

## Regents Chemistry Exam - What You Really Really Need To Know

(Based on; Jun02, Aug02, Jun03, Aug03, Jan04, Jun04, Aug 04, Jan05, Jun05, Aug 05, Jan06, Aug06, Jun07, Jun08)

Use this as your final review. Not everything you need to know is included, but everything that is here was needed for at least one of the past regents exams.

Additionally, Know Thy Reference Table – know where and how every table is used.

### ATOMIC THEORY (Periodic Table)

| Particle |                | Charge | Mass, amu      | Location |
|----------|----------------|--------|----------------|----------|
| Protons  | p <sup>+</sup> | +1     | 1              | nucleus  |
| Neutron  | n <sup>0</sup> | 0      | 1              | nucleus  |
| Electron | e <sup>-</sup> | -1     | 0 <sup>+</sup> | orbital  |

An electron does have mass, but it is insignificant relative to a proton or neutron. The mass of a proton is approximately 2000 times the mass of an electron.

- The nucleus has a net positive charge = # of p<sup>+</sup>. (It does not change for ions.)
- All atoms of the same element have the same number of protons.
- Atomic number = #p<sup>+</sup>; Atomic mass = #p<sup>+</sup> + #n<sup>0</sup>
  - #p<sup>+</sup> = At. Number;
  - #n<sup>0</sup> = At. Mass - At. Number;
  - #e<sup>-</sup> = At. No. for a neutral atom. (For a positive ion, there less electrons. For negative ions, there are more electrons. #p<sup>+</sup> never changes).
- Isotopes are the same element with a different number of neutrons. Same atomic number, same number of protons. Different mass number, different number of neutrons.
  - Isotopic notation:  ${}^{14}_6\text{C}$  = carbon 14 = C-14 =  ${}^{14}\text{C}$
- Ions are the same element with a different number of electrons. Same atomic number = same number of protons. Different charge = different number of electrons.
- Atomic masses are for elements; mass numbers are for a single isotope.
  - Atomic masses are a weighed average of each naturally occurring isotopic mass (and therefore decimals). Atomic masses are closest to the mass of most abundant isotope.
  - Know how to calculate weighted averages as the sum of the masses x % abundance/100.
- Emission (bright line or flame test) occurs when an excited electron loses E as it returns to the ground state.
  - Electrons in higher shells (principal E levels) have higher E than those in lower shells.
  - Electrons gain energy to move from a lower level to a higher level.
  - Elements can be identified by matching them to their unique line spectra.
- Electron configurations – given on the Periodic table.
  - Last number is the number of valence electrons.
  - Noble gas configuration = 8 valence electrons
  - Ground State from Periodic Table. Excited states have one e<sup>-</sup> in a higher orbital.
  - For ions – add one e<sup>-</sup> for each negative charge; subtract one e<sup>-</sup> for each positive charge. Never change the number of protons!
  - S<sup>-2</sup>, Ca<sup>+2</sup>, and Ar all have the same electron configuration & same total # of electrons.
- Rutherford Gold Foil. Shot alpha particles at gold foil. Surprisingly found some to deflect. Evidence for a dense positive nucleus. Most of Atom is empty space.
- Modern Model (also known as the wave mechanical model) is based on the work of many scientists over a long period of time. Electrons are in regions of high probability called orbitals or electron clouds.

