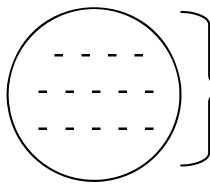
194 Things to Know for the Chemistry Regents Exam

1. *Protons* are positively charged (+).
2. *Neutrons* have no charge.
3. *Electrons* are small and are negatively charged (-).
4. Protons & neutrons are in an atom's nucleus.
5. Electrons are found in "clouds" (*orbitals*) around an atom's nucleus that are regions of most probable electron location.
6. The *mass number* is equal to an atom's number of protons and neutrons added together.
7. The *atomic number* is equal to the number of protons in the nucleus of an atom.
8. The *number of neutrons* = mass number – atomic number.
9. *Isotopes* are atoms of the same element, but differ in the number of neutrons.
10. Positive (+) ions form when an atom loses electrons. They are smaller than their parent atom.
11. Negative ions form when an atom gains electrons. They are larger than their parent atom.
12. Ernest *Rutherford's gold foil experiment* showed that an atom is mostly empty space with a small, dense, positively-charged nucleus.
13. *J.J. Thompson* discovered the electron and developed the "plum-pudding" model of the atom.



The whole atom was positive & negative particles were spread throughout entire atom.

14. Dalton's model of the atom was a solid sphere of matter that was uniform throughout.
15. The Bohr Model of the atom placed electrons in energy levels around the nucleus of an atom.
16. The current, wave-mechanical model of the atom has electrons in "clouds" (orbitals) around the nucleus.
17. USE THE REFERENCE TABLES!!!
18. "STP" means "Standard Temperature and Pressure." (273 Kelvin & 1 atm) – Table A
19. Electrons emit energy (often as light) when they return from higher energy levels to lower (ground state) energy levels. Bright line spectra are produced.
20. Elements are pure substances composed of only one kind of atom.
21. Binary compounds are substances made up of only two kinds of atoms. (examples: KCl, Fe_2O_3, H_2O, NH_3, CO_2)
22. Diatomic molecules are elements that form two atom molecules in their natural form at STP. Remember the phrase – "HONClBrIF twins" (H_2, O_2, N_2, Cl_2, Br_2, I_2, F_2,)

23. Use this diagram to help determine the *number of significant figures* in a measured value…

Pacific  Atlantic

If the decimal point is *present*, start counting digits from the *Pacific* (left) side, start with the first non-zero digit.

$\longrightarrow$ 1 2 3
0.00310 (3 sig. figs.)

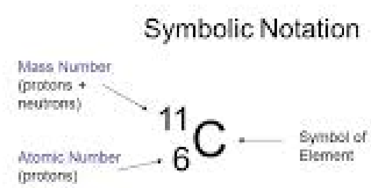
If the decimal point is absent, start counting digits from the Atlantic (right) side, start with the 1st non-zero digit.

3 2 1 $\longleftarrow$
31,400 (3 sig. figs.)

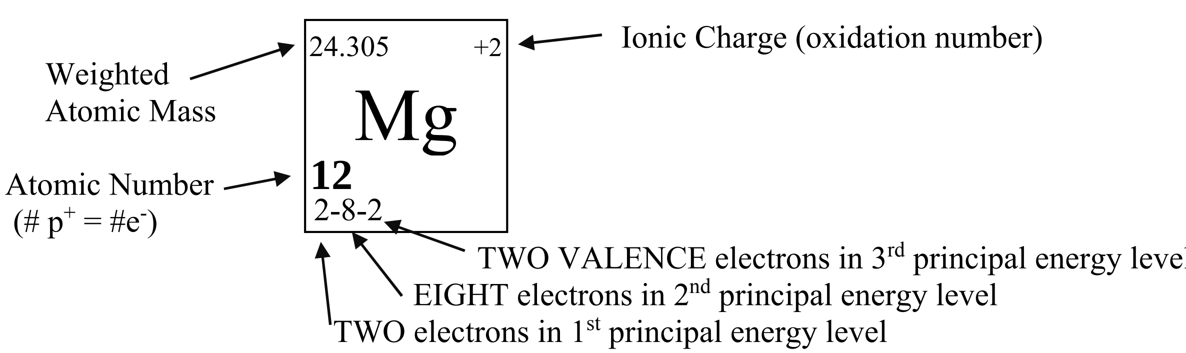
24. *Solutions* are the best examples of *homogeneous mixtures*. (Air, salt water, etc.)
25. *Heterogeneous mixtures* have discernable components and *are not* uniform throughout. (Chocolate-chip cookies, vegetable soup, soil, muddy water, etc.)
26. A *solute* is the substance being dissolved, while the *solvent* is the substance that dissolves the solute. (Water is the solvent in Kool-Aid, while sugar is the solute.)
27. Isotopes are written in a number of ways: C-14 is also carbon-14.
    Here are 4 different isotopes of carbon, with their proper symbols.

