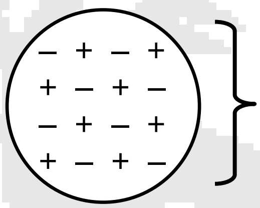
200 Tips to Pass the Chemistry Regents Exam

1. Don't run with scissors.
2. Don't play the horses.
3. ***Protons*** are positively charged (+); ***electrons*** are small and are negatively charged (–).
4. ***Neutrons*** have no charge (they can be thought of as a proton and an electron added together).
5. Protons & neutrons are in an atom's nucleus (***nucleons***).
6. Electrons are found in "clouds" (***orbitals***) around an atom's nucleus.
7. The ***mass number*** is equal to an atom's number of protons and neutrons added together.
8. The ***atomic number*** is equal to the number of protons in the nucleus of an atom.
9. The ***number of neutrons*** = mass number – atomic number.
10. ***Isotopes*** are atoms with equal numbers of protons, but differ in their neutron numbers.
11. ***Cations*** are *positive* (+) ions and form when a neutral atom *loses* electrons. They are *smaller* than their parent atom.
12. ***Anions*** are negative ions (–) and form when a neutral atom *gains* electrons. They are *larger* than their parent atom.
13. **Ernest *Rutherford's*** *gold foil experiment* showed that an atom is mostly empty space with a small, dense, positively-charged nucleus.
14. ***J.J. Thompson*** discovered the electron and developed the "plum-pudding" model of the atom.



Positive & negative charges spread throughout entire atom.

15. ***Dalton's*** model of the atom was a solid sphere of matter that was uniform throughout.
16. The ***Bohr Model*** of the atom placed electrons in "planet-like" orbits around the nucleus of an atom.
17. The current, ***wave-mechanical model*** of the atom has electrons in "clouds" (orbitals) around the nucleus.
18. USE THE REFERENCE TABLES!!! THEY ARE YOUR FRIENDS!!!
19. "***STP***" means "***S***tandard ***T***emperature and ***P***ressure." (273 Kelvin & 1 atm)
20. Electrons emit energy as light when they jump from higher energy levels back down to lower (***ground state***) energy levels. ***Bright line spectra*** are produced.
21. ***Elements*** are pure substances composed of only one kind of atom.
22. ***Binary compounds*** are substances made up of only *two* kinds of atoms. (examples: H_2O, NH_3, CO_2)
23. ***Diatomic molecules*** are elements that form two-atom molecules in their natural form at STP. Remember the phrase – "BrINClHOF" (Br_2, I_2, N_2, Cl_2, H_2, O_2, F_2)
24. Use this diagram to help determine the ***number of significant figures*** in a measured value:

If the decimal point is ***P**resent*, start counting digits from the ***P**acific* (left) side, starting with the first non-zero digit.

→ *1 2 3*

0.00310 (3 sig. figs.)

If the decimal point is ***A**bsent*, start counting digits from the ***A**tlantic* (right) side, starting with the first non-zero digit.

3 2 1 ←

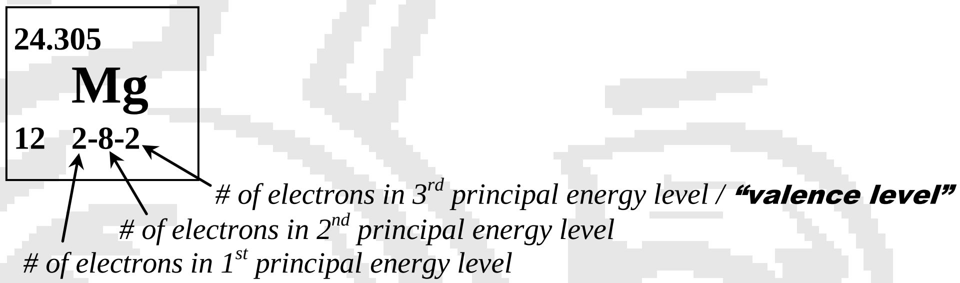
31400 (3 sig. figs.)