## PERIODIC PROPERTIES (Table S)

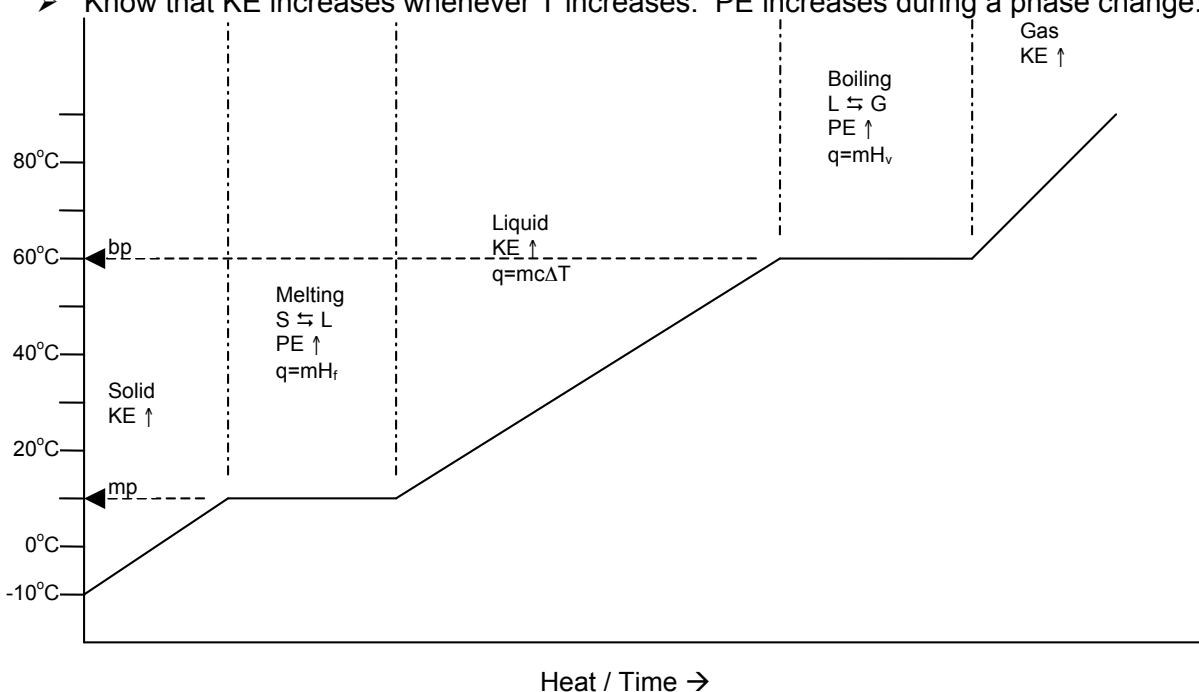
- Properties of elements in groups or periods follow trends; missing data tend to be near the average of the surrounding elements.
- Elements are arranged in order of increasing atomic number.
- Elements in the same groups have similar chemical properties due to the same number of valence electrons.
- Elements in the same period have their valence electrons in same shell.
- Metals – lose  $e^-$  in bonds; form positive ions, malleable, thermal and electrical conductors, low ionization E (left are losers), low E.N. Metals are at the left and bottom. The most reactive metals are at the bottom left (Cs, Fr). Metals are malleable and ductile, and good conductors of heat and electricity. Greatest metallic properties trend ✓
- Non-metals – gain  $e^-$  in ionic bonds; form negative ions, brittle, poor conductors, high ionization E, high E.N. (F is highest). The most reactive non-metals are at the top right (F).
- Metalloids are above and below (Ge and Sb) the “stairs”.
- Allotropes are different forms of the same element (different molecular or crystal structures), e.g, diamond and graphite,  $O_2$  and  $O_3$ . Allotropes have different chemical and physical properties.
- Atomic Radii are explained by the number of shells and/or the number of protons attracting the electrons. More shells, larger radii. More protons, stronger attraction, smaller radii. Greatest Atomic radii trend ✓
  - Negative ions have large radii. Positive ions have small radii.
- Ionization E is the energy needed to remove an electron from a gaseous atom. Greatest I.E. trend ↗
- Electronegativity is the attraction by an element for electrons (in a bond). Same trend as ionization E. F has the highest EN. Greatest EN trend ↗
- Noble gases are group 18. Noble gases are monatomic and unreactive. Most Noble gases have no E.N. value because they don't form bonds.
- Group 17 (BrINClHO<sub>F</sub>)  $F_2$  and  $Cl_2$  are gases.  $Br_2$  is a liquid.  $I_2$  is a solid.
- Every exam has a question requiring a simple graphing of a trend.

