TOP TEN TIPS TO PASS THE CHEMISTRY REGENTS*

1. Substances are pure in composition: homogeneous (same composition throughout).
2. Substances must be either elements or compounds.
3. Mixtures can vary in composition (heterogeneous), or be homogeneous.
4. Aqueous solutions and mixtures of gases are homogeneous mixtures (same throughout).
5. Energy is released in an exothermic reaction; energy is absorbed in an endothermic reaction.
6. Heat is measured in calories or kilocalories (1 Kcal = 1000 calories)
7. Temperature is measured in degrees Celsius, or in Kelvin.
8. $\boxed{K = {}^{o}C + 273}$
9. ºA (absolute temperature scale) = K (Kelvin temperature scale); 1º change in K = 1º change in ºC
10. If temperature is changing, then calories = mass x change in temp $\boxed{Q = m\,\Delta t}$
11. If there is no change in temperature, then…

    calories = mass x heat of fusion (for melting/freezing)

    calories = mass x heat of vaporization (for boiling/condensing)
12. **Use the Reference Tables!**
13. Density = mass / volume
14. Combined Gas Law is $\boxed{\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}}$
15. Ice and liquid water exist in equilibrium at 0 ºC or 273 K.
16. Steam and liquid water exist in equilibrium at 100 ºC or 373 K.
17. There is no change in temperature (kinetic energy) during a phase change.
18. Potential energy changes during a phase change.
19. STP (standard temperature and pressure) is 0 ºC (= 273 K) and 1 atmosphere (= 760 mm Hg or "torr" = 101.3 kPa).
20. 1 mole = 6.02×10^{23} particles = 22.4 L of gaseous volume at STP = molecular weight of substance.
21. Sublimation occurs when a solid changes directly to a gas (common examples: $CO_{2(s)}$ and $I_{2(s)}$).
22. The diatomic gases are the BrINClHOF gang: Br_2, I_2, N_2, Cl_2, H_2, O_2, F_2 .
23. The monatomic gases or molecules are all the Noble Gases: He, Ne, Ar, Kr, Xe, Rn .
24. In neutral atoms the number of protons = the number of electrons.
25. Atomic number = the number of protons in any atom (or ion).
26. Protons and neutrons are both generically referred to as nucleons (particles in the atomic nucleus).
27. *Use the Reference Tables!*
28. Mass number = the number or protons + the number of neutrons.
29. Atomic weights shown on Periodic Table are the weighted average of all an element's isotopes found in nature.
30. Isotopes of an element have the same atomic number (# of protons) but different mass number (neutrons + protons).
31. Gram atomic mass = atomic mass of an element in grams.
32. Atomic weights on the Periodic Table are based on ^{12}C.
33. Electrons give off light energy when they fall from higher energy levels to lower ones.
34. Principal Quantum Number is the same thing as the principal energy level; corresponds to row number on Periodic Table.
35. Atomic sublevels are *s*, *p*, *d*, and *f* ; they have 1, 3, 5, and 7 orbitals, respectively.
36. Any orbital can hold only two electrons maximum; any energy level can hold $\mathbf{2n^2}$ electrons where n = the # of the energy level.
37. The valence electrons in an atom determine the atom's chemical properties.
38. Atoms lose, gain, or share valence electrons.
39. Ionization energy is the amount of energy necessary to remove an electron from an atom.
40. Opposite charges attract; like charges repel.
41. **Uh, Use the Reference Tables!**
42. When atoms bond, their energy conditions become more favorable (more stable/lower energy).
43. When chemical bonds are formed, energy is released; when chemical bonds are broken, energy must be absorbed.
44. Electronegativity is the ability of an atom to attract a pair of electrons while in a covalent bond.
45. Ionic bonds typically form between a <u>metal</u> and a <u>nonmetal</u>; another rule of thumb is if the difference in electronegativity values of the two elements bonding is ≥ **1.7**, the bond is likely ionic.
46. Polar covalent bonds form between two different <u>nonmetals</u>; another rule of thumb is if the difference in electronegativity values of the two elements bonding is **> 0.3 but < 1.7**, the bond is likely polar covalent
47. Nonpolar molecules are formed by molecules with nonpolar covalent bonds (electrons are shared equally) (electronegativity differences of **0.0 to 0.3**).
48. Polar molecules result from bond dipoles (unequal sharing of electrons) which **do not cancel out**. Remember **SNAP!!** (**<u>s</u>**ymmetrical = **<u>n</u>**onpolar, **<u>a</u>**symmetrical = **<u>p</u>**olar)
49. Three nonpolar molecules that have polar bonds which cancel out due to symmetry are CH_4, CCl_4, and CO_2.
50. Covalent bonds between atoms form substances that are molecular in nature.
51. Network solids are large covalent molecules with extremely high melting points and are very durable materials.