25. ***Solutions*** are the best examples of ***homogeneous mixtures***. (Air, salt water, soda, *etc.*)

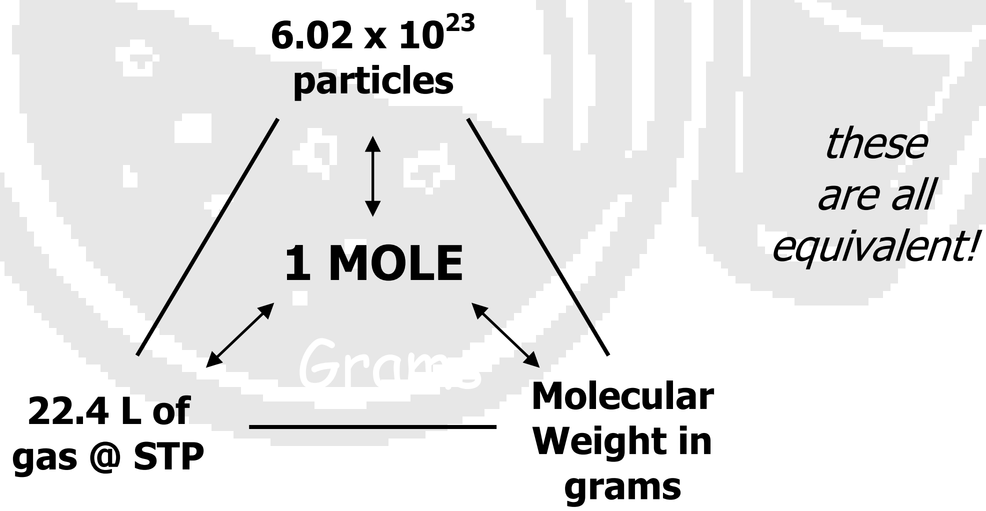
26. ***Heterogeneous mixtures*** have discernable components and are _not_ uniform throughout. (Chocolate-chip cookies, vegetable soup, soil, muddy water, etc.)
27. A ***solute*** is the substance being dissolved, while the ***solvent*** is the substance that dissolves the solute. (Water is the solvent in Kool-Aid, while sugar/flavoring is the solute.)
28. Isotopes are written in a number of ways: C-14 is also Carbon-14, and is also

mass number **(# p$^+$ plus # n^0)** → $^{14}_{6}C$ ← *atomic number* **(# p$^+$)**

29. The distribution of electrons in an atom is its ***electron configuration***.
30. Electron configurations are written in the bottom center of an element's box on the periodic table in your reference tables.



31. Use the ***mole triangle diagram*** below to help you solve conversions between moles, grams, numbers of molecules/atoms, and liters of gases at STP…



32. ***Orbital notation*** is a way of drawing the electron configuration of an atom.

↑↓	↑↓	↑	↑	
1s	**2s**		**2p**	

is carbon's orbital notation *(before electron promotion, of course!)*

32. Polyatomic ions (Table E) are groups of atoms with an overall charge other than zero ($\neq 0$).
 NO_3^{1-}, NH_4^{1+}, SO_4^{2-}, *etc.*
33. ***Coefficients*** are written in front of the formulas of reactants and products in chemical equations. They give us the ratios of reactants and products in a balanced chemical equation.
34. Chemical formulas are written so that the charges of cations and anions neutralize one another.
 Example: *calcium phosphate*:
 $Ca^{2+} PO_4^{3-} = Ca_3(PO_4)_2$ ***"criss-cross method"***
35. When naming binary ionic compounds, write the name of the positive ion (cation) first, followed by the name of the negative ion (anion) with the name ending in "-ide."
 Examples:

KCl	MgS
potassium chloride	***magnesium sulfide***

36. When naming compounds containing polyatomic ions, keep the name of the polyatomic ion the same as it is written in Table E.