## MATTER

- Pure Substances
  - Elements – only one capital letter. Cannot be broken down further.
  - Compounds –  $\geq 2$  elements. Formed or broken down by chemical means (reactions).
  - Different substances have different chemical and physical properties, even if they are the same element, e.g.,  $O_2$  and  $O_3$  (they have different molecular structures)
- Mixtures – separate by physical methods.
  - Homogeneous – uniform throughout, solutions - look for “(aq)” indicating an aqueous solution. Homogeneous solutions are often separated by evaporation or distillation due to different phys prop, e.g., boiling pts)
  - Heterogeneous – non-uniform, often separated by filtration.
- Know the difference between chemical and physical properties.
  - Chemical properties require a change in the chemistry (a chemical reaction) to measure, i.e., a reaction with a new compound is formed, e.g., Mg burns. “Reacts” is a dead give away.
  - Physical properties are just a different arrangement of the particles, e.g., dissolving or phase changes.
- Every exam has questions with particle diagrams (■◆● type symbols to represent atoms). Know how to show a mixture vs. pure compound or element. Know that a diatomic element is an element and not a compound.

| Phase        | Macroscopic   | Microscopic  |
|--------------|---|--|
| Solid (s)    | Definite shape, definite volume                       | Strong attractions, ∴ rigid geometric pattern of close packed vibrating particles (atoms, ions or, molecules). |
| Liquid (l)   | Definite volume, takes shape of container             | Some attractions, ∴ particles can flow. Particles (molecules) still close packed.                              |
| Gas (g)      | Takes shape and volume of container                   | "No" attractions, ∴ particles far apart. Particles in rapid random motion.                                     |
| Aqueous (aq) | <i>Not really a phase, homogeneous water solution</i> | <i>Water molecules surrounding particles (ions or molecules)</i>   |

- Phase Changes (P.E. changes, increase during heating, decrease during cooling):
  - S + heat → L, melting or fusion, endothermic
  - L → S + heat, freezing, exothermic
  - L + heat → G, boiling or vaporization, endothermic
  - G → L + heat, condensation, exothermic
  - S + heat → G, sublimation, endothermic
  - G → S + heat, deposition, exothermic
- Heating / Cooling Curves
  - Know where each of the phase changes and phases occur on a heating or cooling curve. At phase changes, two phases are in equilibrium at the same temperature.
  - Know that when a substance has sufficient KE, the attractive forces (PE) can be overcome and a phase change will occur.
  - Know that KE increases whenever T increases. PE increases during a phase change.



- Know heat change formulas and how to use them (Tables T and B),  $q = mc\Delta T$ ,  $q = mH_f$  for calorimetry calculations.
- Heat is a flow of Energy. Heat (thermal energy) always flows from a warmer object to a colder object.
- Any apparent violation of the law of conservation of energy is usually due to heat loss to the environment.
- Temperature is a measure of the average kinetic energy.
  - If the temp increases, the KE increases and visa versa.
  - Average KE only depends on T, not mass, e.g., 10 g of water at 50°C has a higher ave. KE than 1000 g of water at 40°C.
  - The KE is related to the motion (and mass) of the particles.

- Know how to use Table H to determine the vapor pressure at any temperature. When  $P_{\text{vap}} = P_{\text{atm}}$ , boiling will occur. Vapor pressure depends on temperature because molecules have more kinetic energy.

### **GASES** (Tables A and T)

- Know the PTV relationships and how to show each graphically, e.g, P vs T. (PV is indirect and a hyperbole.)
- Know how to use the Combined Gas Law equation. T must be in K!!
- Equal volumes of any gas contain an equal number of moles (or molecules) at the same T and P, e.g., 4 L of  $\text{O}_2(g)$  has the same number of molecules as 4L of  $\text{CO}_2(g)$  or 4L of  $\text{CH}_4(g)$ .
- Real gases behave like ideal gases when they have weak attractive forces and the particles are far apart (high Temp and low P).

