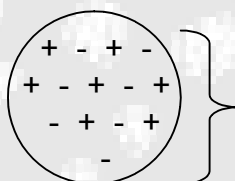


200 Ways to Pass the Chemistry Physical Setting Regents Exam

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 (**nucleons**).
5. Electrons are found in "clouds" (**orbitals**) around an atom's nucleus.
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 with equal numbers of protons, but differ in their neutron numbers.
10. **Cations** are *positive* (+) ions and form when a neutral atom *loses* electrons. They are *smaller* than their parent atom.
11. **Anions** are negative ions and form when a neutral 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.



Positive & negative particles 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 "planet-like" orbits 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 "**S**tandard **T**emperature and **P**ressure." (273 Kelvin & 1 atm)
19. 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.
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: 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 – "BrINClHOF" (Br_2 , I_2 , N_2 , Cl_2 , H_2 , O_2 , F_2)
23. Use this diagram to help determine the **number of significant figures** in a measured value...

Pacific Atlantic

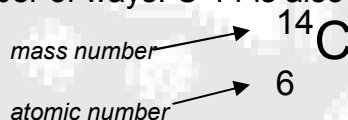
If the decimal point is **present**, start counting digits from the **Pacific** (left) side, starting with the first non-zero digit.

→ 1 2 3
0.00310 (3 sig. figs.)

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

3 2 1 ←
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, and is also



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 in your reference tables.

24.305
Mg
12 2-8-2

of electrons in 3rd principal energy level
of electrons in 2nd principal energy level
of electrons in 1st principal energy level

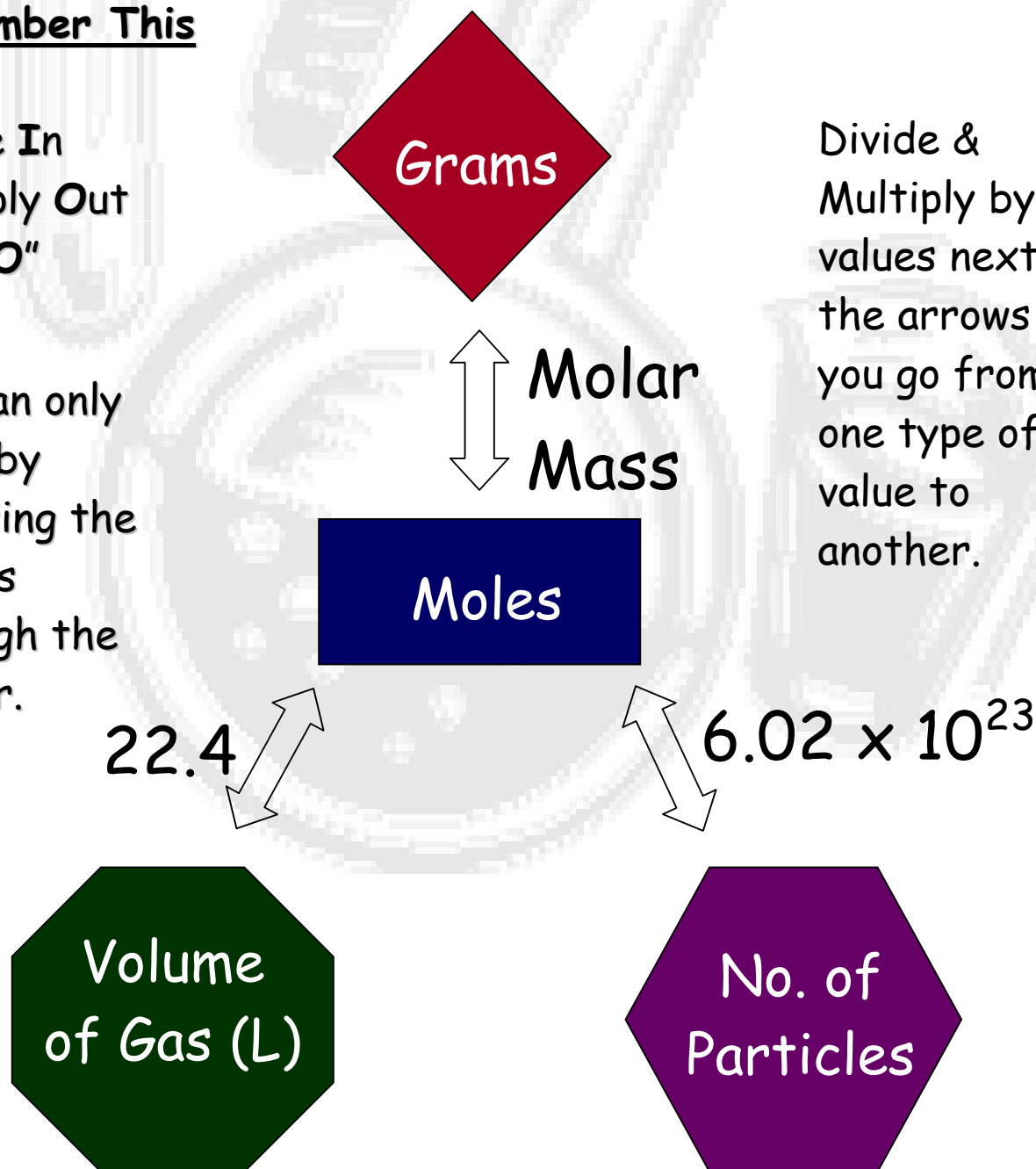
30. Use the **mole triangle diagram** on the next page to help you solve conversions between moles, grams, numbers of molecules/atoms, and liters of gases at STP...