Example:

NH_4Cl	NH_4NO_3
ammonium chloride	***ammonium nitrate***

37. ***Physical changes*** do not form new substances. They merely change the appearance of the original material, and <u>not</u> the identity. (*e.g.* the melting of ice)
38. ***Chemical changes*** result in the formation of new substances out of old ones. (*e.g.* the burning of hydrogen gas to produce water vapor)
39. ***Reactants*** are on the left side of the reaction arrow (the "ingredients") and ***products*** are on the right side (what results).
40. ***Endothermic reactions*** absorb heat. The energy value *is on the left side* of the reaction arrow in a forward reaction (energy is an <u>ingredient</u> necessary to make the reaction happen).
41. ***Exothermic reactions*** release energy and the *energy is a <u>product</u>* in the reaction (*on the right side* of the reaction arrow in a forward reaction).
42. *Only* coefficients can be changed when balancing chemical equations! DON'T TOUCH THE SUBSCRIPTS IN THE CHEMICAL FORMULAS!
43. ***Synthesis reactions*** occur when two or more reactants combine to form a single product.

    Example: $2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$
44. ***Decomposition reactions*** occur when a single reactant forms two or more products.

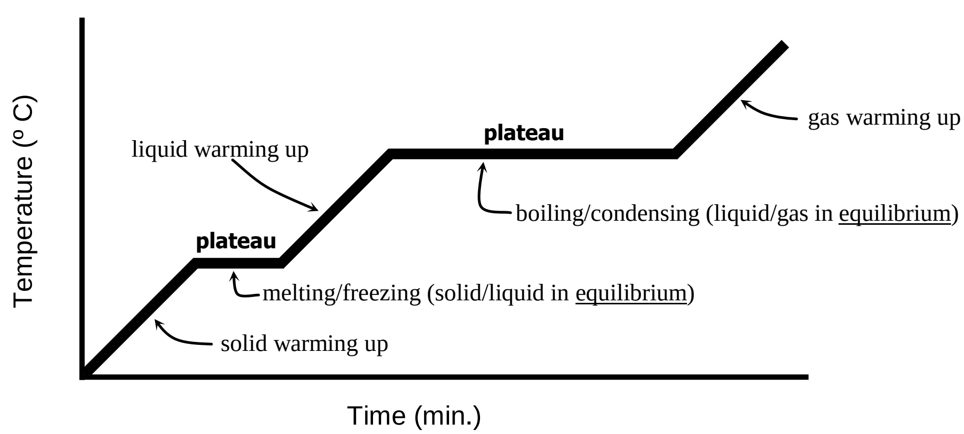
    Example: $CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$
45. ***Single replacement reactions*** occur when one element replaces another element in a compound.

    Example: $Mg + 2HCl \rightarrow MgCl_2 + H_2$
46. ***Double replacement reactions*** occur when two compounds react to form two new compounds.

    Example: $AgNO_3 + KCl \rightarrow AgCl + KNO_3$
47. The sum of the masses of the reactants in a chemical equation is always equal to the sum of the masses of the products. "***Law of Conservation of Mass***"
48. The gram formula mass of a substance is the sum of the atomic masses of all of the atoms in it.

    e.g. **H_2SO_4 = 98 g/mole**

    $$\left.\begin{array}{l} 2 \times H = 2 \times 1 \text{ g/mole} = 2 \text{ g/mole} \\ 1 \times S = 1 \times 32 \text{ g/mole} = 32 \text{ g/mole} \\ 4 \times O = 4 \times 16 \text{ g/mole} = 64 \text{ g/mole} \end{array}\right\} \rightarrow \text{sum} = 98 \text{ g/mole}$$
49. Know how to calculate the percentage composition of a compound. (Formula is on Table T)
50. ***6.02 x 10^23*** is called ***Avogadro's number*** and is the number of particles in ***1 mole*** of a substance.
51. The particles in a ***solid*** are rigidly held together.
52. ***Solids*** have a definite shape and volume.
53. ***Liquids*** have closely-spaced particles that easily slide past one another.
54. ***Liquids*** have no definite shape, but have a definite volume.
55. ***Gases*** have widely-spaced particles that are in random motion.
56. ***Gases*** are easily compressed and have no definite shape or volume.
57. Be able to read and interpret heating/cooling curves such as pictured below.



