13. Electrons have a higher energy state in the excited state
14. Energy is released when electrons fall back (excited → ground)
15. Bright light spectrum is produced when electrons return to ground state
16. Valence electrons – outermost electrons
17. Valence electrons affect chemical properties of an element
18. Isotopes – same number of protons, different number of neutrons
19. Atomic mass – average of all naturally occurring isotopes
20. # neutrons = mass # - atomic #
21. Positive ion = lost electrons; negative ion = gained electrons
22. Ground state electron configuration matches periodic table; excited state does not
23. Lewis dot structure; a dot represents a valence electron (# dots = # valence electrons)
24. *** be able to calculate the atomic mass of an element, given masses and percent abundance of isotopes (% ÷ 100 * mass and add them all together)

Nuclear Chemistry

1. Stability of an isotope is based on the ratio of neutrons to protons in the nucleus
2. Most nuclei are stable. Some are unstable (atomic # 83 and above- naturally)- they will spontaneously decay, emitting radiation
3. Each radioactive isotope has a specific mode and rate of decay. (Table N + O)
4. Transmutations – a change in the nucleus of an atom that converts it from one element to another. This can occur naturally or can be artificial (2 reactants) by the bombardment of the nucleus by high energy particles (+ β →) or natural (1 reactant- Table N)
5. Spontaneous decay – release of alpha, beta, positrons and/or gamma radiation from the nucleus of an unstable isotope
6. Use table O for mass, charge, symbol (gamma has the most penetrating power b/c small mass and no charge)
7. Fusion: Hydrogen + Hydrogen → Helium + energy (only on the sun)
8. Fission – is the splitting of a heavy nucleus into 2 lighter nuclei. Reactants: U-235, Pu- 239
9. Fusion – requires very high temperature; lighter nuclei combine to form heavier
10. Fission – produces radioactive wastes that remain for long periods of time; heavier nuclei break down to form lighter nuclei
11. Radon-222, Krypton-85, N-16 are radioactive wastes that can be safely dispersed into the environment
12. Solve nuclear equations by making the mass and charge = on both sides of the →
13. Energy released in a nuclear reaction (fission or fusion) comes from a fractional amount of mass converted into energy. Nuclear changes convert matter into energy
14. Energy released during nuclear reactions is much greater than energy released during chemical reactions
15. Risks associated with radioactivity and the use of radioactive isotopes include: biological exposure (causing cancer), long term storage, disposal, and nuclear accidents
16. Radioactive isotopes have many beneficial uses:

    Medicine: short half life, quick elimination from body
    Tc-99 brain tumor
    I-131 thyroid disorder
    Ra and Co-60 for cancer treatment

    Age: Living things – C-14 : C-12
    Of rocks, minerals – U-238: Pb-206
17. Half life problems: Use table T and the arrow method

$$\#\text{ half life periods} = \frac{\text{total time}}{\text{Time of one half-life (table N)}}$$

Periodic Table

1. Elements are arranged according to atomic number
2. An elements location gives you an indication of physical and chemical properties
3. Atomic # = # of protons; identifies element
4. Mass # = # protons + # neutrons
5. Metals- left of staircase; non-metals- right of staircase
6. Metalloids: B,Si,Ge,As,Sb,Te (not Al- single, single, double, double)
7. Noble gases- group 18; Halogens – group 17; Alkali metals – group 1; Alkaline Earth metals – group 2
8. Metals- good conductors of heat and electricity, malleable and ductile
9. Elements can be differentiated by physical properties and by chemical properties
10. Groups 1,2,13-18- elements within the group have the same # valence electrons (except He)
11. Elements within the same group have similar chemical properties
12. Use Reference table S to predict trends in ionization energy, electronegativity and atomic radius (across a period, down a group)
13. Non-metals- poor conductors of heat and electricity, brittle solids, gases
14. Transition elements – groups 3-11; elements in this area make colorful solutions
15. Ionization energy – the amount of energy it takes to remove the most loosely bound electron