$^{11}_{6}C$ $^{12}_{6}C$ $^{13}_{6}C$ $^{14}_{6}C$

28. The distribution of electrons in an atom is its *electron configuration*.



29. Electron configurations are written in the bottom center of an element's box on the periodic table.



30. For mole conversions;
1 mole = molar mass (sometimes for ionic compounds this is called the gram formula mass)
1 mole of a GAS at STP = 22.4 Liters
1 mole = 6.02×10^{23} particles

31. Coefficients are the big numbers written in front of the formulas of reactants and products in chemical equations. They give us the MOLE RATIOs of reactants and products in balanced chemical equation.

32. The coefficients in a balanced equation indicate the number of moles of each substance in the reaction.
33. Polyatomic ions (Table E) are groups of atoms with an overall charge, they act as a PARTICLE.
NO_3^{1-}, NH_4^{1+}, SO_4^{2-}, etc.

34. Chemical formulas are written so that the charges of cations + anions balance, or neutralize one another.
Use the "criss-cross" method of making formulas. Ex: *calcium phosphate*:

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

35. When naming binary ionic compounds, write the name of the cation first, followed by the name of the negative anion, changing the atomic name to end with "-ide."
Ex: KCl is *Potassium chloride* MgS is *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. Ex: NH_4Cl is *Ammonium chloride* NH_4NO_3 is *Ammonium nitrate*

37. *Physical changes* do not form new substances. They only change the appearance of the original substance, they are PHASE CHANGES.
38. *Chemical changes* result in the formation of different substances. They involve bonding/unbonding. (Example: The burning of hydrogen gas in oxygen to produce water vapor.)

39. *Reactants* are on the left side of the reaction arrow and *products* are on the right.
40. *Endothermic reactions* absorb heat. The energy value *is on the REACTANT side* of the reaction arrow in a forward reaction.
41. *Exothermic reactions* release energy and the *energy is a PRODUCT* in the reaction.

42. *Only* coefficients can be changed when balancing chemical equations; NOT subscripts! You can not change the formula of a compound, it is what it is.
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 solution.
Example: $Mg_{(G)} + 2HCl_{(AQ)} \rightarrow MgCl_{2(AQ)} + H_{2(G)}$

46. *Double replacement reactions* occur when two compounds react to form two different compounds, a SOLID PRECIPIATATE and a new aqueous solution.

Example: $AgNO_{3(AQ)} + KCl_{(AQ)} \rightarrow AgCl_{(S)} + KNO_{3(AQ)}$

47. In reactions, the mass of the reactants equals the mass of the products. "*Law of Conservation of Mass*."

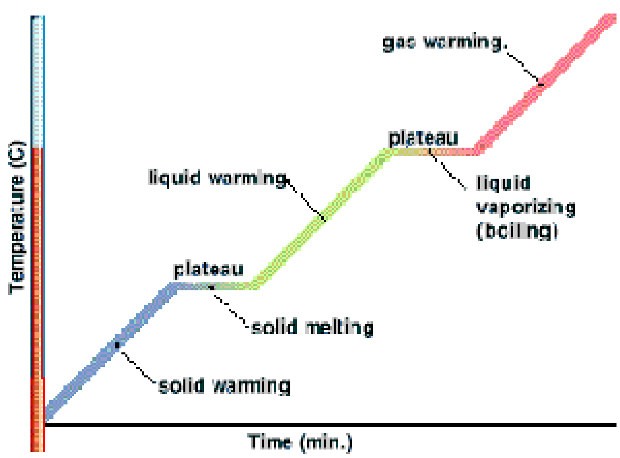
48. The gram-formula mass of a substance is the sum of the atomic masses of all of the atoms in it.