The Mole Triangle

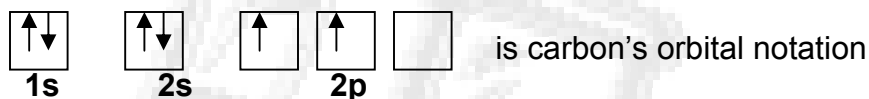
Remember This
Way!

Divide In
Multiply Out
"DIMO"

You can only
move by
following the
arrows
through the
center.



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

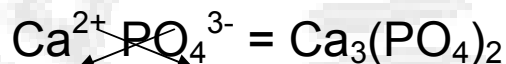


32. Polyatomic ions (Table E) are groups of atoms with an overall charge.
 NO_3^{1-} , NH_4^{1+} , SO_4^{2-} , etc.

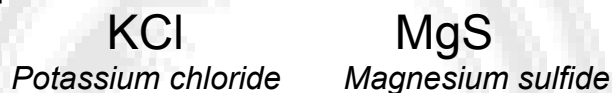
33. **Coefficients** are written in front of the formulas of reactants and products in chemical equations. They give us the ratios of reactants and products in a balanced chemical equation.

34. Chemical formulas are written so that the charges of cations and anions neutralize one another.

Example: *calcium phosphate*:



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.” Example:



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

Example:



37. **Physical changes** do not form new substances. They merely change the appearance of the original material. (The melting of ice)

38. **Chemical changes** result in the formation of new substances.
(The burning of hydrogen gas 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 left 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!

43. **Synthesis reactions** occur when two or more reactants combine to form a single product. Example: $2\text{H}_{2(g)} + \text{O}_{2(g)} \rightarrow 2\text{H}_2\text{O}_{(g)}$

44. **Decomposition reactions** occur when a single reactant forms two or more products. Example: $\text{CaCO}_{3(s)} \rightarrow \text{CaO}_{(s)} + \text{CO}_{2(g)}$

45. **Single replacement reactions** occur when one element replaces another element in a compound.



46. **Double replacement reactions** occur when two compounds react to form two new compounds.



47. The masses of the reactants in a chemical equation is always equal to 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. $\text{H}_2\text{SO}_4 = 98 \text{ g/mole}$

$$2 \times \text{H} = 2 \times 1 \text{ g/mole} = 2 \text{ g/mole}$$

$$1 \times \text{S} = 1 \times 32 \text{ g/mole} = 32 \text{ g/mole}$$

$$4 \times \text{O} = 4 \times 16 \text{ g/mole} = 64 \text{ g/mole}$$

$$\text{sum} = 98 \text{ g/mole}$$

49. Know how to calculate the percentage composition of a compound. (Formula is on Table T.)

50. 6.02×10^{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.

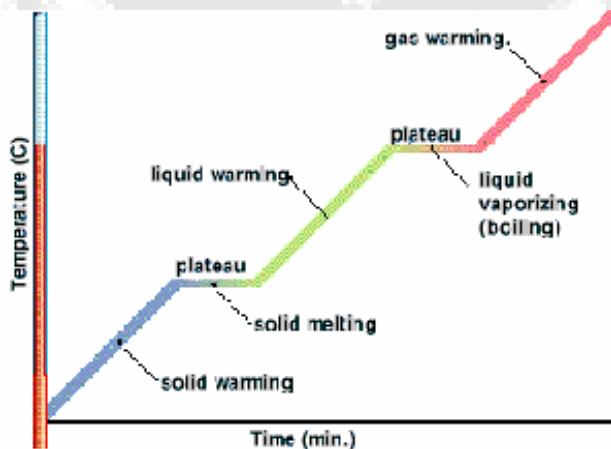
53. **Liquids** have closely-spaced particles that easily slide past one another.

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 as pictured below.



58. Substances that **sublime** turn from a solid directly into a gas. (CO_2 & I_2)

59. Degrees Kelvin = $^{\circ}\text{C} + 273$

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 it's 4.18

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 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 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 **pressure** on a gas increases, **temperature** increases.