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* *okay, so it's obviously more than 10 tips, just get off my back, okay?*

52. Diamond, SiO_2 (silicon dioxide), SiC (silicon carbide), and graphite are all examples of network solids.

53. All metals form <u>metallic</u> bonds (a sea of mobile valence electrons).

54. Ionic compounds conduct electricity in aqueous solutions or in their liquid ("molten") state.

55. Covalent substances or molecules do not conduct heat or electricity well.

56. **Bonds** are what form between atoms <u>within</u> a molecule to keep a molecule together.

57. **Attractions** are what is felt <u>between</u> molecules to keep a group of molecules together.

58. Coordinate covalent bonds are formed when one atom/ion is the source of <u>both</u> of the electrons shared in the bond pair.

59. Two examples of species with a coordinate covalent bond in them are the ammonium ion (NH_4^+) and the hydronium ion (H_3O^+).

60. Dipole–dipole attractions occur between two <u>polar</u> molecules.

61. Hydrogen bonding is <u>not</u> *actually* a bond, but an attractive force between two molecules.

62. Hydrogen bonding is a strong dipole attraction involving hydrogen and a strongly electronegative atom (F or O or N) ("FON bonds")

63. Van der Waals attractions occur between nonpolar molecules.

64. Van der Waals attractions get stronger with the increasing molecular weight of the molecules involved (*i.e.* they get larger).

65. Molecule–ion attractions occur in aqueous solutions or hydrated molecules.

66. Likes dissolve likes. (Polar molecules dissolve other polar molecules, nonpolar molecules dissolve other nonpolar molecules.)

67. "ic" acids form "ate" ions; "ous" acids form "ite" ions.

68. Binary compounds are composed of two different elements and end in "ide."

69. Ternery compounds are composed of three different elements and contain a polyatomic ion.

70. Names of common polyatomic ions appear on Reference Table E.

71. The total charge on the atoms in a compound must add up to <u>zero</u>.

72. Group 1 elements are called alkali metals.

73. Group 2 elements are called alkali earth metals.

74. Group 17 elements are called halogens.

75. Groups 18 elements are called noble gases, inert gases, or monatomic gases.

76. Elements from Groups 3 through 11 are called transition metals.

77. Transition metals form ions that are colored when dissolved in solution.

78. Metals form ions that are smaller than their neutral atoms by losing electrons.

79. Nonmetals form ions that are larger than their neutral atoms by gaining electrons.

80. Metallic characteristics increase as you go farther down and to the left on the Periodic Table (*e.g.* Cs and Fr).

81. Nonmetallic characteristics increase as you go farther up and to the right on Periodic Table (*e.g.* O and F).

82. Br (bromine) is the only nonmetal that is a liquid at room temperature.

83. Hg (mercury) is the only metal that is a liquid at room temperature.

84. All elements atomic number 83 and greater are considered to be radioactive (have no naturally occurring stable isotopes).

85. The periodic table is arranged by increasing atomic number (number of protons in an atom).

86. Transition metals are found pure in nature; Group 1 and Group 2 metals are only found in compounds in nature (highly reactive).

87. Fluorine only has an oxidation state of –1.

88. N_2 contains a triple covalent triple bond in it and is very stable.

89. About $^3/_4$ of all the elements are metals.

90. Metals have low ionization energies (easy to remove electrons) and low electronegativities (do not want to gain electrons).

