

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

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.

### ATOMIC THEORY (*Periodic Table*)

Particle		Charge	Mass, amu	Location
Protons	p <sup>+</sup>	+1	1	nucleus
Neutron	n <sup>o</sup>	0	1	nucleus
Electron	e <sup>-</sup>	-1	0*	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 nuclear charge (positive) = # 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> PLUS #n<sup>o</sup>  
#p<sup>+</sup> = Atomic Number;  
#n<sup>o</sup> = Atomic Mass - Atomic Number;  
#e<sup>-</sup> = Atomic No. for all atoms (all are neutral)  
Cations + have less electrons. Anions - have more electrons. The #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, different number of neutrons.  
Isotopic notation:  ${}^1_6\text{C}$  = carbon 14 = C-14 =  ${}^{14}\text{C}$   
*This carbon atom (this isotope of carbon) has mass of 14 amu, 6 protons, 6 electrons, 8 neutrons*

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 weighted average of each naturally occurring isotopic mass (and therefore decimals).  
Atomic mass on periodic table are closest to the mass of the 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 (emits) energy as it returns to the ground state.  
Electrons in higher shells (principal energy levels) have higher energy 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, which are known.
- 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, he 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, heat and electrical conductors, low ionization energy, low electronegativity values. 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.
- Non-metals – gain  $e^-$  in ionic bonds; form negative ions, they are brittle, poor conductors of heat or electricity, have a high first ionization energy, high electronegativity values (F is highest). The most reactive non-metals are at the top right not including the noble gases.
- Metalloids are touch the “staircase”. But do not include Al or Po (AlPo is the dogfood exception)
- 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. Negative anions have a larger radius than the atoms they formed from. Positive cations have a smaller radius than the atoms they formed from.
- First Ionization energy is the energy needed to remove a mole of electrons from a mole of atoms
- Electronegativity is the attraction by an element for electrons (in a bond). F has the highest electronegativity value of 4.0 this scale is arbitrary (there are no units) and the scale is arbitrary in that the numbers don’t actually mean anything, they are just for ranking.
- Noble gases are group 18. Noble gases are monatomic and unreactive. Most Noble gases have no electronegativity values because they don’t form bonds.
- Group 17 (halogens)  $F_2$  and  $Cl_2$  are gases.  $Br_2$  is a liquid.  $I_2$  is a solid. This is primarily due to electron dispersion attraction (an intermolecular attraction)

## MATTER

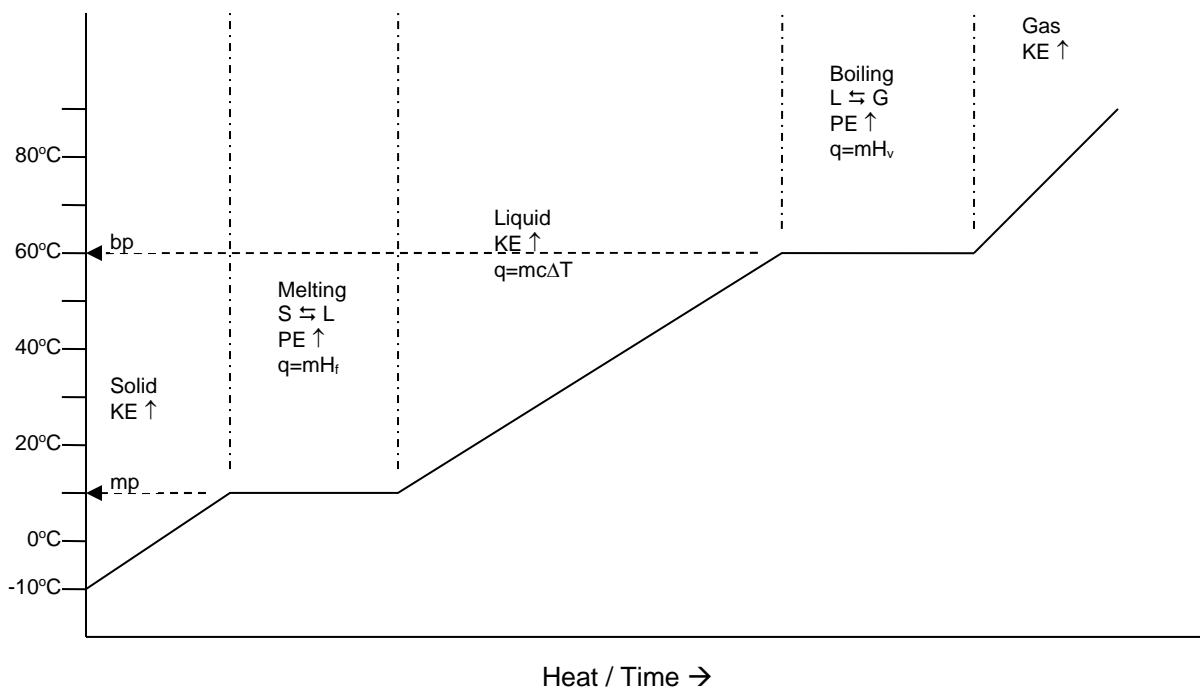
- Pure Substances include elements and compounds.  
Elements cannot be broken down further.  
Compounds contain 2 or more elements bonded together. They are formed or broken down by chemical means (reactions).  
Different substances have different chemical and physical properties, even if they are the same element, e.g.,  $O_2$  and  $O_3$  (these allotropes have different molecular structures)
- Mixtures – these are physical blends of stuff, not chemically bonded together. They can be separated by physical methods, not chemical reactions.  
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 giveaway. 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 (potential energy changes, increase during heating, decrease during cooling):