### **MATH** (Tables C, D, and T)

- Significant Figures – Atlantic / Pacific rule.
  - Least # sig figs when you multiply or divide.
  - Least # of decimal when you add or subtract.
- Know how to calculate % error, See Table T.
- Use comma = decimal for mL  $\leftrightarrow$  L
- Convert from  $^{\circ}\text{C} \leftrightarrow \text{K}$ , Table T

### **CHEMICAL FORMULAS AND REACTIONS** (Tables E and T)

- Gram formula mass = sum of all the atomic masses. Multiply each element by its subscripts (e.g.,  $\text{Ba}(\text{NO}_3)_2 = \text{Ba } 137 \times 1, \text{ N } 14 \times 2, \text{ O } 16 \times 6$ )
- Empirical Formulas – lowest whole # ratio, i.e., subscripts can not be reduced.
- Molecular Formula subscripts must be some multiple of the E.F. subscripts. Molecular Formula Mass is a whole number multiple of the empirical formula mass.
- % Composition – total mass for one element / gram formula mass. Know how to calculate the % water in a hydrate (theoretical and experimental).
- Know how to calculate the number of moles in a given mass of a substance (See Table T).
- Equal moles contain an equal number of particles.
- Names / Formulas
  - BrINClHO F – diatomic elements. All are gases except  $\text{Br}_2(l)$  and  $\text{I}_2(s)$
  - Binary (two elements, two capital letters) mostly end in “ide”. For “ate” or “ite” go to Table E. Roman numerals give the charge / oxidation number for metals. They are NOT the subscript, e.g, Iron (III) oxide is  $\text{Fe}_2\text{O}_3$ . Copper(II) hydroxide is  $\text{Cu}(\text{OH})_2$ . They like to try and catch you with things like tin (IV) oxide  $\text{SnO}_2$  or titanium (II) oxide,  $\text{TiO}$ .
  - Molecular compounds are named with prefixies, e.g.,  $\text{CO}_2$  carbon dioxide,  $\text{N}_2\text{O}_5$  dinitrogen pentoxide. IUPAC name for  $\text{H}_2\text{O}$  is water and  $\text{NH}_3$  is ammonia.
  - The dot in a hydrate means “+”. The number of  $\text{H}_2\text{O}$  in  $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$  is 7. The number of O atoms is 11. (The number of moles of O atoms in 1 mole of  $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$  is 11.)
- Reactions
  - The coefficients represent the number of moles of each substance in a chemical reaction.
  - Law of conservation of matter – must have the same number of atoms on each side of arrow. Know how to show with particle (Frizzle) diagrams. (Charge and energy are also conserved during a reaction).
  - Know the types of chemical reactions we discussed in class.
    - Synthesis: two things become one ( $\text{A} + \text{B} \rightarrow \text{C}$ )
    - Decomposition: one thing becomes two ( $\text{C} \rightarrow \text{A} + \text{B}$ )
    - Single replacement: an element and a compound forms a new element and a compound ( $\text{A} + \text{BC} \rightarrow \text{AC} + \text{B}$ )
    - Double replacement: two compounds form two new compounds ( $\text{AB} + \text{CD} \rightarrow \text{AD} + \text{CB}$ )
  - Mole  $\leftrightarrow$  mole ratios problem

- Math: given moles / it's coefficient = x / it's coefficient. (Mole ratios in reactions do not depend on subscripts.)
- Subscripts represent the number of moles of atoms in a compound.

### **BONDING** (Table H, S)

- Lewis dot structures: a dot represents a **valence** electron (B has 5 e<sup>-</sup> but only 3 valence e<sup>-</sup>).
  - For atoms, start with 2 on one side, and then spread out the rest on the three remaining sides.
  - Covalent – show as sharing with 8 e<sup>-</sup> per atom
  - Ionic – show as ions, show charge on metal with no e<sup>-</sup>, show bracket around non-metal with 8 e<sup>-</sup>. Show charge on both metal and non-metal
- Covalent bond – sharing of **two** electrons, two non-metals.
- Ionic bond – transfer of electrons, metal + non-metals.
- Metallic – sea of electrons, metal + metal. Highly mobile valence electrons explain conductivity.
- An ionic compound containing a polyatomic ion, e.g., MgSO<sub>4</sub> has both ionic and covalent bonding.
- Polar covalent bonds – between **different** nonmetal elements. The most polar have the greatest electronegativity difference (Table S).
- Non polar molecules have a symmetrical distribution of charge (polar bonds). Examples are CO<sub>2</sub> and CF<sub>4</sub>.
- O<sub>2</sub> has a double bond (4 shared e<sup>-</sup>, 2 shared pairs). N<sub>2</sub> has a triple bond (6 shared e<sup>-</sup>). All other BRINCLHOF have single bonds. CO<sub>2</sub> has two double bonds.
- Polar molecules have attractive forces holding them together.
  - Attractive forces (intermolecular) are strongest between molecules that can form hydrogen bonding. HF, HO, HN. Water forms strong H bonds.
  - A higher attractive (intermolecular) force means higher boiling points and less volatility.