91. Allotropes are different forms of the same element (O_2 [pure oxygen gas] *vs.* O_3 [ozone], graphite *vs.* diamond *etc.*).

92. Boiling and melting points are directly proportional to the strength of a substance's intermolecular forces (the stronger the forces, the higher the melting point and boiling point).

93. ***Fer cryin' out loud, use your Reference Tables!***

94. $$\boxed{\text{Percent Error} = \frac{\left|\text{Experimental Value} - \text{Accepted Value}\right|}{\text{Accepted Value}} \times 100\%}$$

95. $$\boxed{\text{Percent Composition by Mass} = \frac{\text{mass contributed by part}}{\text{total mass (molecular weight)}} \times 100\%}$$

96. An empirical formula is the lowest, whole number ratio of elements in a compound.

97. A molecular formula of a compound is a multiple of the empirical formula (all elements multiplied through by a common factor).

98. Some empirical formulas may also happen to serve as a compound's molecular formula (*e.g.* P_2O_5)

99. Zeroes after a decimal point but <u>before</u> the first non-zero digit are <u>**not**</u> considered significant (they just hold the place of the decimal).

100. Zeroes between any non-zero digits automatically <u>are</u> significant.

101. Zeroes after the decimal <u>and</u> after any non-zero digits <u>are</u> considered significant (help specify how precise a value is).

102. When adding together numbers that have different amounts of significant figures, always report the result to the <u>same</u> number of decimal places as has the value with the <u>least</u> number of significant figures.

103. Always add acid to water, <u>not</u> the other way around (it's more dilute that way, and therefore safer).

104. The greater the amount of air (O_2) present, the cleaner a Bunsen burner flame burns.

105. Heat of reaction = stored energy of products minus stored energy of reactants $\boxed{\Delta \mathbf{H} = \mathbf{H}_{\text{products}} - \mathbf{H}_{\text{reactants}}}$

106. When ΔH = (–), the reaction is exothermic (heat/energy is given off); when ΔH = (+), the reaction is endothermic (heat/energy is absorbed).

107. The products of an exothermic reaction are more stable (have less energy) than the reactants.

108. The greater the exothermic nature of the reaction (*i.e.*, the more negative ΔH), the more stable the products.

109. Remember to use Chart I !

110. The rate of a chemical reaction will generally increase if the surface area, temperature and/or concentration of the reactants is increased, or if a catalyst is used.

111. At equilibrium, the rate of the forward reaction = the rate of the reverse reaction (net amounts of reactants and products remain constant).

112. Solution equilibrium is found in a solution that is saturated.

113. An increase in temperature generally increases the solubility of solids in a liquid.

114. An increase in temperature generally decreases the solubility of gases in a liquid.

115. An increase in concentration will favor the direction of reaction that will use up the increase.

116. An increase in pressure will favor the direction of reaction that yields less total moles/less total volume, for gases only.

117. An increase in temperature favors the direction of reaction that is endothermic in nature (uses up the heat).

118. A catalyst has no net effect on equilibrium, only how fast equilibrium is reached.

119. $\boxed{\mathbf{K}_{eq} = \dfrac{[products]^{coefficients}}{[reactants]^{coefficients}}}$

120. Only gases and aqueous solutions are included in K_{eq} expressions; ignore pure solids and pure liquids

121. Large values of K_{eq} (>1) means products are favored; small values of K_{eq} (<1) means reactants are favored.

122. A change in **temperature** is the only factor which can change the value of K_{eq}.

123. The larger the value of K_{sp}, the greater the solubility a substance has (easier it is for it to dissolve).

124. The larger the value of K_a (or K_b), the greater the degree to which the substance ionizes, and the better an electrolyte it is.

125. $\boxed{\mathbf{\Delta G = \Delta H - T\Delta S}}$

126. When ΔG = (–), a chemical reaction will occur spontaneously; when ΔG = (+), the reaction is not spontaneous; when $\Delta G = 0$, the reaction is at equilibrium.