H_2SO_4
2 x H = 2 x 1 g/mole = 2 g/mole
1 x S = 1 x 32 g/mole = 32 g/mole
4 x O = 4 x 16 g/mole = 64 g/mole
sum = 98 g/mole

49. Know how to calculate the percentage composition by mass of a compound.
50. Be able to calculate percent error using the formula

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 as pictured below.



Heating/cooling curves indicate changes in potential energy and kinetic energy

58. Substances that *sublime* turn from a solid to gas. (CO_2 & I_2)
Deposition is the opposite; gas to solid. (frost)
59. Converting from Kelvin to Celcius or back, use this formula $K = °C + 273$

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 constant (J/g·°K) … for water it's 4.18 J/g·K
ΔT = temperature change in Kelvin

THIS IS IMPORTANT: the ΔT C = ΔT K
Think: what is the ΔT C between freezing and boiling for water? That would be 100 – 0 = 100 C.
What is the ΔT K between freezing and boiling for water? That would be 373 – 273 = 100 K.

The temperatures are different, but the change is the same! Like running a race.
If you measure it in feet or yards, the numbers are different but the distance is the same.

61. The heat absorbed or released when 1 gram of a substance changes between the solid and liquid phases is the substance's *heat of fusion*. (H_F = 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 the substance's *heat of vaporization*. (H_V = 2260 J/g for water)

63. Volume and Pressue are inversely proportional: as the *volume* of a gas ↓, the pressure ↑ (constant T)
64. As *temperature* of a gas ↑ *pressure* ↑. (or both ↓) Temp + Pressure are directly proportional (constant V)

65. As the *temperature* of a gas ↓ *volume* ↓ (or both ↑). Temp + volume directly proportional (constant P)
66. *Always use Kelvin* 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 + are attracted to one another, and thus do not *ideal gases*.
68. Real gases behave most like an ideal gas at *low pressures and high temperatures.*
(less collisions and less likelihood of sticking together means more likely to STAY A GAS)

Separating Mixtures apart is NOT chemical, it is a physical thing. You must take advantage of differences in physical properties.
69. *Distillation* separates mixtures based on DIFFERENT boiling points.
70. *Filtration* separates mixtures of solids and liquids based upon DIFFERENT particle sizes.
71. *Chromatography* can also be used to separate mixtures of liquids and mixtures of gases based on DIFFERENT molecular polarity, or DIFFERENT particle sizes.
72. Magnets separate 2 substances because of that DIFFERENCE in being magnetic or not.

73. *The Periodic Law* states that the properties of elements are periodic functions of their *atomic numbers.* That the properties of elements is orderly, if the elements are arranged in order of increasing atomic number, in the weird shape that the periodic table must take to make this true.

74. *Periods* are horizontal rows on the Periodic Table. The only similarity is same number of orbitals.
75. *Groups* are vertical columns on the Periodic Table. These have many chemical similarities because they all share the same number of VALENCE electrons, where 99% of all bonding takes place.
76. *Metals* are found left of the "staircase" on the Periodic Table, *nonmetals* are to the right. The *metalloids* touch it. BUT NOT HYDROGEN, and not AlPo – they touch the staircase but they are not metalloids

77. *Noble gases* (Group 18) are mostly inert + stable because their valence electrons orbitals are full.
78. FIRST *Ionization energy* increases as you go up and to the right on the Periodic Table.
79. The *Atomic radius of atoms will decrease* left to right across a period due to increasing net nuclear charge (more protons in the nucleus mean greater inward attraction on the electrons).
80. *Atomic radii increases* as you go down a group because you ADD ORBITALS going down the group.

81. *Electronegativity* is a measure of an element's atom's attraction for electrons in a bond.
82. Electronegativity *increases* as you go up and to the right on the Periodic Table.
83. The elements in Group 18 are the *noble gases*.
84. Use *Table S* to compare and look up the physical properties of specific elements.