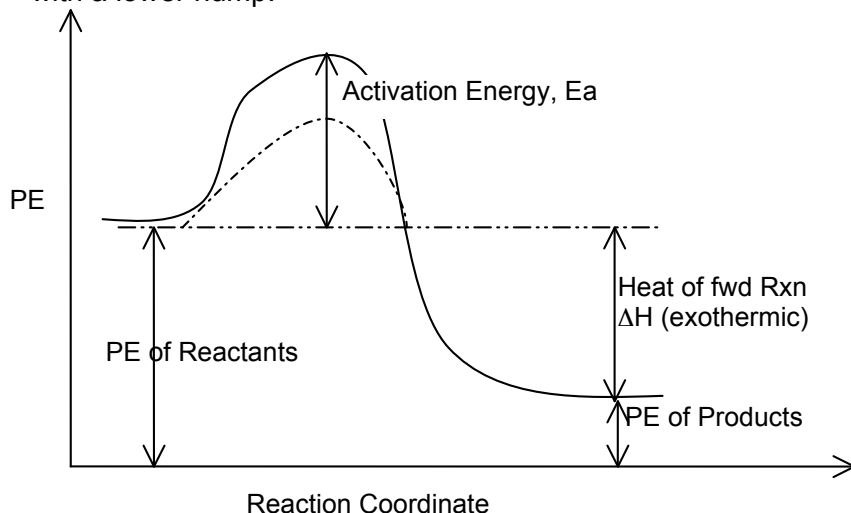
| <b>Properties</b>    | <b>Ionic Compound</b>   | <b>Covalent (Molecular) Cmpd</b>   | <b>Metals</b>                            |
|----------------------|---|--|--|
| <b>Melting point</b> | Highest   | Lowest   | High                                     |
| <b>Conductivity</b>  | Aqueous <b>solutions</b> conductive (electrolytes). Molten or “fused” conduct | Non conductive   | Good conductors of electricity and heat. |
| <b>Volatility</b>    | Non volatile  | Volatile, have odor  | Non volatile                             |
| <b>Solubility</b>    | Soluble in polar solvents - water   | Polar soluble in polar solvents. Non-polar soluble in non polar solvents | Insoluble in everything                  |
| <b>Hardness</b>      | Hard and brittle  | Soft   | Hard (malleable and ductile)             |

- Flow of charged particles (ions or electrons) are necessary for conductivity (also as electrolytes in sol'n)

### **KINETICS** (Tables I, )

- Faster rates depend on the higher concentrations of reactants, higher temperature, more surface area, and nature of reactants (aqueous ionic are fast); all explained by the number of effective collisions (collision theory) .
- Breaking bonds requires (absorbs) E, forming bonds releases E. (I + I → I<sub>2</sub> releases E)
- Potential E diagrams
  - Exothermic (shown) – Product PE lower than Reactant(s) PE

- A catalyst lowers the activation energy but does not effect the equilibrium (provides an alternative reaction pathway). Shown on PE diagram starting and ending the same with a lower hump.