65. As the **temperature** of a gas increases, **volume** increases.

66. *Always use Kelvins* for temperature when using the **combined gas law**.

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

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 mixtures of gases.

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.

74. **Groups** are vertical columns on the Periodic Table.

75. **Metals** are found left of the "staircase" on the Periodic Table, **nonmetals** are above it, and **metalloids** border it.

76. Memorize this chart.

Metals	Malleable	Ductile	Lustrous	Good conductors of heat & electricity	Low ionization energy and electroneg.	Tend to form + ions
Nonmetals	Brittle when solid	Mostly gases at STP	Dull	Good insulators	High ionization energy and electroneg.	Tend to form - ions

77. **Noble gases** (Group 18) are inert and stable due to the fact that their valence level of electrons is completely filled.

78. **Ionization energy** increases as you go up and to the right on the Periodic Table.

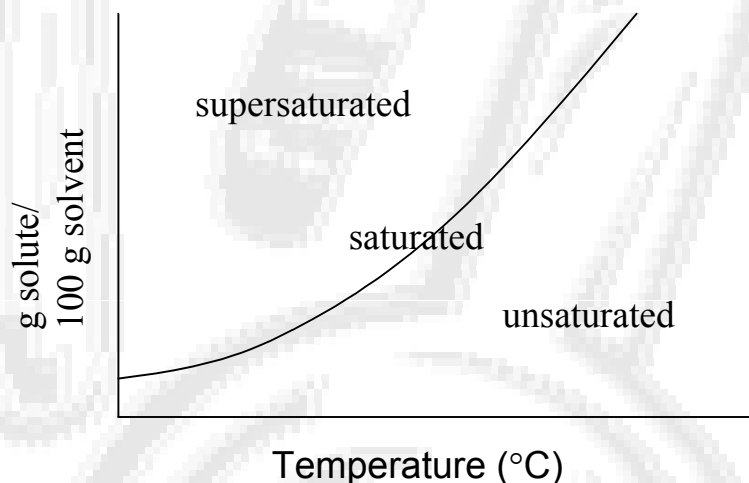
79. **Atomic radii** decrease left to right across a period due to increasing nuclear charge.

80. **Atomic radii** increase as you go down a group.
81. **Electronegativity** is a measure of an element's attraction for electrons.
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 the bond is.
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.
94. **Covalent bonds** form when two atoms **share** a pair of electrons.
95. **Ionic bonds** form when one atom **transfers** an electron 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 0.4 and 1.7.
98. **Ionic bonds** form when the electronegativity difference between two bonding atoms is *greater than* 1.7.
99. Substances containing mostly covalent bonds are called **molecular substances**.
100. Substances containing mostly ionic bonds are called **ionic compounds**.
101. Memorize this table.

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)

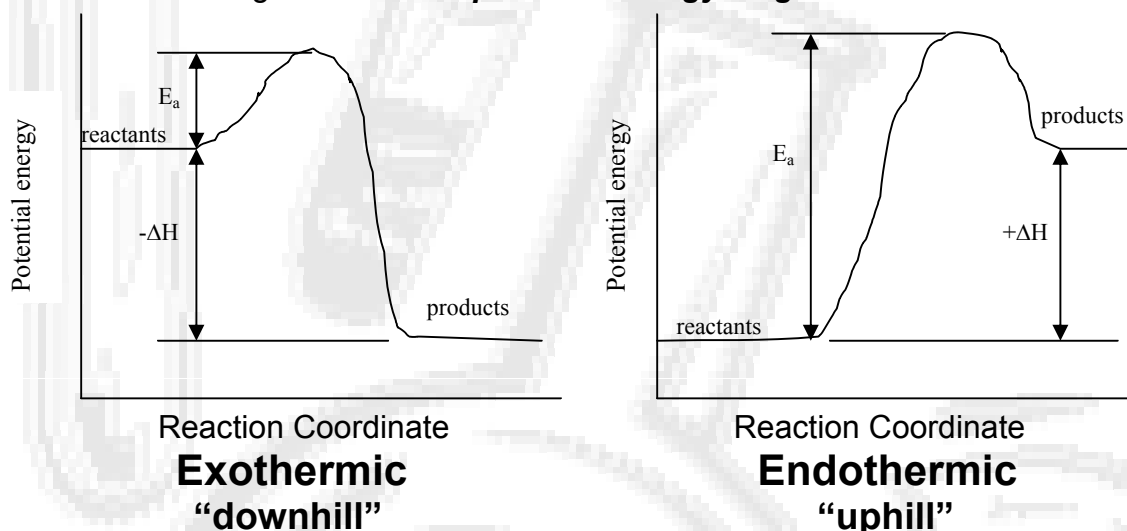
102. **Hydrogen bonds** form when hydrogen bonds to the elements N, O, or F and gives the compound unusually high melting and boiling points.
103. Use Table F to predict the solubilities of compounds.
104. Remember substances tend to be soluble in solvents with similar properties....
"Like dissolves like"
105. As temperature increases, solubility increases for most solids.