- Solid + heat → Liquid, melting or fusion, endothermic
- Liquid → Solid + heat, freezing, exothermic
- Liquid + heat → Gas, boiling or vaporization, endothermic
- Gas → Liquid + heat, condensation, exothermic
- Solid + heat → G, sublimation, endothermic
- Gas → Solid + 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 formulas and how to use them:  $q = mc\Delta T$ ,  $q = mH_f$  and  $q = mH_v$ .
- Heat is a flow of energy. Heat (thermal energy) always flows from a warmer object to colder
- 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 temperature increases, the Kinetic energy increases and vice versa.  
Average KE only depends on temperature, not mass. 10 g of water at 50°C has a higher average kinetic energy 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. That means when vapor pressure equals the air pressure that is a boiling point. Boiling depends on 2 things: internal or inter molecular attractions and air pressure pushing down on the surface of a liquid. Vapor pressure depends on temperature because molecules have more kinetic energy.

### GASES (Tables A and T)

- Know the pressure, volume and temperature relationships and how to show each graphically, ex: P vs T. (P and V is inversely proportional and makes a curved graph)
- Know how to use the Combined Gas Law equation. Temperature only in Kelvin
- Avogadro's Hypothesis: Equal volumes of different gases at the same temperature and pressure contain an equal number of moles (or molecules) ex: 4 L of  $O_2(g)$  has the same number of molecules as 4L of  $CO_2(g)$  or 4L of  $CH_4(g)$ . THIS IS FOR GASES.
- Real gases behave most like ideal gases when they have weak attractive forces and the particles are far apart (high temperature, low pressure).

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

- Significant Figures – Atlantic / Pacific rule. Use least # SF's when you do math.
- Know how to calculate % error, See Table T.
- Convert from °C to K,  $K = C + 273$  Always, always, always use a freaking formula. Write it out and get it right, do NOT just do math in your calculators. LOOK at it on paper.

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

- Gram formula mass = Molar Mass. Multiply each element by its subscripts (e.g.,  $Ba(NO_3)_2 = Ba\ 137 \times 1, N\ 14 \times 2, O\ 16 \times 6$ ) SUM UP THE LAST COLUMN. Do this on PAPER too!
  - Empirical Formulas – lowest whole # ratio, i.e., subscripts cannot be reduced. (it's math, not chem)
  - Molecular Formula subscripts must be some multiple of the empirical formula 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  
HONCIBrIF twins – diatomic elements. All are gases except  $Br_2(L)$  and  $I_2(S)$   
Binary (two elements, two capital letters) mostly end in "ide".  
Roman numerals give the ion charge for metals. They are NOT the subscript, e.g, Iron (III) oxide is  $Fe_2O_3$ .  
Copper (II) hydroxide is  $Cu(OH)_2$ . The regents likes to try and fool you with things like tin (IV) oxide  $SnO_2$  or titanium (II) oxide  $TiO$ .

Molecular compounds are named with prefixes, e.g., CO<sub>2</sub> carbon dioxide, N<sub>2</sub>O<sub>5</sub> dinitrogen pentoxide. IUPAC name for H<sub>2</sub>O is water and NH<sub>3</sub> is ammonia.

