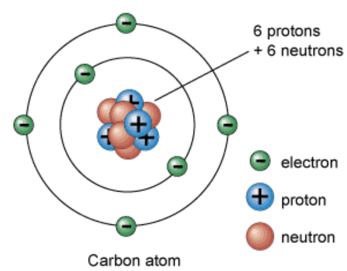
CHEM REVIEW: 281 ways to pass the Chemistry Regents

Safety, Scientific Method and Graphing:

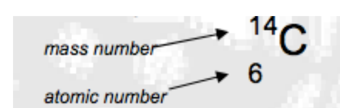
1. One Safety precaution to follow if you have **long hair** is tie back your hair
2. When you spill a liquid on your arm first rinse your arm and then tell the teacher
3. After a laboratory experiment is over **dispose of chemicals properly**
4. ***Scientific Method***: When doing an experiment, all variables must remain the same except the variable you are testing
5. ****Graphing***: Make an even scale of numbers and circle final points

Unit 1 Atomic:

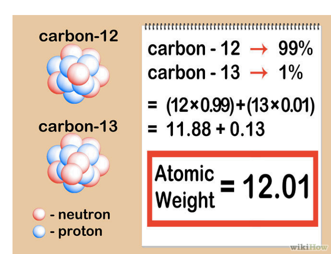
6. Today we know an atom contains **subatomic particles** (protons, neutrons, electrons) that make it divisible
7. *******Protons** have a +1 charge.
8. *******Neutrons** have a 0 charge.
9. ****Mass of Neutron = 1 amu (atomic mass unit) = Mass of Proton
10. ***Electrons** have a charge of -1 and a mass of 0.
11. *Charge of electron is -1, charge of proton +1 (same magnitude, opposite charge)
12. ***TABLE O*** shows symbols, mass and charges of particles (electrons are represented as beta)
13. *****Protons and Neutrons are located in the nucleus of an atom.
14. **Charge of an atom's nucleus** = (+) number of protons
15. **Atoms have a positively charged nucleus and negatively charged electrons located in "clouds" (**orbitals**) around an atom's nucleus.
16. *********Mass number**= #protons + #neutrons
17. *******Atomic number =** number of protons (All atoms of the same element have the same atomic number)
18. **Number of neutrons** = Mass number – atomic number



19. ********Isotopes** are atoms with the same number of protons, different number of neutrons (different mass number)
20. *********Isotope notation**: top number of isotope notation is Mass number and bottom number is atomic number



21. **Other notations**: (C-14 or Carbon-14) Number after an element represents mass number
22. **Atomic mass** is the weighted average of all the naturally occurring isotopes for that element
23. ****Average Atomic mass**
    = (isotope1 mass)(% in decimal form) + (isotope 2 mass)(% in decimal form)
24. **Abundance:** Whatever whole number the atomic mass is closest to on the Periodic Table means that isotope is most abundant
25. ***Rutherford's Gold Foil experiment** shows an atom is mostly empty space with a small, dense, positively charged nucleus.
26. **Thomson** and **Bohr's models** showed electrons present in an atom
27. *********Wave-mechanical model** (electron cloud model) shows that an orbital (cloud) is the most probable location of electrons



28. ***Neutral Atom:** An Atom has the same number of protons and electrons as long as there is no charge (total charge of 0)
29. -*****Total (Net) Charge** of an atom = # protons - # electrons
30. ** **Ion:** a charged element (it has lost or gained electrons): electron configuration will change if it is an ion (possible charges are found on PT)