58. Substances that ***sublime*** turn from a solid directly into a gas. (common examples: CO_2 & I_2)
59. Degrees Kelvin = °C + 273
60. Use this formula to calculate heat absorbed/released by substances:

$q = mc\Delta t$

q = heat absorbed or released (Joules)
m = mass of substance in grams
c = specific heat capacity of substance (J/g•°C) … for water, the value is 4.18 J/g•ºC
Δt = temperature change in degrees Celsius

61. The heat absorbed or released when 1 gram of a substance changes between the solid and liquid phases is called the substance's ***heat of fusion***. (334 J/g for water)
62. The heat absorbed or released when 1 gram of a substance changes between the liquid and gaseous phases is called the substance's ***heat of vaporization***. (2260 J/g for water)
63. As the ***pressure*** on a gas increases, the ***volume*** decreases proportionally.
64. As the ***temperature*** on a gas increases, ***pressure*** increases proportionally.
65. As the ***temperature*** of a gas increases, ***volume*** increases proportionally.
66. *Always use Kelvins* for temperature when using the ***Combined Gas Law***:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

67. ***Real gas*** particles have volume and are attracted to one another, and thus do not always behave like ***ideal gases***.
68. Real gases behave more like ideal gases at *low pressures and high temperatures.*
69. ***Distillation*** separates mixtures with different boiling points.
70. ***Filtration*** separates mixtures of solids and liquids.
71. ***Chromatography*** can also be used to separate mixtures of liquids and/or mixtures of gases, based on varying mobilities of the different *phases* present.
72. ***The Periodic Law*** states that the properties of elements are periodic functions of their *atomic numbers.*
73. ***Periods*** are horizontal rows on the Periodic Table: properties slowly change across a period.
74. ***Groups/families*** are vertical columns on the Periodic Table: elements share similar properties.
75. ***Metals*** are found left of the "staircase" on the Periodic Table, ***nonmetals*** are above it, and ***metalloids*** live on it.
76. It would be real handy to know this chart! —

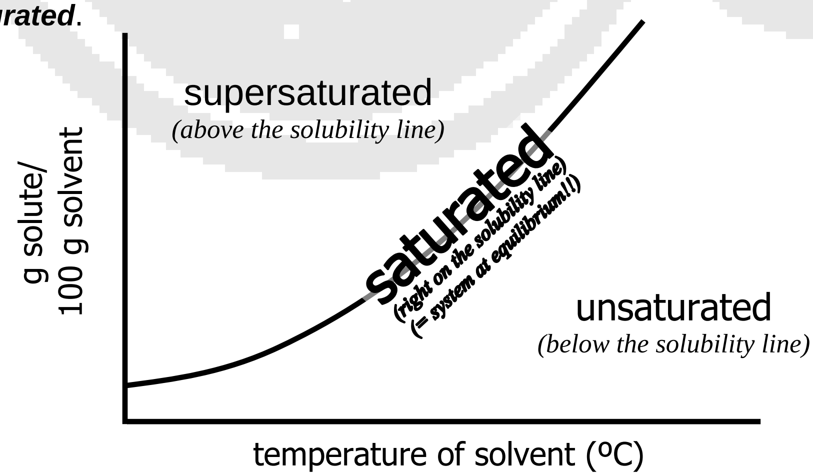
Metals	Malleable	Ductile	Lustrous	Good conductors of heat & electricity	Low ionization energy and electronegativity values	Tend to form (+) ions
Nonmetals	Brittle when solid	Mostly gases at STP	Dull	Good insulators	High ionization energy and electronegativity values	Tend to form (–) ions