The dot in a hydrate means “+”. The number of H<sub>2</sub>O molecules in MgSO<sub>4</sub>·7H<sub>2</sub>O is 7. The number of O atoms is 11. (The number of moles of O atoms in 1 mole of MgSO<sub>4</sub>·7H<sub>2</sub>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 ( $A + B \rightarrow C$ )
- Decomposition: one thing becomes two ( $C \rightarrow A + B$ )
- Single replacement: an element and an AQ solution forms a new element and new AQ solution which looks like this:  $A + BC_{(AQ)} \rightarrow AC_{(AQ)} + B$
- Double replacement: two AQ solutions form a new AQ solution and a solid compounds which looks like this:  $AB_{(AQ)} + CD_{(AQ)} \rightarrow AD_{(AQ)} + CB_{(S)}$

### BONDING (Table H, S)

- Lewis dot structures: a dot represents a **valence** electron (Boron, rhymes with... has 5 electrons but only 3 valence electrons).

Drawing Lewis Dot diagrams 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 electrons per atom – NO BRACKETS

Ionic – USE BRACKETS to show ions, put charge on metal ion with no dots, put brackets around the non-metal with 8 dots. Show charges on both metal and non-metal

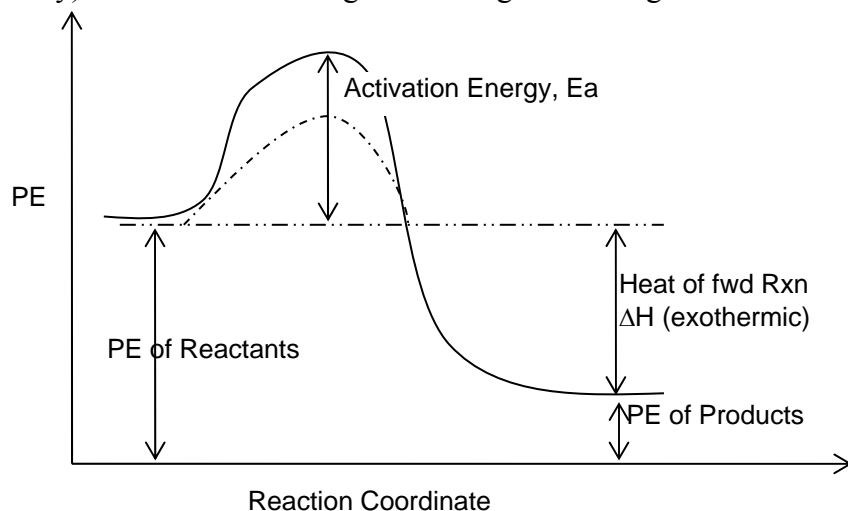
- Covalent bond – sharing of two electrons, two non-metals. No brackets
- Ionic bond – transfer of electrons, metal + non-metals. Use brackets
- Metallic – sea of loose valence electrons. Highly mobile valence electrons explain conductivity and malleability.
- An ionic compound containing a polyatomic ion, e.g., MgSO<sub>4</sub> has both ionic + covalent bonding.
- Polar covalent bonds – nonmetal atoms with a difference in electronegativity values
- Nonpolar 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 nonpolar covalent bond (4 shared e<sup>-</sup>, 2 shared pairs). N<sub>2</sub> has a triple nonpolar covalent bond (6 shared e<sup>-</sup>). All other HONClBrIF twins have single nonpolar covalent bonds.  
CO<sub>2</sub> has two polar 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. Must contain H in a polar molecule)  
A higher attractive (intermolecular) force means higher boiling points and less volatility.

Properties	Ionic Compound	Covalent/Molecular Compounds	Metals
Melting point	Highest	Lowest	High
Conductivity	Aqueous solutions conductive (electrolytes). Molten or "fused" conduct	Non conductive	Good conductors of electricity and heat.
Volatility (easily turn to gases)	Non volatile	Volatile, have odor	Non volatile
Solubility	Most are soluble in polar solvents like water	Polar soluble in polar solvents. Non-polar insoluble in water	Always insoluble
Hardness	Hard + brittle	Soft	Hard (malleable + ductile)

### KINETICS

Faster rates of reaction 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). Anything that increases collisions, increases rates of reaction.

- Breaking bonds requires (absorbs) energy, When bonds form energy is released.  
ex:  $I + I \rightarrow I_2$  releases energy.  $N_2 \rightarrow N + N$  is endothermic)
- Potential E diagrams  
Exothermic (shown) – Product potential energy is lower than the Reactants potential energy  
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 or decrease energy (heat), change in pressure affects only gases, or add/remove reactants or products cause shifts

- Heats of reaction - Know how to use Table I.  
Negative  $\Delta H$  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** 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.
- MOST 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 liquids are IMMISCIBLE + 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. When making a solution add water last to finish volume.
- Know how to use Table G. Values are for 100 g (which is 100 mL) 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. NOTHING CAN SUPERSATURATE ON TABLE G.
- 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. Pressure only affects gases
- Gases are more soluble at high pressure and low temperature.
- Saturated solutions are an equilibrium between the solid  $\rightleftharpoons$  and the dissolved substance.
- Solution formation can be exothermic or endothermic. Exothermic solution formation will increase the water temperature.