24.305
Mg
12 2-8-2

31. ***Electron configuration**: Shows location of electron in their shells. Example: 2-8-2 (2 electrons in first energy level, 8 electrons in second energy level, 2 electrons in third energy level) CAN BE FOUND ON PERIODIC TABLE
32. **First shell has less energy than 2nd shell
33. First shell can have a max of 2 electrons and Second shell can have a max of 8 electrons
34. ****Valence electrons**: electrons found in the outer most shell (last number in electron configuration)
35. ****Lewis dot diagram for a single atom** just shows the valence electrons (electrons represented by dots, drawn in pairs)
36. *******When excited electron (at higher energy level) moves to **ground state** (lower energy level), a specific amount of energy is emitted (sometimes as light/**bright line spectrum**)
37. *****Excited electron configuration** is not the same as the configuration on the reference table (there will be one less electron in one energy level and one more electron in another energy level)
38. Energy emitted from excited electron can be used to determine identity of element
39. **When viewing a Bright line spectrum- elements must line up exactly to be part of the mixture in the spectrum

<u>Unit 2 Nuclear</u>:

40. *Table O shows a **positron and beta particles** have same mass (0) & opposite charge (Beta negative, positron positive)
41. **Beta particle has less mass and greater penetrating power than a +alpha particle (**Gamma radiation has the greatest penetrating power**)
42. ***All Nuclear reactions are **transmutations** (examples: fission, fusion, decays)
43. **Any element after Po is **naturally unstable** and will spontaneously decay
44. **Stable Isotopes** are not on table N (do not spontaneously emit particles)
45. ***Decay modes and half lives on **table N** (show alpha, beta, positron decay)
46. **Natural Transmutations** show Spontaneous radioactive decay = 1 reactant → 2 products (elements must change)
47. **Nuclear decays** release the decay particle
48. *****Completing nuclear equations**: remember to add up mass number and atomic numbers on each side of the arrow (must be equal on both sides)

$$\overbrace{\underbrace{^{14}_{7}N + ^{4}_{2}He}_{9}}^{18} \rightarrow \overbrace{\underbrace{^{17}_{8}O + ^{1}_{1}\ldots}_{9}}^{18}$$

49. ****Fusion**: light nuclei combine to form a heavy nucleus and a lot of energy (energy is sometimes in the form of a neutron)

$$^{2}_{1}H + ^{2}_{1}H \rightarrow {}^{4}_{2}He + \text{ENERGY}$$

$$^{1}_{0}n + ^{235}_{94}U \rightarrow ^{139}_{56}Ba + ^{94}_{36}Kr + 3\,^{1}_{0}n$$

50. Fusion produces more energy than **fission**
51. ****Nuclear reactions** (such as fission or fusion) releases more energy than a chemical reaction (redox, substitution, neutralization)
52. ****Nuclear reactions: **mass is converted into energy**
53. **Half life** the length of time it takes for ½ mass of a sample to decay
54. **1 half life = ½ sample remains, 2 half lives = ¼ sample remains, 3 half lives = 1/8 sample
55. **Half life question use table or **timeline method**

2. What fraction of a sample of N-16 remains undecayed after 42.78 seconds? Half life of N-16 on table N is 7.13 seconds.

Time elasped	Fraction left
0	1
7.13 s	½
14.26 s	¼
21.39 s	1/8
28.52 s	1/16
35.65 s	1/32
42.78 s	**1/64**

56. Radioisotopes are used for **dating of geological formations** (C-14)
57. **I-131** used to **diagnose thyroid disorders**
58. Radioisotopes can be used to **detect diseases**
59. Radioisotopes can **treat cancer** but can also cause mutations in healthy cells (Co-60)

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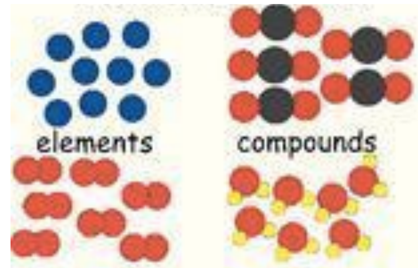
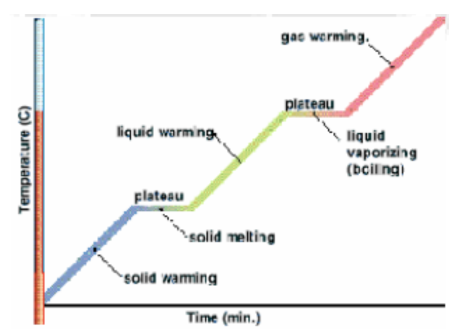