- Entropy – less order, less structure
  - more gas particles = greater entropy
  - Higher Temp > lower temp; G > L > S
- Equilibrium reactions
  - Concentrations are constant, not equal. Reactions do not stop (dynamic).
  - Forward and reverse reactions have equal **RATES** (not concentrations)
- Equilibrium Shifts (LeChatelier) shift to minimize the stress (change).
  - Increase concentration – shifts away from more (explain by collision theory).
  - Increase Temp. – shifts from “heat” in equation
  - Increase pressure – shifts to side with less moles of gas.
- Heats of reaction -Know how to use Table I.
  - Negative values are exothermic. Values are given for the number of moles (coefficients) shown in the equation in Table I.
  - Exothermic heat of reaction will show as a positive value when written as part of the equation, e.g.,  $C + O_2 \rightarrow CO_2 + 394 \text{ kJ}$  ( $\Delta H = -394\text{kJ}$ )

**SOLUTIONS** (Tables I, F, G, T) All solutions are **homogeneous mixtures**.

- Solute is the stuff being dissolved. Solvent is the stuff the solute is dissolved in.
- Solutions have lower freezing points and higher boiling points than pure substances.
- Ionic compounds form ions in solution and are electrolytes (conducts electricity when dissolved in water). Conductivity depends on the concentration of ions. Only soluble compounds are electrolytes (see Table F).
- Know how to use Table F to determine if a compound is soluble.
- Like dissolve like, polar and non-polar will not form solutions together.
  - Water is polar! Substances that dissolve in water are polar, those that don't aren't.
- Know how to calculate molarity. See table T. Remember comma = period for mL  $\rightleftharpoons$  L; 1,500 mL = 1.5 L. When making a molar solution add water to a final (total) volume.
- Know how to use Table G. Values are for 100 g of water.
  - Solids are more soluble at high temperature (see Table G). The lines represent saturation limits for each compound.
  - Additional solute will precipitate out or not dissolve.
- The rates (not amount) of solution are affected by temperature, mixing, and surface area – explained by collision theory.
- The solubility of solids is not affected by pressure.
- Gases are more soluble at high pressure and low temperature.
- Saturated solutions are an equilibrium between the solid  $\rightleftharpoons$  and the dissolved substance.

- Solutions can be exothermic or endothermic. Exothermic solutions will increase the water temperature.  $\Delta H < 0$  on Table I.

### ORGANIC (Tables P, Q, R)

- Carbon can form single, double, or triple bonds, but will always have 4 and only 4 bonds. Carbon can form long chains or rings.
- Hydrocarbons contain only C and H. A longer chain means stronger intermolecular forces and therefore a higher boiling point.
- Compounds with the same functional groups have similar chemical properties.
- Know how to recognize functional groups on Table R; count the CC double bonds and CO single / double bonds. Every exam has one or two of these questions.
- Organic molecules with oxygen tend to be polar (H bonding if have -OH or -NH) and have higher boiling points and water solubility.
- Only organic acids (R-COOH) are electrolytes.
- Saturated – no double bonds, alkanes; unsaturated contains double bonds, alkenes or alkynes – see table Q.
- Isomers have the same chemical formula, but different structural formulas. They do not necessarily have the same functional groups, e.g., alcohols and ethers may be isomers.
- Types of Reactions (one of these is usually on the exam).
  - Polymerization reactions – joining many small (monomers)
  - Fermentation – sugar + enzyme  $\rightarrow$  CO<sub>2</sub> + ethanol
  - Esterification – acid + alcohol  $\rightarrow$  ester
  - Substitution (saturated) –  $C_xH_y + Br_2 \rightarrow C_xH_{y-1}Br + HBr$
  - Addition (unsaturated) –  $C_xH_y + Br_2 \rightarrow C_xH_yBr_2$
  - Combustion –  $C_xH_y + O_2 \rightarrow CO_2 + H_2O$
  - Saponification – makes soap from fat or oil.