Formulas/Math of Chemistry

1. Compounds – can be chemically decomposed
2. A compound is a substance composed of 2 or more elements chemically combined
3. Empirical formula – simplest whole number ratio of atoms
4. Molecular formula – the actual ratio of the atoms in a molecule
5. Structural formula – the arrangement of atoms shown
6. There is a conservation of mass, energy and charge in all chemical reactions
7. Balance equations: # atoms reactants = # atoms products (add coefficients)
8. Formula mass – sum of the atomic masses of the atoms
9. Gram formula mass (molar mass) – one mole of a given substance
10. $\% = \frac{part}{whole} * 100$ Note: Part = mass of element asked about; Whole = gram formula mass
11. $\%\ hydrate = \frac{part}{whole} * 100$
12. A hydrate has water as part of its crystalline structure; to remove the water you can heat it
13. Types of chemical reactions:
    - Synthesis: two substances combine to make one product; $A + B \rightarrow AB$
    - Decomposition: compound is broken down into simpler substances: $AB \rightarrow A + B$
    - Single Replacement: Element + Compound → "New" element + "New" compound
      $A + BC \rightarrow B + AC$
    - Double Replacement: Compound + Compound → "New" compound + "New" compound
      $AB + CD \rightarrow AD + CB$
14. Naming:
    - One oxidation state (charge): metal + nonmetal -ide
      $MgCl_2$ magnesium chloride
    - One oxidation state (charge) w/ Table E: Never change name of table E ion
      Na_2SO_4 sodium sulfate
    - More than one oxidation state (charge): Roman Numerals
      metal (Roman Numeral) + nonmetal -ide
      Fe_2O_3 iron (III) oxide
    - Nonmetals: Use prefixes
      P_2O_5 diphosphorous pentoxide
15. Formula Writing: 1.) Criss-cross oxidation #'s (charges)
    2.) reduce

Calcium Nitrate	Magnesium Sulfate	copper (II) chloride
$Ca^{+2}\ (NO_3)^{-1}$	$Mg^{+2}\ (SO_4)^{-2}$	$Cu^{+2}\ Cl^{-1}$
$Ca(NO_3)_2$	$MgSO_4$	$CuCl_2$

Bonding

1. Compounds can be differentiated by their chemical and physical properties
2. Ionic compounds (metal-nonmetal); poor conductors as solids- good as liquids, aqueous
3. Molecular compounds (nonmetal-nonmetal); covalent bonding
4. Ionic bond- transfer of electrons; metal-nonmetal
5. Polar covalent bond- unequal sharing of electrons; nonmetal-nonmetal
6. Nonpolar covalent bond- equal sharing of electrons; diatomics (HOFBrINCl's)
7. Metallic Bonding- "sea of mobile electrons;" metals only (ex. Ag, Mg, Ca)
8. Triple bond- three pairs of electrons shared; double bond- two pairs shared
9. Shape and distribution of charge decides molecular polarity (SNAP)
10. Nonpolar molecules- symmetrical charge; CO_2, CH_4, CCl_4, diatomics
11. Polar molecules- asymmetrical charge; HCl, NH_3 and H_2O
12. Negative ions- gains electron(s), radius increases
13. Positive ions- loses electron(s), radius decreases
14. When a bond is broken, energy is absorbed; when a bond is formed, energy is released- Absorb to break, release to make
16. Atoms bond together to obtain a total of 8 valence electrons (become stable)
17. Noble gases- stable valence configurations; tend to not bond
18. Hydrogen bonding- H and FON; strong intermolecular forces; high melting/boiling points (H_2O, NH_3, HF); low vapor pressure
19. Dipole forces- (polar molecules) strong; higher melting points, soluble in water
20. Nonpolar- weak intermolecular forces; insoluble in water, low melting points, high vapor pressure
21. Lewis dot- one dot = one valence electron (**for compounds you must get formula first)
22. Electronegativity- a measurement of an atoms ability to gain electrons
23. Most polar or most ionic- largest electronegativity difference
24. Least polar or least ionic- smallest electronegativity difference
25. Both ionic and covalent bonds- polyatomic ions- hint answer must have at least 3 elements