Unit 3 Matter:

60. *__Substance__ = compound or element
61. ********__Elements__ cannot be broken down by chemical means (it is an element if it is on the Periodic table/Table T)
62. **Compounds** can be broken down by chemical means
63. **Same compound = same chemical property, different compound = different chemical properties
64. *__7 diatomics__ (two of the same atom bonded together) – BrINClHOF (Br_2, I_2, N_2, Cl_2, H_2, O_2, F_2)
65. *Melting Point (MP) and Boiling Point (BP) of elements on **table T**
66. *__Liquid at a specific temperature__ MP < Specific Temperature <BP
67. *__Mixture__ can vary in proportion of its components (example: Salt water)
68. ***__Homogeneous mixtures (solutions):__ even distribution of particles (aq-dissolved in water) substance has to be soluble to mix with water
69. *__Heterogeneous mixtures__= not even throughout. Contains a substance that will not be soluble in water.
70. *When substances are **mixed**, they retain their properties
71. *Mixtures that contain substances with different density and particle size **can be separated by physical means**
72. Mixture can be separated by chromatography, distillation and filtration
73. *__Distillation__ separates liquids with different boiling points (water and alcohol)
74. *__Chromatography__- method of separating particles by solubility and polarity
75. **Evaporation** separates a salt dissolved in water
76. *__Chemical property__ is how substances react
77. ****__Chemical change__ results in the formation of a difference substance (example: burning)
78. **__Physical change__ = do not form new compounds, commonly phase changes (change in distance between molecules)
79. ***__Solids__ = atoms close together, liquid in the middle, **gas** = atoms far apart
80. Solids have a **definite shape and definite volume**
81. **Deposition** = gas to solid phase change
82. **__Sublimatio__n: Solid→Gas phase change (ex: CO_2)
83. *****In a **phase change diagram**, the flat parts represent the phase changes (Potential Energy [PE] changes and Kinetic Energy [KE] remains the same)
84. In a phase change diagram, the **sloped lines** represent heating or cooling (PE remains the same and KE changes)
85. *__Density = mass/volume__ (g/ L or g/cm^3)
86. **Higher density sinks** to bottom of tank
87. Density never changes for each element (Found on **Table S** for elements)



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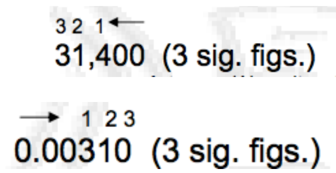
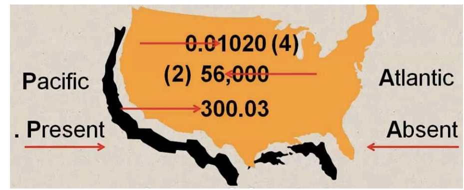
Math:

88. *Sig figs: Atlantic pacific
    -decimal **absent**, count from the first nonzero number on the **Atlantic** side/right and count all numbers to the left;
    -decimal **present**, count from the first nonzero number on the **pacific** side/left and count all number to the right

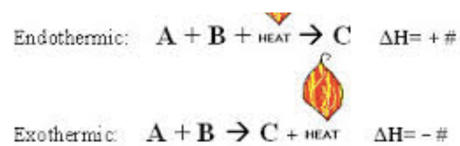


89. **When **Multiplying/Dividing**: answer should be the lowest number of sig figs
90. ***% error** formula on table T



Unit 4 Energy:

91. **Kelvin = ° C +273** formula on table T (0 ° C = 273 K)
92. **1kJ = 1000J**
93. **Forms of energy**: chemical, thermal, electromagnetic, electrical, nuclear, mechanical
94. ***Thermal energy (heat)** is measured in joules (J) = random motion of atoms and molecules
95. ********Average kinetic energy = **temperature**
96. When two substances have the same temperature, the substance with the **greater mass has more thermal energy**
97. **Heat of vaporization** is the amount of heat required to vaporize a substance (table B for water constants) = 2260J or 2.26 x 10^3 J
98. Heat of fusion (heat it takes to melt a substance) is less than heat of vaporization because it requires less heat to melt a substance than boil a substance.
99. *****Heat flows from hot to cold**
100. ****$Q= mC\Delta T$ (q is heat, m is mass, C is specific heat capacity [found on table B for water], ΔT is change in temperature) All info on table T
101. **Exothermic** = energy exits (is released)
102. ****Endothermic** = energy absorbed (heat is shown on left side of equation) (examples of endothermic phase changes: s→l, l→g, s→g)



Unit 5 Gas Laws:

103. **Pressure only effects gases**
104. Pressure is measured in **pascals**
105. * **STP** (**S**tandard **T**emperature and **P**ressure) on table A (273 K and 1 atm or 101.3 kPa and 0C)
106. Pressure and temperature have **direct relationship** (as pressure increases, temperature increases)
107. ***** $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ (if something is constant, you can cross it out of the formula)
108. ***Same volume** = same number of molecules
109. *****Conditions for ideal gas**: PLIGHT pressure low ideal gas high temperature
110. *Ideal gases move in random, constant, straight line motion
111. Ideal gases are separated by great distances compared to their size
112. Ideal gases have **no attractive forces**
113. Collisions of gas may result in a transfer of energy
114. ****Table H**, dotted line is the normal boiling of the substance
115. Gases have the **weaker IMF** than solids

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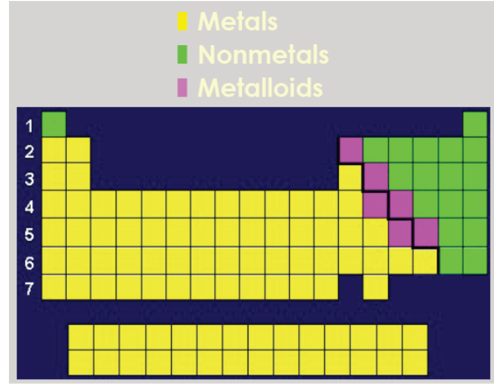

Unit 6 Periodic Table:

116. **Mendeleev** organized his PT by atomic mass
117. ****Modern Periodic Table**: Elements are arranged in order of **atomic number**
118. **Periods** are horizontal rows on the Periodic Table
119. **Groups** are vertical rows on the Periodic Table
120. **Liquids** on the periodic table = bromine and mercury
121. *****Metals** (left of the staircase) are good conductors of heat, malleable
122. Metals are malleable because of the nature of the bonds between the atoms
123. Metals have fewer valence electrons than nonmetals
124. Commonly **Metals react with nonmetals**
125. ****Nonmetals** to the right of the staircase
126. ***Metalloids** are on the staircase
127. **Transition metals** form colored solutions in aqueous ion form
128. ***Group 17 (halogen group)** – forms halide ions in solution
129. *******Noble gases** (group 18) are stable because of their stable electron configuration (8 valence electrons), they are not reactive
130. **Elements will gain or lose electrons to be like noble gases
131. ******MELPS- **M**etals **E**lectrons **L**ost form **P**ositive **S**maller ions
132. *Some groups will form the same **oxidation numbers** when gaining or losing electrons
133. ********Same group/family** = **similar chemical properties** because they have the same number of valence electrons
134. ******Attraction for electrons in a chemical bond = **electronegativity** (found on table T)
135. *****First Ionization energy and electronegativity increases left to right (across a period) and decreases top to bottom (down a group)
136. *******Atomic radius** (on table T) decreases left to right (across a period) and increases top to bottom (down a group)
137. *Atomic radius increases down a group because more energy shells are added

Unit 7 Naming, Formulas and Equations:

138. *******Ionics formulas**: metal comes first, nonmetal second. Criss cross charges to get formula. Example: Calcium phosphate

$Ca^{2+}\ PO_4^{3-} = Ca_3(PO_4)_2$

139. ******Naming Ionics** with table E: say metal and then ion from table E. Ammonium examples shown:

KCl Potassium chloride; MgS Magnesium sulfide
NH_4Cl Ammonium chloride; NH_4NO_3 Ammonium nitrate

140. **Roman numeral** represents the charge of the metal
141. ***Polyatomic ions** names on table E
142. **********Balancing**: number of each atom needs to be the same on both sides of the equation (use tallies). Use **coefficients** to balance

$C_3H_8 + 5\,O_2 \rightarrow 4\,H_2O + 3\,CO_2$

	LEFT SIDE	RIGHT SIDE
C	3	3 ✓
H	8	8 ✓
O	10	10 ✓

143. ****Mass, charge and energy** are conserved in a chemical reaction
144. ***Types of chemical reactions**= decomposition, single replacement, double replacement, synthesis
145. **Synthesis**: two or ore reactants combine to form one product

$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$

146. ***Decomposition**: one compound is broken down into two

$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$

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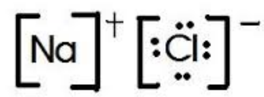
147. ****Single replacement** is when one element switches partners $Mg + 2HCl \rightarrow MgCl_2 + H_2$

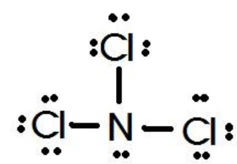
148. **Double Replacement:** think Do Si Do $AgNO_3 + KCl \rightarrow AgCl + KNO_3$

Unit 8 Bonding:

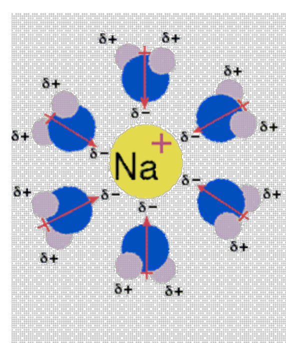
149. ********BARF**= bond BROKEN energy ABSORBED (endothermic), energy RELEASED bond FORMED (exothermic)
150. **Ionic compound** (metal and nonmetals or table E ions) and molecular compounds are all nonmetals with no table E ions
151. **Ionic compounds with a table E ion have covalent and ionic bonds
152. Ionic compounds- hard brittle solids with a high melting point, poor conductors as solids and good conductors in aq (higher concentration = better conductor)
153. **Metallic bonding**- metals only!
154. ****Ionic bonds** = bond between metal and nonmetal or ion from table E
155. Ionic bonds **transfer electrons** from the valence shell of one atom to the valence shell of another atom



156. *****Lewis dot diagrams** have brackets for ionics. Molecular compounds: make sure all elements have eight electrons (except H) and the total number of electrons in diagram = total number of valence electrons



157. *****Covalent bond** = nonmetals (molecular compound)
158. ***Nonpolar covalent bond** is a bond between two of the same elements
159. *****Multiple covalent bond**- double (four electrons shared/ 2 pairs), triple (six electrons shared/ 3 pairs)
160. Diatomic Oxygen molecule has double bond, diatomic nitrogen has triple
161. *******Electronegativity difference = polarity of bond** (higher the difference, the more polar the bond)



162. ****Molecule Polarity: SNAP = **symmetrical nonpolar, asymmetrical polar**
163. ******Stronger **Intermolecular forces** (IMF) = High boiling point
164. ***Example of IMF: **hydrogen bonding** (between H and F,O or N)
165. Water molecules (Oxygen is partially negative and hydrogen is partially positive)
166. Positive and negative charges attract, like charges repel

Unit 9 Stoichiometry:

167. Atomic mass on the Periodic Table is in **grams/ 1 mole**
168. *****GFM** (gram formula mass, formula mass or molar mass) is found by getting the mass of each element (multiplied by their subscript if they have one) and add all masses together
169. ********% Composition by mass**= mass of part / mass of whole (Table T)
170. % Composition can be calculated given a periodic table and chemical formula

H_2SO_4 = 98 g/mole

$2 \times H = 2 \times 1$ g/mole = 2 g/mole

$1 \times S = 1 \times 32$ g/mole = 32 g/mole

$4 \times O = 4 \times 16$ g/mole = 64 g/mole

sum = 98 g/mole

171. ******Mole** = mass/GFM (Table T)
172. ******Mole : mole ratio = coefficient : coefficient ratio**
173. *******Empirical formula**: simplified formula, molecular is normal not simplified formula
174. ****Conservation of mass**: mass of products = mass of the reactants

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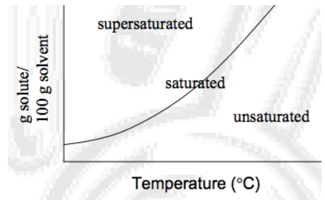
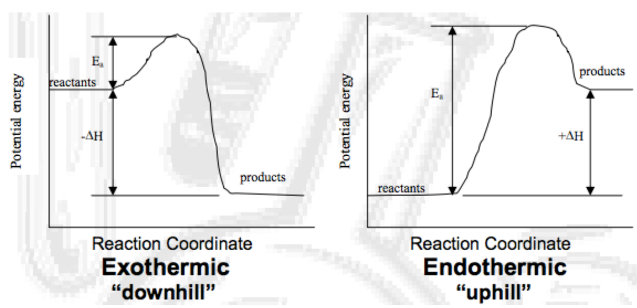
Unit 10 Solutions:



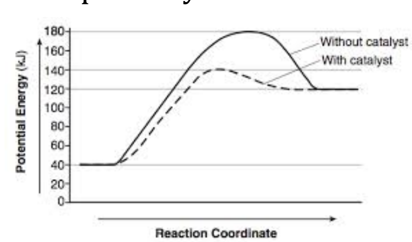
175. Solubility is dependent on temperature for solids (upward curve on table G)
176. ********Table G**- use when given a temperature and asking about a salt dissolved in water (on the line is saturated, below = unsaturated, above = super)
177. ***Likes dissolve likes** (example: polar mixes with polar substances)
178. ***Table F** (insoluble compounds will not fully dissolve in water and will be in solid phase, soluble compounds will dissolve in water and be in aq phase)
179. More ions more change in BP/FP
180. ***Add something to water (such as a salt) = **freezing point decreases** (more salt = lower the freezing point)
181. ***Add something to water (such as a salt) = **BP increases** (more salt = higher BP)
182. **Concentration units** m/L or PPM
183. *****PPM** = (grams of solute / grams of solution) x 1,000,000 [formula on table T]
184. *****Molarity = moles/L** formula on Table T

Unit 11 Kinetics:

185. *For a chemical reaction to occur you need **sufficient energy** and **proper orientation** for new bonds to form
186. Chemical reactions need **effective collisions**
187. *******Increase temperature**, increases reaction rate (more collisions)
188. ***Increase pressure**: increases reaction rate (more collisions)
189. **Lower concentration** slow reaction rate because of less collisions
190. **Greater surface area** = higher rate of reaction
191. **Potential energy diagrams** show energy of reaction and when a catalyst is add the energy of the reactants and products do not change
192. ***Activation energy** goes from the reactants to top of curve
193. **Reverse activation energy** goes from the products to top of curve



194. *Potential energy diagram starts high and ends low for an exothermic reaction
195. *******Heat of reaction** = potential energy of products – potential energy of the reactants
196. Heat of reaction (**Δ**H) is in the middle of a potential energy diagram
197. ******Table I** show heat of reaction (energy absorbed + **ΔH** energy released – **ΔH**
198. ******Catalyst** speeds up a reaction by providing a different reaction pathway that lowers the activation energy