127. Entropy is measure of randomness or disorder in a system.

128. Nature always seeks ideally to minimize total energy (decrease ΔH) and maximize randomness (increase ΔS) in any given system.

129. Acids produce H^+ or H_3O^+ ions when dissolved in aqueous solutions.

130. Bases produce OH^- (either directly or by reacting with water) when dissolved in aqueous solutions.

131. Acids and bases both conduct electricity when dissolved in aqueous solutions.

132. Aci**d**s turn litmus re**d** (or pink) and do not affect the color of phenolphthalein.

133. **B**ases turns litmus **b**lue and phenolphthalein pink.

134. Acids are proton or hydrogen ion donors.

135. Bases are proton or hydrogen ion acceptors.

136. The terms "amphitprotic" and "amphoteric" each refer to substances that can act either as an acid or a base.

137. $\boxed{\mathbf{M}olarity = \dfrac{\#\ moles\ of\ solute}{volume\ of\ solution\ in\ Liters}}$

– or – $\boxed{\#\text{ moles of acid or base} = \text{volume x molarity}}$

138. The two members of a conjugate acid-base pair differ only by one hydrogen ion.

139. Strong acids completely ionize (products strongly favored) and have a very large K_a; they are found at or near the top of Chart L.

140. Weak acids do not ionize completely (reactants are favored) and have a small K_a; they are found at or near the bottom of Chart L.

141. Strong bases are those that contain alkali metals — Group 1, *e.g.* LiOH and NaOH.

142. $\boxed{\mathbf{K_w = [H^+][OH^-] = 1.0 \times 10^{-14}}}$

143. $\boxed{\mathbf{pH = -\log[H^+]}}$

144. pH = 7 is a neutral solution; pH > 7 is a basic solution; pH <7 is an acidic solution.

145. For titrations, $\boxed{\mathbf{M_aV_a = M_bV_b}}$

146. Salts of strong acids and weak bases hydrolyze to produce acidic solutions.

147. Salts of strong acids and strong bases do not hydrolyze (they neutralize).

148. Electrolytes are substances which form solutions capable of conducting electrical currents when the substance is dissolved in water.

149. Oxidation is a loss of electrons (which means an increase in oxidation number). **"LEO"**

150. Reduction is a gain of electrons (which means a decrease [*reduction*] in oxidation number). **"GER"**

151. The oxidation number of any pure, uncombined element is **zero**.

152. The total oxidation numbers in a neutral compound must add up to zero.

153. The oxidation numbers of all the atoms in an ion must add up to the charge on the ion.

154. An oxidizing agent is the species that undergoes reduction (and thereby facilitates oxidation).

155. A reducing agent is the species that undergoes oxidation (and thereby facilitates reduction).

156. When using charts for standard reduction potentials, remember to reverse the direction of the indicated reaction and the sign on its corresponding E^0 value when you need an oxidation reaction.

157. If the net E^0 = (+), the underlying half-reaction combination will be spontaneous; if the net E^0 = (–), the reaction is not spontaneous; if the net E^0 = 0, the reaction is already at equilibrium.

158. Electrochemical cells produce electricity from a spontaneous redox reaction only.

159. Nonspontaneous redox reactions can be forced to occur by applying an electrical current, in an electrolytic cell.

160. Reduction always occurs at the cathode. **RED CAT**

161. Oxidation always occurs at the anode. **AN OX**

162. The signs on the anode/cathode are exactly opposite when comparing an electrochemical cell with an electrolytic cell (electrochemical: cathode = (+), anode = (–); electrolytic: cathode = (–), anode = (+)).

163. In an electrolytic cell, the object being plated is attached to or actually is the cathode itself.

164. Carbon has the ability to form four covalent bonds which are directed out in the shape of a regular tetrahedron.

165. Organic compounds are generally nonpolar, insoluble in water (likes dissolve likes), are non-electrolytes, have low melting and boiling points and high vapor pressures.