77. ***Noble gases*** (Group 18) are inert and stable due to the fact that their valence levels of electrons are completely filled.
78. ***Ionization energy*** increases as you go up and to the right on the Periodic Table.
79. ***Atomic radii*** *decrease* left to right across a period due to increasing nuclear charge.
80. ***Atomic radii*** *increase* as you go down a group, due to increasing number of energy levels, and electron shielding effect.
81. ***Electronegativity*** is a measure of an element's attraction for electrons while in a chemical bond.
82. Electronegativity *increases* as you go up and to the right on the Periodic Table.
83. The elements in Group 1 are the ***alkali metals***.
84. The elements in Group 2 are the ***alkaline earth metals***.
85. The elements in Group 17 are the ***halogens***.
86. The elements in Group 18 are the ***noble gases***.
87. Use ***Table S*** to compare and look up the properties of specific elements.
88. Energy is *released* when a chemical bond *forms*. The more energy that is released, the more stable is the bond that is forming.
89. The last digit of an element's group number is equal to its ***number of valence electrons***.
90. Draw one dot for each valence electron when drawing an element's or ion's ***Lewis diagram***.
91. The ***kernel*** of an atom includes everything in an atom *except* the atom's valence electrons.

92. Metallic bonds can be thought of as a crystalline lattice of kernels surrounded by a "sea" of mobile valence electrons.
93. Atoms are most stable when they have 8 valence electrons (an ***octet***) and tend to form ions to obtain such a configuration of electrons (the "**Octet Rule**").
94. ***Covalent bonds*** form when two atoms ***share*** a pair of electrons.
95. ***Ionic bonds*** form when one atom ***transfers*** an electron(s) to another atom when forming a bond with it.
96. ***Nonpolar covalent bonds*** form when two atoms of the *same element* bond together.
97. ***Polar covalent bonds*** form when the electronegativity difference between two bonding atoms is between ≈ 0.3 and 1.7.
98. ***Ionic bonds*** form when the electronegativity difference between two bonding atoms is *greater than* 1.7 (typically, a metal bonding to a nonmetal or polyatomic ion).
99. Substances containing covalent bonds are called ***molecular substances***.
100. Substances containing ionic bonds are called ***ionic compounds***.
101. It would be real handy to know this chart! —

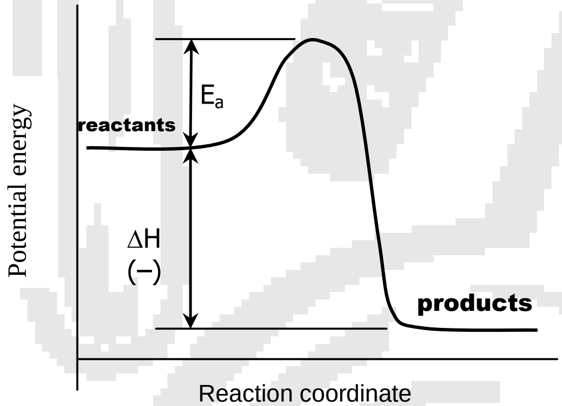
Substance Type	Properties
Ionic	Hard High melting and boiling points Conduct electricity when molten or when aqueous
Covalent (Molecular)	Soft Low melting and boiling points Do not conduct electricity (insulators)

102. ***Hydrogen bonds*** form when hydrogen bonds to the elements N, O, or F ("FON bonds") and gives the compound unusually high melting and boiling points.
103. Use Table F to predict the solubilites (soluble/insoluble, that is no precipitate forms *vs.* a precipitate forms) of compounds.
104. Remember substances tend to be soluble in solvents with similar properties….
"Likes Dissolve Likes"
105. As temperature increases, solubility increases for most solids.
106. At low temperatures and high pressures solubility *increases* for most gases.
107. Use *Table G* to determine whether a particular solution is ***saturated***, ***unsaturated***, or ***supersaturated***.