**ORGANIC**

- Carbon can form single, double, or triple bonds, but will always have 4 and only 4 bonds. Carbon can form 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<sub>2</sub>) and have higher boiling points and water solubility.
- Organic acids (R-COOH) are electrolytes.
- Alkanes are Saturated hydrocarbons– no double or triple bonds
- Alkenes are unsaturated - they contain ONE double bond in our class.
- Alkynes are unsaturated - they contain ONE triple bond in our class.
- Isomers have the same chemical formula, but different structural formulas. They do not have to have the same functional groups, e.g., alcohols + ethers may be isomers.

- Types of Organic Reactions (one of these is usually on the exam).  
 Polymerization reactions – joining many small molecules (monomers)  
 Fermentation – sugar + enzyme  $\rightarrow$  CO<sub>2</sub> + ethanol  
 Esterification – acid + alcohol  $\rightarrow$  ester  
 Substitution (saturated) – C<sub>x</sub>H<sub>y</sub> + Br<sub>2</sub>  $\rightarrow$  C<sub>x</sub>H<sub>y-1</sub>Br + HBr  
 Addition (unsaturated) – C<sub>x</sub>H<sub>y</sub> + Br<sub>2</sub>  $\rightarrow$  C<sub>x</sub>H<sub>y</sub>Br<sub>2</sub>  
 Combustion – C<sub>x</sub>H<sub>y</sub> + O<sub>2</sub>  $\rightarrow$  CO<sub>2</sub> + H<sub>2</sub>O  
 Saponification – makes soap from fat or oil plus base.

### REDOX (Table J)

- Most common oxidation numbers: = zero for ALL elements (including metals); ion oxidation number is their CHARGE ex: -2 for O, +1 for H
- Redox reactions are due to gain or loss of electrons. LEO the Lion says GER;  
Loss Of Electrons is oxidation; Ag  $\rightarrow$  Ag<sup>+</sup> + e<sup>-</sup> this silver atom is oxidized into Ag<sup>+</sup>  
 Metals will lose electrons  
 The Gain of Electrons is Reduction; nonmetals will gain electrons  
 ex Br + e<sup>-</sup>  $\rightarrow$  Br<sup>-</sup> this Bromine atom is reduced into the bromide anion
- 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 (higher on list) atoms will replace less active atoms in a compound.
- Double replacements never are Redox. There is NO CHANGE in oxidation number, just partners.
- Electrochemical cells  
 RedCat – reduction at cathode, The ca+hode is positive;  
 Ergo: oxidation only happens at the anode.  
 Electrons flow from anode to cathode.  
 Voltaic (batteries) – chemical energy makes electrical energy; spontaneous  
 A salt bridge allows the flow of ions.
- Electrolytic cells  
 Electrical energy makes chemistry happen, requires energy



## ACID BASE

Arrhenius acid – give hydrogen ions as only positive ions in solution. A hydrogen ion can be described as the (dumb) hydronium ion,  $\text{H}_3\text{O}^+$ . (A hydrogen ion is just a proton.)

- Arrhenius base – have hydroxide ions as only negative ions in solution
- Acids + bases are electrolytes (as are soluble ionic compounds, salts).
- Acids react with most metals to form  $\text{H}_2$ . Active metals above  $\text{H}_2$  on Table J.
- Acids have a pH less than 7, high  $\text{H}^+$  concentration. Bases have pH greater than 7 with a high  $\text{OH}^-$  concentration. Every whole number change in pH represents a 10X change in hydrogen ( $\text{H}^+$ ) concentration. A solution with pH of 3 is 1000x more acidic than pH of 6.
- Know how to use the (FIXED) Titration Equation. Acids + bases may be delivered from burets.
- 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 you could say... hydrogen ions react with hydroxide ions to form water.
- An alternative theory defines acids as  $\text{H}^+$  donors and bases as  $\text{H}^+$  acceptors.

## NUCLEAR CHEM

- The stability of an isotope depends on the ratio of neutrons to protons.
- Radioactive decays modes and half-life is on 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 – spontaneous radioactive decay. Just happens, we can't control it.  
Artificial – bombardment of element with high energy particle. Caused by humans.
- Fission (division) is a heavier element splitting into 2 or more 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), ex:  $80\text{g} \rightarrow 40 \rightarrow 20 \rightarrow 10 \rightarrow 5\text{g}$  is 4 half lives.  
If  $1/8$  is left;  $1/8 \rightarrow 1/4 \rightarrow 1/2 \rightarrow 1$  is 3 half lives.

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