85. Be able to determine Periodic trends using periodic table and the element data in Table S.
86. The number of electron shells in an atom of an element is the number of the period where the element is found on the Periodic Table.
87. When bonds form, Energy is *released.* Ex: $N + N \rightarrow N_2$ + energy
88. Energy is required to break a chemical bond. Ex: Cl_2 + energy $\rightarrow Cl + Cl$

89. Electron configurations of the atoms 72-89 are shown funky. Every atom fits a 2-8- in the first 2 orbitals.

90. Draw one dot for each valence electron when drawing an element's *Lewis dot diagram*. (AKA electron dot) Atoms get dots and NO brackets. Ions get Brackets… metal cations get NO DOTS, anions get 8 dots.

91. Be able to draw electron-dot diagrams for atoms, ions, and molecules.
92. Metallic bonds can be thought of as an organized pattern of positive ions 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. Noble gases are the "most stable" of the atoms, hence they hardly make any bonds or comounds ever.
94. A *Covalent bond* forms when two atoms *share* a pair of electrons.
95. Ionic *bonds* form when metal atoms *transfer* one or more electrons to another nonmetal atom when bonding. These cations and anions must form simultaneously. You can't have a jar of aluminum cations alone.

96. *Nonpolar covalent bonds* form when two atoms of with the same ELECTRONEGATIVITY VALUE bond.
97. Polar covalent bonds form when there is an electronegativity difference between two bonding atoms.

98. Ionic *bonds* form when metals and nonmetals form ions, and are attracted together (wildly) strongly.
99. Substances containing covalent bonds are called *molecular substances*.
100. Substances containing ionic bonds are called *ionic compounds*.
101. Know this:

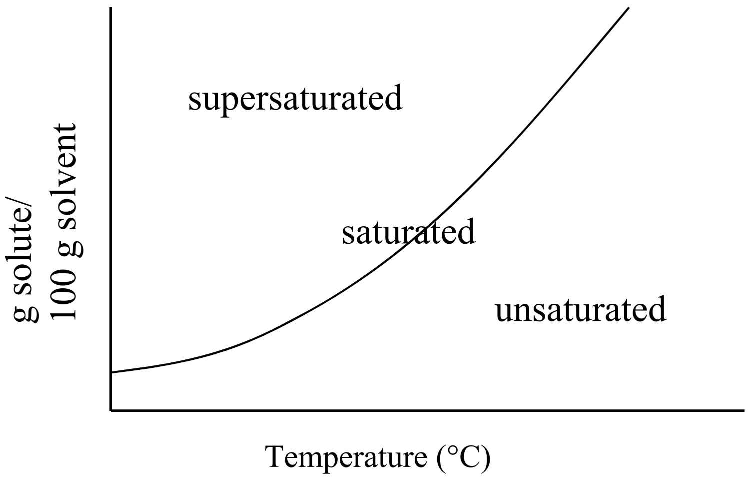
Ionic Compounds: are Hard, have High melting + boiling points, Conduct electricity when molten or AQ

Covalent (Molecular) Compounds: are Soft, have Low melting + boiling points, and
do not conduct electricity (they are insulators)

102. *Hydrogen bonds* form when hydrogen atoms bond in polar molecules, and gives the compound strong intermolecular attractions, resulting in unusually high melting point and boiling points.

103. Use Table F to predict the solubilites of compounds in WATER. Insoluble means the amount of the compound that will dissolve is nearly zero (but not actually zero).
104. Remember substances tend to be soluble in solvents with similar properties "Like Dissolves Like"
105. As temperature increases, solubility increases for most solids.
106. At low temperatures and high pressures solubility *increases* for gases.

107. Use *Table G* to determine whether a solution is *saturated*, *unsaturated*, or *supersaturated*.



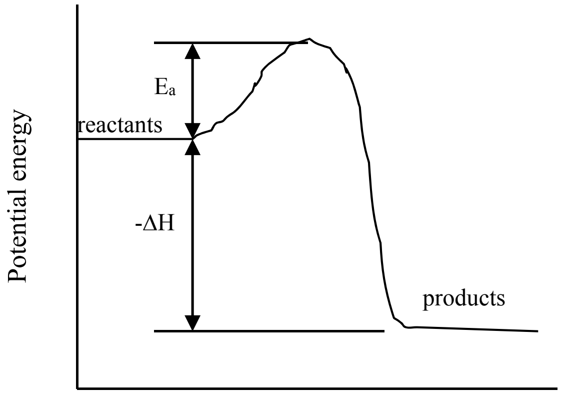
108. *Molarity* is a way to measure the *concentration* of a solution. Molarity is equal to the number of moles of solute divided by the number of liters of solution. The formula is on the back of the reference tables.