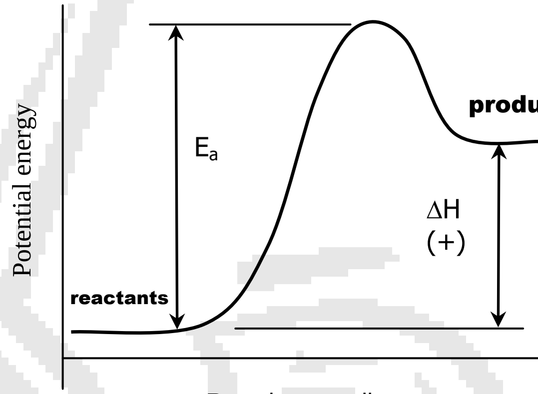


108. ***Molarity*** is one way to measure the *concentration* of a solution. Molarity is equal to the number of moles of solute divided by the number of liters of solution. ($M = {}^{mol}/_{liters}$)
109. ***Percent by mass*** = mass of the part / mass of the whole x 100%
110. ***Parts per million (ppm)*** = grams of solute / grams of solution x 1,000,000
111. Non-volatile solutes raise the boiling points and lower the melting points of pure solvents.
112. Liquids ***boil*** when their vapor pressure is equal to the atmospheric pressure.
113. The ***normal boiling point*** of a substance is the temperature at which it boils at 1 atm of pressure. (Take note of Table H)
114. Covalently bonded substances tend to react more slowly than ionic compounds.

115. Increasing the concentration of reactants will increase reaction rate.
116. Increasing the surface areas of the reactants will increase reaction rate.
117. Increasing temperature of the reactants increases reaction rate.
118. Increasing the pressure on gases increases reaction rate.
119. ***Catalysts*** speed up reactions by lowering their ***activation energies***. They are not changed themselves and can be reused many times over.
120. Be able to recognize and read ***potential energy diagrams***.



Reaction coordinate
exothermic
"downhill"



Reaction coordinate
endothermic
"uphill"

121. ΔH is (+) for endothermic reactions and is (–) for exothermic reactions.
122. The rates of the forward and reverse reactions are equal at equilibrium.
123. ***Adding*** any reactant or product to a system at equilibrium will shift the equilibrium away from the added substance.
124. ***Removing*** any reactant or product from a system at equilibrium will shift the equilibrium point toward that removed substance.
125. An ***increase in temperature*** shifts an equilibrium system in the ***endothermic direction***.
126. A ***decrease in temperature*** shifts an equilibrium system in the ***exothermic direction***.
127. ***Increasing the pressure*** on a gaseous equilibrium will shift the equilibrium point toward the side with ***fewer moles of gas***.
128. ***Decreasing the pressure*** on a gaseous equilibrium will shift the equilibrium point toward the side with ***more moles of gas***.
129. ***Catalysts*** have ***no effect*** on an ***equilibrium system***, other than to just establish itself quicker.
130. ***Enthalpy (H)*** is the heat energy gained or lost in a reaction.
131. ***Entropy (S)*** is high in a highly unorganized system, such as a gas, a messy room, etc.
132. For the hypothetical reaction $\mathbf{wA + xB \rightarrow yC + zD}$, $K_{eq} = \dfrac{[C]^y[D]^z}{[A]^w[B]^x}$
133. ***Oxidation*** is the ***loss of electrons*** by an atom or ion. The oxidation number *increases* as a result. The electrons are on the *right side* of the reaction arrow.

$$\mathbf{Zn \rightarrow Zn^{2+} + 2e^-}$$

134. ***Reduction*** is the ***gain of electrons*** by an atom or ion. The oxidation number *decreases* (is **reduced**!) as a result. The electrons are on the *left* side of the reaction arrow.

$$\mathbf{Cl + e^- \rightarrow Cl^-}$$

135. Redox reactions ***always*** involve the exchange of ***electrons***.
136. Remember…. "LEO says GER!"

Lose	**G**ain
Electrons	**E**lectrons
Oxidation	**R**eduction

137. One way to ***identify redox reactions*** is by seeking an uncombined element on one side of a reaction that is in a compound on the other side.