### REDOX (Table J)

- Most common oxidation numbers: = zero for free elements (including metals); = charge for ions; = -2 for O, = +1 for H; = 1 for Group 1 in cmpds; = 2 for Group 2 in cmpds.
- Redox reactions are due to gain or loss of electrons. LEO say GER;
  - loss of e<sup>-</sup> is oxidation;  $Ag \rightarrow Ag^+ + e^-$ ; Metals are losers (generally)
  - gain of e<sup>-</sup> is reduction;  $Br + e^- \rightarrow Br^-$
- Elements undergoing oxidation increase their oxidation numbers, elements undergoing reduction reduce (decrease) their oxidation number.
- Half reactions represent only oxidation or reduction. The number of electrons must provide conservation of charge.
- Number of electrons lost = number of electrons gained in a REDOX reaction.
- Single replacement reactions are always Redox – know how to use Table J to predict which will occur. More active free metal will replace less active metal in a compound. It helps to remember that the coinage metals (Ag, Au, Cu) are relatively non reactive.
- Double replacements never are Redox.
- Electrochemical cells
  - RedCat – reduction at cathode, ca+hode is positive;
  - AnOx – oxidation at anode.
  - Electrons flow from a to c (a to z) anode to cathode.
  - Voltaic (batteries) – chemical E  $\rightarrow$  electrical E; spontaneous; produce E.
  - A salt bridge allows the flow of ions.
- Electrolytic cells
  - Electrical E  $\rightarrow$  chemical change, requires energy; cathode become negative (but AnOx and RedCat still apply).
  - Negative ions flow toward the positive anode and positive ions flow toward the negative cathode.

### ACID BASE (Tables J, K, L, M, T)

- Arrhenius acid – give hydrogen ions as only positive ions in solution. A hydrogen ion is better described as the hydronium ion,  $\text{H}_3\text{O}^+$ . (A hydrogen ion is just a proton.)
- Arrhenius base – give hydroxide ions as only negative ions in solution (Metal hydroxides only,  $\text{CH}_3\text{OH}$  is not a base nor an electrolyte!)
- Acids and bases are electrolytes (as are soluble ionic compounds, salts).
- Acids react with some metals to form  $\text{H}_2$ . Active metals above  $\text{H}_2$  on Table J react with acids.
- Acids  $\text{pH} < 7$ , high  $\text{H}^+$  concentration. Bases  $\text{pH} > 7$  high  $\text{OH}^-$  concentration. Every 1 unit change in pH represents a 10X change in hydronium ( $\text{H}^+$ ) concentration. pH of 3 is 1000x more acidic than pH of 6.
- Know how to use the Titration Equation (see table T). Both acid and base may be delivered from a buret.
- Know how to use indicator Table M to predict color changes. Phenolphthalein is pink with bases. Litmus is blue with a base, red with an acid.
- Neutralization: Acid + Base  $\rightarrow$  salt (ionic compound) + water; or hydrogen ions react with hydroxide ions to form water.
- An alternative theory defines acids as hydrogen ion donors and bases as hydrogen ion acceptors.

### NUCLEAR CHEM (Table N, O, T)

- The stability of an isotope depends on the ratio of neutrons to protons.
- Radioactive decays modes and half-life's – see Table O
- Particle charge and mass given in Table O:  $\frac{\text{mass}}{\text{charge}} X$ .
  - Higher mass have lower penetration power. Alpha is least, gamma is greatest.
- Nuclear reactions must balance (conserve) mass and charge.
  - Nuclear reactions lose some mass due to conversion of mass (matter) to energy.
  - Nuclear reactions produce more energy than chemical reactions due to the conversion of mass to energy.
- Transmutations – changes one element into another.
  - Natural Transmutation – spontaneous radioactive decay. Only one thing on left side of arrow.
  - Artificial Transmutation – bombardment of element with high E particle. Two things on left side of arrow.
- Fission (division) is a heavier element forming lighter ones.
- Fusion is two lighter elements form a heavier one, eg  $\text{H} + \text{H} \rightarrow \text{He}$  on the sun.
- Know some risks (mutations, cancer) and benefits (energy, medicinal) of nuclear chemistry.
- $^{14}\text{C}$  used for radioactive dating,  $^{131}\text{I}$  is used to treat thyroid disorders.  $^{60}\text{Co}$  is used to treat cancer.
- Half lives are constants, they do not change during decay.
  - Count the halving of the mass (arrows), e.g,  $80\text{g} \rightarrow 40 \rightarrow 20 \rightarrow 10 \rightarrow 5\text{g}$  is 4 half lives. If  $1/8$  is left;  $1/8 \rightarrow 1/4 \rightarrow 1/2 \rightarrow 1$  is 3 half lives.

### LABS

- Know the procedure and math for:
  - Density by displacement
  - Water of Hydration
  - Titrations of acids and bases
  - Percent error

*That's all, not too bad. Good Luck!!*