199. ********Entropy (disorder)**: gases have the highest entropy, then liquid/aq and solids have the lowest
200. **Nature undergoes changes towards **higher entropy and lower energy**

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Equilibrium:

201. Forward reaction → Reverse reaction ← **Equilibrium ←→**

202. ***At equilibrium the **concentration of reactants and products remain constant** because reaction rates of the forward and reverse reaction are equal

203. *****Rate of forward reaction = rate of reverse reaction**

204. *Need a **closed system** to maintain equilibrium

205. Shift to right make more products, shift to left make more reactants

206. *****UP and AWAY, DOWN and TOWARDS

207. **Decrease pressure** shift to the side with **most number of moles**

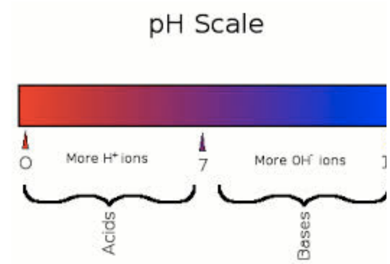
208. Increase temperature shift away from heat

209. Adding catalyst= no shift

210. Solution equilibrium: **Rate of dissolving equals rate of crystallization**

211. Dynamic phase equilibrium = two phase changes going on in a sealed flask

Unit 12 Acids and Bases:



212. ********Electrolytes**: acids, bases, salts

213. Electrolytes **conduct in aqueous** solutions not in solid phase

214. **Arrhenius theory**- describes the behavior of acids and bases

215. *******Arrhenius theory- **acids yield H+ ion** and **bases have an OH-**

216. *Acids and bases on **Table K/L** (know CH_3COOH is an acid, all other CH compounds are not electrolytes)

217. *******Alternate theory- **BAAD:** BASES ACCEPT **H+** ACIDS DONATE **H+**

218. Most acidic compounds = lowest pH

219. **Lower the pH (more acid) = **more hydronium ions**

220. ****Concentration of hydronium ions increases by **a factor of 10** when the pH goes down **by 1 pH unit.**

221. Moles of OH- and H+ are equal in a **neutral solution**

222. ********Neutralization: Acid + base → salt + water**

223. Any metals above H_2 on table J will react spontaneously with an acid

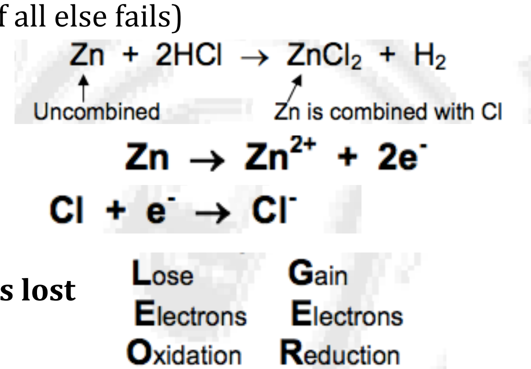
224. *******$M_aV_a = M_bV_b$ remember M is called molarity or concentration

225. ***Titration** determines the concentration of unknown acids or bases

226. **********Indicators**: range on the reference table (table M) is when the color is in-between two primary colors

Unit 13 Redox:

227. *******Oxidation number**: Single atom = 0, look on reference table or use algebra for atoms with multiple charges (can also use reverse criss cross if all else fails)



228. *****Redox reactions electrons are transferred** (charge changes)

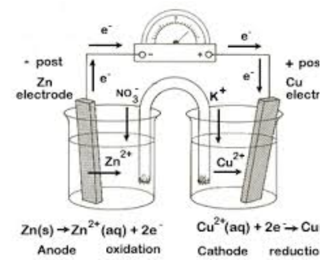
229. ******Oxidation: electrons are lost** (electrons on right), **Reduction: electrons are gained** (electron on left)
LEO the lion says GER