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$$

Zn is uncombined *Zn is combined with Cl*

138. ***Oxidizing agents*** are what *get reduced* in a redox reaction; ***reducing agents*** are what *get oxidized* in a redox reaction.
139. ***Electrochemical cells*** produce electricity with a *spontaneous* redox reaction.
140. The *left electrode* is usually the site of *oxidation* in an electrochemical cell diagram.
141. A helpful mnemonic is "I have ***AN OX*** and a ***RED CAT***." :
In electrochemical cells, the ***AN***ode gets ***OX***idized and ***RED***uction occurs at the ***CAT***hode.
142. ***Electrolytic cells*** use an applied electrical current to force a ***nonspontaneous*** redox reaction to take place.
143. Electrolytic cells are usually used for metal plating of objects.
144. ***Acids*** and ***bases*** are both ***good electrolytes****;* their solutions conduct electricity well.
145. Weak acids taste *sour*.
146. Weak bases taste *bitter*.
147. Acids and bases turn ***indicators*** different colors; see ***Table M*** for details and know how to use it!
148. Acids have a pH < 7; bases have a pH > 7.
149. A pH of 7.0 means exactly neutral: the solution contains equal amount of H^+ and OH^- ions.
150. ***Tables K & L*** list names and formulas of common acids and bases asked about on the Regents.
151. The metals <u>above</u> H_2 on ***Table J*** will react with acids to make produce H_2 gas (bubbles).
152. ***Arrhenius*** says:
"Acids give off H^+ or H_3O^+ ions in solution."
"Bases give off OH^- ions in solution."
153. ***Brønsted*** says:
"Acids *donate* protons to a proton acceptor."
"Bases *accept* protons from a proton donor."
154. Acids and bases react in ***neutralization*** reactions to make ***water*** and an ***ionic salt***.
155. ***Titrations*** are controlled neutralization reactions used to find the concentration of an acid or base sample. Note the formula for it on Table T.
156. ALL organic compounds contain the element ***carbon***.
157. *Each* **carbon atom** ALWAYS seeks to make ***four bonds*** in molecules (no more, no less).
158. ***Saturated*** hydrocarbons have only *single* bonds between the carbon atoms within them (alkanes).
159. ***Unsaturated*** hydrocarbons have at least one *double* or *triple* bond between the carbons atoms in them (alkenes & alkynes, respectively).
160. ***Hydrocarbons*** contain ONLY the elements hydrogen and carbon.
161. The ***homologous series*** of hydrocarbons' formulas are on ***Reference Table Q***.
162. The ***functional groups*** of organic molecules are listed on ***Reference Table R***.
163. ***Structural isomers*** of organic compounds have *different* structural formulas but the *same* molecular formula (same chemical formulas, but different physical structure).
164. Number the parent carbon chain in an organic molecule from the end closest to the alkyl group(s) (the branches hanging off the root chain).
165. ***Combustion reactions*** occur when a hydrocarbon reacts with oxygen to make CO_2 and H_2O.
166. ***Organic substitution reactions*** occur when an alkane and a halogen (Group 17) reacts such that one or more hydrogen atoms on the alkane are replaced with a halogen.
167. ***Organic addition reactions*** occur when an alkene or alkyne combine with a halogen to make one product (a "*halocarbon*").
168. ***Esterification*** occurs when an organic acid and an alcohol react to make water and an ***ester***.
169. ***Saponification*** occurs when an ester reacts with a base to make alcohol and a ***soap***.
170. ***Fermentation*** reactions occur when yeast catalyzes a sugar ($C_6H_{12}O_6$) to make carbon dioxide and an alcohol (typically ethanol).
171. ***Polymers*** are long carbon chains built up from repeating units called ***monomers***.
172. Polymers form by ***polymerization*** reactions.