109. *Percent Composition 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

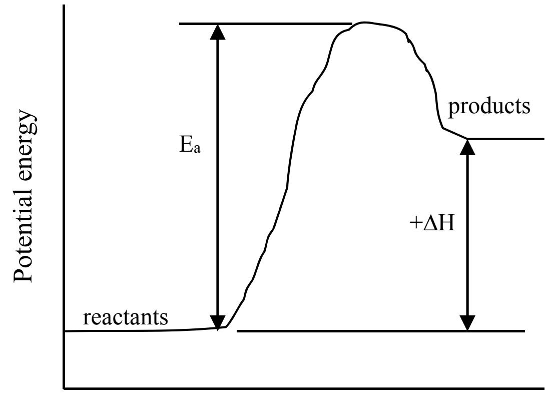
111. Solutes raise the boiling points <u>and</u> lower the melting points of solvents.
112. Liquids *boil* when their vapor pressure is equal to the external pressure on the liquid.
113. The *normal boiling point* of a substance is the temperature at which it boils at 1 atm of pressure. (Table H)

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 solid reactants will increase reaction rate.
117. Increasing the pressure on gases increases reaction rate.
118. *Catalysts* speed up reactions by providing an alternate reaction pathway having a lower *activation energy*. They are not changed themselves and can be reused many times over.
119. Increasing temperature increases reaction rate.

120. Be able to recognize and read *potential energy diagrams*.



Reaction Coordinate
Exothermic
(products lower than reactants)



Reaction Coordinate
Endothermic
(products higher than reactants)

121. ΔH is (+) for endothermic reactions and ΔH 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*. It just establishes itself quicker.

130. Natural systems tend to change to achieve a state of lower energy and higher entropy.
131. *Entropy (S)* is high in a highly disordered system, such as a gas, a messy room, etc.
132. Know processes that can separate mixtures and the concepts that allow the separations.

133. *Oxidation* is the *loss of electrons* (LEO) by an atom or ion. The oxidation number *increases* as a result. The electrons are on the *right side* of the reaction arrow.

Oxidation: $Zn \rightarrow Zn^{2+} + 2e^-$

134. *Reduction* is the *gain of electrons* (GER) 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.

$Cl^{o} + e^- \rightarrow Cl^{-1}$

135. Redox reactions *always* involve the transfer of *electrons*.

136. Remember…. "LEO says GER!" Lose Electrons Oxidation. Gain Electrons is Reduction.

137. *Identify redox reactions* by checking for a change in oxidation number such as an uncombined element on one side of a reaction that is in a compound on the other side.
138. $Zn + 2HCl \rightarrow ZnCl_2 + H_2$ is really: $Zn^{0} + 2H^{+1}Cl^{-1} \rightarrow Zn^{+2}Cl_2^{-1-1} + H_2^{0}$

139. In a redox reaction, the total number of electrons lost = the total number of electrons gained.
140. *Voltaic cells* produce electricity with a *spontaneous* redox reaction.
141. Use Table J to determine if a proposed redox reaction is spontaneous.
142. A helpful saying… "Leo is a *RED CAT*."
In electrochemical cells, *RED*uction occurs at the *CAT*hode. That's both volataic and electrolytic.

143. *Electrolytic cells* use an applied electrical current to force a nonspontaneous redox reaction to take place.
144. Electrolytic cells are can be used for metal plating of objects or separating a compound into its elements.
145. *Acids* and *bases* are both *good electrolytes*. Their solutions conduct electricity well.
146. Acids and bases turn *indicators* different colors. They're listed on *Table M*.
147. Acids have a pH < 7.
148. Bases have a pH > 7.
149. *Tables K & L* list names and formulas of common acids and bases asked about on the Regents.
150. The metals <u>above</u> H_2 on *Table J* will react with acids to make H_2 gas bubbles.
151. *Arrhenius* says:
"Acids give off Hydrogen Ions (H^+) or Hydronium Ions (H_3O^+) ions in solution."
"Bases give off Hydroxide Ions (OH^-) ions in solution."
152. *An alternate theory* says: "Acids *donate* protons." "Bases *accept* protons."