230. ***Half reactions** show charges and electrons lost

231. ***Balanced redox reaction has **same number of electrons lost and gained** (atoms and charges balanced)

232. *Lower on table J is less reactive

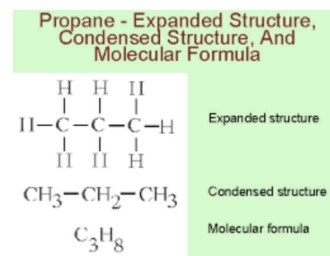
233. *More active metals are higher on table J
234. ***Higher on table J** more likely to be **oxidized** for metal side, lower more easily reduced
235. *If the oxidized element is higher on table j then the reaction will be spontaneous
236. *****AN OX RED CAT** (anode oxidation, reduction cathode)
237. ***Electrons flow from anode to cathode**
238. *Mass of anode will decrease because **solid metal is going into aqueous**
239. *****Voltaic cell**: chemical energy is converted to electrical energy spontaneously



240. Electrons flow through a wire for voltaic cells
241. *Ions move through the **salt bridge** for a voltaic cell
242. ****Electrolytic cell**: electrical energy to chemical energy
243. ***Battery (power source)** is needed for an electrolytic cell to provide electrical energy (electrolysis or electroplating)
244. Cathode in an electrolytic cell is what is electroplated (key or spoon)

<u>Unit 14 Organic:</u>

245. **Organic compounds** must contain at least one carbon and one hydrogen
246. **Hydrocarbon** contain only carbon and hydrogen
247. ***Saturated hydrocarbons** have all single bonds (2 electrons shared)
248. *******Unsaturated compounds** have double or triple bonds between carbons
249. Chemical formulas for organic: **empirical, structural, molecular**
250. **Structural formula** shows bonds
251. **Carbon (atomic number 6) can form **chains, rings and networks**
252. **Table p** tells number of carbons (remember carbon always makes 4 bonds!)



253. ***Homologous series** on table Q (alkanes, alkenes, alkynes)
254. ***Naming alkenes** = (use table Q), start from side with double bond and count number of carbons, put number to represent location of double bond
255. Count carbons from side closest to double bond or functional group
256. *******Isomers** have same number of C's and H's but are in a **different structural arrangement (same molecular formula)**
257. ***Different structure = different physical/chemical properties
258. *******Function groups: table R**
259. Different Functional groups have **different chemical properties**
260. *Drawing structure of **alcohol** includes putting an OH
261. **Halides** contains group 17 elements
262. **Organic Acid** structure has C with a double bonded O and OH
263. **Amines** contain N
264. **Type of organic reactions**: esterification, addition, saponification, polymerization, fermentation, etc.
265. ****Addition reaction**: binary compound (two atoms bonded together) gets added to double/triple bonded compound to reduce the number of bonds
266. **Saponification** = organic reaction used to make soap
267. **Polymerization** = adding same compound together to make a chain
268. ***Fermentation**: sugar + enzyme → ethanol (alcohol) + carbon dioxide

CHEM REVIEW: 281 ways to pass the Chemistry Regents

Test taking skills:

269. **Use the reference tables!**
270. Read paragraphs completely: information you need MIGHT be in there
271. Answer every question.
272. If you don't know the answer take the best guess you can make.
273. Refer to the reference table for any question you may have trouble on
274. Use the test to take the test. The answer or a hint may be in another question
275. Your first choice is usually your best choice (unless you read the question incorrectly the first time around)
276. Even if you think you know a chemical symbol, formula or charge….look it up an the reference table anyway (Table S and Periodic Table help a lot with this)
277. Skip a question if you are having a hard time and go back to it later.
278. **You have plenty of time to take the regents (3 hours). Take a 5 minute break after you are done with the test and then look over your answers.**
279. Eat a healthy meal the night before and for breakfast as well.
280. Get a good night's sleep. A tired mind is not as sharp and clear as a well-rested one.
281. **Relax! You have seen all this stuff before…it is somewhere in your brain!**