173. ***Addition polymerization*** occurs when unsaturated monomers join into a long polymer chain.

$$n\ C_2H_2 \rightarrow (C_2H_2)_n$$

174. ***Condensation polymerization*** occurs when monomers join to form a polymer *by removing water.* Water is a product!
175. ***Natural polymers*** include starch, cellulose, and proteins.
176. ***Synthetic polymers*** include nylon, rayon, polyester, and plastics.
177. Unstable atoms that are radioactive are called ***radioisotopes***. (***Table N***)
178. Radioisotopes can decay by giving off any of the particles/emissions listed in ***Table O***.
179. ***Alpha particles*** (see Table O) are positively charged (+).
Beta particles (see Table O) are negatively charged (–).
180. For nuclear equations, the sum of the mass numbers and atomic numbers (charge numbers) must be equal on both sides of the reaction arrow:
for example:

$$\overbrace{{}^{14}_{7}\mathrm{N} + {}^{4}_{2}\mathrm{He}}^{\textit{mass sum = 18}}_{\textit{charge sum = 9}} \longrightarrow \overbrace{{}^{17}_{8}\mathrm{O} + {}^{1}_{1}\mathrm{H}}^{\textit{mass sum = 18}}_{\textit{charge sum = 9}}$$

181. ***Fission reactions*** split heavy nuclei into smaller ones and release energy.

$${}^{1}_{0}\mathrm{n} + {}^{235}_{92}\mathrm{U} \longrightarrow {}^{139}_{56}\mathrm{Ba} + {}^{94}_{36}\mathrm{Kr} + 3\,{}^{1}_{0}\mathrm{n} + \text{lots o' energy}$$

182. ***Fusion reactions*** occur when light nuclei combine to form a heavy nucleus and *a lot of energy*.

$${}^{2}_{1}\mathrm{H} + {}^{2}_{1}\mathrm{H} \longrightarrow {}^{4}_{2}\mathrm{He} + \text{LOTS O' ENERGY!}$$

183. The ***half life*** of a radioisotope is the length of time it takes for exactly one half of the atoms in a sample to radioactively decay into its corresponding products (and the other half remains behind unchanged). (Table N)
184. C-14 is used to determine the ages of organic material up to 23,000 years old.
185. U-238 is used to determine the ages of rocks.
186. I-131 is used to treat thyroid disorders.
187. Co-60 is used to treat cancer tumors.
188. Radiation can be used to kill bacteria on foods to slow the spoilage process.
189. Disposal of radioactive waste is a problem associated with nuclear reactors. Where do you store stuff so incredibly toxic and deadly?
190. USE THE REFERENCE TABLES!!!
191. Be sure to answer every question, even if you must (take an educated) guess. Some chance of getting it right is better than none at all, and you're guaranteed to be wrong if you write nothing!
192. You have three hours to take the test, so take your time. But remember to pace yourself!
193. Try substituting different words for words that seem confusing. Sometimes this helps the question make more sense. (ex.: substitute the word "false" for "not true")
194. Consider on every question if the answer is in the reference tables or if the reference tables could help you. Very frequently the tables will help, and oftentimes will help *tremendously*!
195. Your first answer is usually your best one. Only change an answer if you are certain you are wrong, or if you find an obvious mistake when checking your work.
196. Even if you think you know a formula, look it up. It's free, and that's why you've got the reference tables in the first place: so you don't have to *memorize* a lot of stuff! Most are on Table T (the last page).
197. Skip a question if it is giving you a hard time. Go back to it later. Something else in the test may help you answer the harder problem.
198. Eat a healthy meal the night before and for breakfast as well. Your brain needs food to run!
199. Get a good night's sleep. A tired mind is not as sharp and clear as a well-rested one.
200. *Relax* — you've seen all this stuff before!