153. Acids and bases react in *neutralization* reactions to make *water* and a *salt*.

154. *Titrations* are controlled neutralization reactions used to find the concentration of an acid or base sample. Note the formula for it on Table T. Notice: M_A means concentration of acid ions H^+ and M_B means concentration of base ions OH^{-1}.

155. ALL organic compounds contain the element *carbon*.
156. *Carbon* ALWAYS makes *four bonds* in molecules.
157. *Saturated* hydrocarbons have all *single* bonds within their molecules (alkanes).
158. *Unsaturated* hydrocarbons at least one *double* or *triple* bond in their molecules (alkenes & alkynes).
159. *Hydrocarbons* contain ONLY the elements hydrogen and carbon.
160. The *homologous series* of hydrocarbons' formulas are on *Reference Table Q*.

161. The *functional groups* found in some organic molecules are listed on *Reference Table R*.
162. *Structural isomers* of organic compounds have *different* structural formulas but the *same* molecular formula.
163. *Combustion reactions* occur when a hydrocarbon reacts with oxygen to make CO_2 and H_2O.
164. *Organic substitution reactions* occur when an alkane and a Group 17 element react so that one or more hydrogen atoms on the alkane are replaced with an atom of the Group 17 element.
165. *Organic addition reactions* occur when an alkene or alkyne combine with a compound to make one product by adding the compound across the multiple bond.

166. *Esterification* occurs when an organic acid and an alcohol react to make water and an *ester*.
167. *Saponification* occurs when an ester reacts with a base to make alcohol and a *soap*.

168. *Fermentation* reactions occur when yeast catalyzes a sugar ($C_6H_{12}O_6$) to decompose into carbon dioxide and ethanol.
169. *Polymers* are long chains of repeating units called *monomers*.
170. Polymers form by *polymerization* reactions.
171. *Addition polymerization* occurs when unsaturated monomers join in a long polymer chain. (n represents a large number; usually in the thousands)

$$nC_2H_2 \rightarrow (C_2H_2)_n$$

172. *Condensation polymerization* occurs when monomers join to form a polymer *by removing a small molecular compound usually water.* (Water is a product.)

173. Unstable atoms that are radioactive are called *radioisotopes*. (*Table N*)
174. Radioisotopes can decay by giving off any of the particles/emanations listed in *Table O*.
175. *Alpha particles* (see Table O) are positively charged (+).
176. *Beta particles* (see Table O) are negatively charged (-).
177. The sum of the mass numbers and atomic numbers must be equal on both sides of the reaction arrow for nuclear equations.

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

178. *Fission reactions* split heavy nuclei into smaller ones.

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

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

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

180. The *half life* of a radioisotope is the *length of time* it takes for one half of the atoms in a sample to radioactively decay. (Table N)
181. C-14 is used to determine the ages of organic material up to 23,000 years old.
182. U-238 is used to determine the ages of rocks.
182. I-131 is used to treat thyroid disorders.
183. Co-60 is used to treat cancer tumors.
184. Radiation can be used to kill bacteria on foods to slow the spoilage process.
185. Disposal of radioactive waste is a HUGE problem associated with nuclear reactors.
186. No Blanks!!! Answer every question. If you don't know the answer, take an educated guess. Some chance of getting it right is better than none at all.
187. You have three hours to take the test, so take your time, ALL OF IT.
188. On every question consider if the answer is in the reference tables or if the reference tables could help you.
189. Your first choice is usually your best one. Only change an answer if you find an obvious mistake when checking your work.
190. Even if you think you know a formula, look it up. Most are on Table T (the last page). 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.
191. Eat a healthy meal the night before and for breakfast as well.
192. Get a good night's sleep. A tired mind is not as sharp and clear as a well-rested one.
193. Relax – you've seen all this stuff before!