106. At low temperatures and high pressures solubility *increases* for most 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 by mass** = mass of the part / mass of the whole $\times 100\%$
110. **Parts per million (ppm)** = grams of solute / grams of solution $\times 1,000,000$
111. Solutes raise the boiling points and lower the melting points of 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 the pressure on gases increases reaction rate.
118. **Catalysts** speed up reactions by lowering their **activation energies**. 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**.



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**. It just establishes 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 $wA + xB \rightarrow yC + zD$, $K_{eq} = \frac{[C]^y[D]^z}{[A]^w[B]^x}$

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.



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.



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

136. Remember.... "LEO says GER!"

L ose	G ain
E lectrons	E lectrons
O xidation	R eduction

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

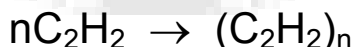


↑
Uncombined

↑
Zn is combined with Cl

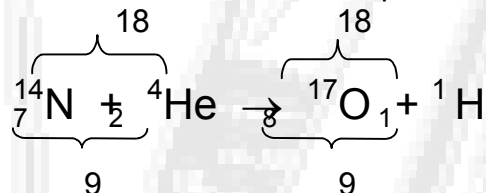
138. **Oxidizing agents** are what *get reduced* in a redox reaction.
Reducing agents are what *get oxidized* in a redox reaction.
139. **Electrochemical cells** produce electricity with a *spontaneous* redox reaction.
140. The *left electrode* is usually the site of *oxidation* in an electrochemical cell diagram.
141. Memorize this saying... "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. They're listed on **Table M**.
148. Acids have a $\text{pH} < 7$.
149. Bases have a $\text{pH} > 7$.
150. **Tables K & L** list names and formulas of common acids and bases asked about on the Regents.
151. The metals above H_2 on **Table J** will react with acids to make 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."
"Bases *accept* protons."
154. Acids and bases react in **neutralization** reactions to make **water** and a **salt**.

155. **Titration**s 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. **Carbon** ALWAYS makes **four bonds** in molecules.
158. **Saturated** hydrocarbons have all *single* bonds within them (alkanes).
159. **Unsaturated** hydrocarbons have *double* or *triple* bonds in them (alkenes & alkynes).
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** on organic molecules are listed on **Reference Table R**.
163. **Structural isomers** of organic compounds have *different* structural formulas but the *same* molecular formula.
164. Number the parent carbon chain in an organic molecule from the end closest to the alkyl group(s).
165. **Combustion reactions** occur when a hydrocarbon reacts with oxygen to make CO₂ and H₂O.
166. **Organic substitution reactions** occur when an alkane and a halogen (Group 17) reacts so that one or more hydrogen atoms on the alkane are replaced with oxygen.
167. **Organic addition reactions** occur when an alkene or alkyne combine with a halogen to make one product (halide).
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 catalyze a sugar (C₆H₁₂O₆) to make carbon dioxide and ethanol.
171. **Polymers** are long chains of repeating units called **monomers**.
172. Polymers form by **polymerization** reactions.
173. **Addition polymerization** occurs when unsaturated monomers join in a long polymer chain.

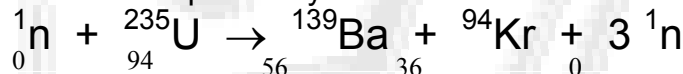


174. **Condensation polymerization** occurs when monomers join to form a polymer by *removing water*. Water is a product!
175. **Natural polymers** include starch, cellulose, and proteins.
176. **Synthetic polymers** include plastics such as nylon, rayon, and polyester.
177. Unstable atoms that are radioactive are called **radioisotopes**. (**Table N**)
178. Radioisotopes can decay by giving off any of the particles/emanations listed in **Table J**.
179. **Alpha particles** (see Table J) are positively charged (+).
Beta particles (see Table J) are negatively charged (-).

180. The sum of the mass numbers and atomic numbers must be equal on both sides of the reaction arrow for nuclear equations.



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



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



183. 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)
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.
190. USE THE REFERENCE TABLES!!!
191. Be sure to answer every question. If you don't know the answer, take a guess. Some chance of getting it right is better than none at all.
192. You have three hours to take the test, so take your time.
193. Try substituting words that seem confusing with a different word. Sometimes this makes 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.
195. Your first choice is usually your best one. Only change an answer if you find an obvious mistake when checking your work.
196. Even if you think you know a formula, look it up. 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.
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!