166. Each carbon atom must have four bonds or eight electrons around it.

167. Saturated compounds contain all single bonds (the molecule is "saturated" with atoms: no new atoms can be added).

168. Unsaturated compounds must contain at least one double bond or triple bond in it somewhere.

169. Hydrocarbons contain only hydrogen and carbon atoms.

170. Homologous series of hydrocarbons have each member differing from the next by one C and two H's (one CH_2 group).

171. Homologous series of hydrocarbons include:

alkanes $\longrightarrow C_nH_{2n+2}$	all single bonds
alkenes $\longrightarrow C_nH_{2n}$	at least one C=C double bond
alkynes $\longrightarrow C_nH_{2n-2}$	at least one C≡C triple bond
benzene $\longrightarrow C_nH_{2n-6}$	ring structures

172. Isomers have same molecular/chemical formula but different physical structures.

173. As the number of atoms in a hydrocarbon increases, the number of possible isomer structures increases as well.

174. Functional groups identify different types of organic compounds.

175. Alcohols contain one or more hydroxyl (—OH) groups.

176. Aldehydes contain a —CHO group, which are always attached to a terminal (end) carbon.

177. Ketones contain a $R\text{—}\overset{\overset{\displaystyle O}{||}}{C}\text{—}R'$ group and cannot be attached to a terminal (end) carbon.

178. Organic acids contain the carboxyl group (—COOH), which must always be attached to a terminal (end) carbon.

179. Organic acids are the only organic molecules that typically are considered electrical conductors, and even then the conductivity is only weak/slight.

180. Ethers contain a R—O—R' group (two sides of a molecule linked together by an "oxygen bridge"); ethers are made by the dehydration of two primary alcohols.

181. Esters contain the —COO— group and have very characteristic smells and flavors.

182. Amines have the functional group R R' R"—N and show up in everything from pharmaceuticals to amino acids

183. Amides have the functional group R—CONH—R'

184. Exotic alkyl halides (halogen derivatives of parent compounds) can be produced by adding certain halogens (*i.e.* F, Cl, Br, I) to a hydrocarbon molecule (called "halogenation")

185. Substitution reactions involve saturated hydrocarbons (only) and form two products.

186. Addition reactions involve unsaturated hydrocarbons (only) and form one product.

187. Polymers are large molecules made up repeating, smaller units called monomers; two important types are condensation polymerization and addition polymerization

188. Fermentation reactions usually use zymase as the enzyme (organic catalyst) to generate alcohol and CO_2 from sugars.

189. Oxidation of an organic compound usually produces CO_2 and H_2O, and releases a tremendous amount of heat/energy.

190. Half-life is the amount of the time required for the nuclei of a radioactive isotope to decay to one half its original mass

191. Fission is the splitting of heavy nuclei into more-stable nuclei of intermediate mass; fusion is the combining of light-mass nuclei to form a heavier, more-stable nucleus

192. Five main classes of reactions are synthesis, decomposition, combustion, single replacement, and double replacement.

The RIGHT WAY to Take A Test:

Before the test:

1. Make sure you study for the test (duh).
2. ~~Pay someone smart to take the test for you.~~
3. ***Get a good night's rest before the test, even if it means less last minute cramming — all the cramming in the world won't help you if you can't stay awake during the test.***
4. ***Eat a good meal before you leave for school, or at very least drink a glass of milk — your brain needs lots of protein to think, so don't starve it!***
5. Give yourself plenty of time to get to school and get settled in so you do not feel rushed.
6. Remember all necessary test-taking materials such as pens, pencils, calculators, *etc.*

During the test:

1. Pace yourself so you have enough time to complete the entire test. If necessary, bring a watch or similar so you always know what time it is.
2. If you are struggling with a particular question, circle it and move on — don't waste time on it, you can always return back to it later if time permits.
3. Eliminate the choices that are obviously wrong and then choose the best answer from those that remain.
4. Often the correct answer will be the one that doesn't "fit the pattern" that the other three answers establish — if you must guess, try to find the answer that does not fit the pattern.
5. **DON'T LEAVE ANY ANSWERS BLANK** — guesses cost nothing but you will certainly get a question wrong if you don't write *anything*.
6. When finished, check your answers over and look for any obvious mistakes, unanswered/skipped questions and so on.