230 Facts You Need To Know To Pass The Regents Chemistry Exam

1. Atoms are the smallest unit of matter.
2. Atoms consist of a central nucleus.
3. The nuclei contain protons and neutrons.
4. Electrons occupy regions around the nucleus called **orbitals**.
5. Protons are positively charged.
6. Protons have a mass of 1 amu.
7. Neutrons have no charge (are neutral).
8. Neutrons have a mass of 1 amu.
9. Electrons are negatively charged.
10. Electrons have an effective mass of 0 amu (1/1836 of a proton).
11. The nucleus as a whole is positively charged.
12. **Isotopes** are atoms with the same atomic number but with different atomic masses (different # of neutrons).
13. There are three common isotopes of hydrogen:
14. Protium (also known as Hydrogen-1) is one isotope (1 p^+, 0 n^0)
15. Deuterium (also known as Hydrogen-2) is a second isotope (1 p^+, 1 n^0)
16. Tritium (also known as hydrogen-3) is a third isotope (1 p^+, 2 n^0)
17. Tritium is radioactive! (decays by β emission)
18. **Mass number** = the sum of protons and neutrons
19. Atomic mass (as found on the Periodic Table) is the weighted average of all the isotopes of that element (based on their relative abundances in nature).
20. Bohr's model of the atom is commonly accepted as a simple model of the atom.
21. The energy levels are named K, L, M, N, O, P *etc*. or numbered 1, 2, 3, 4, 5 *etc*.
22. These energy levels are also known as **principal quantum numbers**.
23. Each orbital holds a distinct numbers of electrons ($2e^-$)
24. Electrons give off distinct colors when they return to their normal energy levels (the "**ground state**") after they are energized; characteristic "**spectral lines**" are emitted as electrons move from high energy level to low energy levels.
25. The last/highest energy level with electrons in it is called the **valence shell**.
26. An atom's "**kernel**" is its nucleus and all energy levels *except* the valence energy level.
27. The Periodic table of the elements is an organized chart of all known elements.
28. Elements on the periodic chart are arranged by atomic number (*i.e.* increasing nuclear charge).
29. Elements are classified as metals, nonmetals, and metalloids.
30. Metalloids are elements that display properties of both metals and nonmetals.
31. Elements can be grouped according to their valence shell electron number.
32. Elements with 1, 2, and/or 3 valence shell electrons are metals.
33. Elements with 5, 6, and/or 7 valence shell electrons are nonmetals.
34. Metalloids are elements along the "staircase" boundary between metals and nonmetals.
35. **Noble gases** are in Group 18 of the periodic chart.
36. Noble gases have 8 valence shell electrons (already satisfy the "**Octet Rule**") (except Helium, which has 2 valence electrons).
37. Vertical **groups** (columns) of elements display similar chemical properties.
38. Horizontal rows are called **periods**; properties change slowly from metallic to nonmetallic across a period.
39. The period (row) number represents the number of occupied principal energy levels for a given element.
40. Atomic radius is the closest distance to which one atom can approach another of the same type of atom.
41. Covalent radius is the distance from the center of the nucleus to the outer valence shell of that atom (or, half the distance between two covalently bonded atoms of the same element).
42. Ionic radii of metals are *smaller* than their atomic radii (lose electrons).
43. Ionic radii of nonmetals are *larger* than their atomic radii (gain electrons).
44. **Electronegativity** is a measure of the ability of an atom to attract another atom's electrons while in a chemical bond.
45. Metals generally have low electronegativies and give away their electrons (have low **ionization energies**).
46. Nonmetals have higher electronegativies and gain electrons (have high ionization energies).
47. Fluorine has the highest electronegativity value of all elements (4.0).
48. Fluorine has the greatest attraction for electrons.
49. Electronegativity differences between bonded atoms help predict the bond type (0 to 0.3 / 0.3 to 1.7 / > 1.7)
50. Ionic bonds and covalent bonds are the two most common bond types.
51. An electronegativity difference of 1.7 or more between the atoms involved indicates an ionic bond.
52. An electronegativity difference less than 1.7 but greater than 0.3 between the atoms involved indicates a polar covalent bond.

53. An electronegativity difference of 0.0 to 0.3 indicates a nonpolar covalent bond.
54. Don't forget to check atom types! Metal bonded to nonmetal = ionic bond!! Ditto for polyatomic ions!
55. A coordinate covalent bond is one where one atom contributes both electrons shared in the bond, and the other atom contributes none.
56. Remember the SNAP rule! (**s**ymmetrical = **n**onpolar, **a**symmetrical = **p**olar)
57. Group 1 elements are known as **alkali metals**; Group 2 metals are known as the **alkaline earth metals**.
58. Group 1 and 2 metals are so reactive that they are only found in nature already bound up in compounds.
59. Elements in Group 13 all have oxidation numbers of +3.
60. Nitrogen (in Group 15) forms a triple covalent bond when it bonds with itself (diatomic as N_2).
61. Group 17 elements are known as the **halogens.**
62. Halogens are the only group of elements that show all three phases of matter at room temperature (F_2, Cl_2 are gases; Br_2 is a liquid; I_2 is a solid).
63. The **noble gases** are group 18 and are very unreactive elements ("**inert**") due to their complete outer energy levels.
64. Diatomic molecules are elements that are combined with themselves; think **BrINClHOF** !
65. Diatomic molecules have electronegativity differences of zero, so they all are bonded with nonpolar covalent bonds.
66. Chemical formulas tell the small, whole number ratio of elements as they are combined with each other.
67. An empirical formula is the smallest, whole number ratio of atoms in a formula; it cannot be reduced.
68. Structural formula shows the physical arrangement of atoms as they are combined together.
69. In a chemical formula the sum of the oxidation number (totals) is assumed equal to zero, unless otherwise noted by a nonzero charge.
70. Oxygen has an oxidation number of -2 except in peroxides (*e.g.* H_2O_2) where it is –1, or bonded to a halogen (*e.g.* OF_2) where it is +2.
71. Hydrogen has an oxidation number of +1 except in metal hydrides (*e.g.* LiH) where it is –1.
72. Compounds classified as binary have only two different types of elements in the formula (*e.g.* NaCl or H_2O).
73. Compounds classified as ternary have three different elements in their formula (*e.g.* $NaNO_3$ or $Al_2(SO_4)_3$); think **polyatomic ions** (look on Table E).
74. Polyatomic ions have a net charge either positive or negative (see Table E).
75. Use a roman numeral when naming compounds where the positive element has more than one positive choice (*e.g.* Cu^{+1} = "copper(I)", Cu^{+2} = "copper(II) *etc.*).
76. Acids with only two different elements are called **binary acids** (*e.g.* HCl) (see table K).
77. Acids with three different elements are called **ternary acids** (*e.g.* HNO_3 ; H is +1 and NO_3 is –1)
78. Chemical equations are statements of a reaction.
79. Chemical equations must be balanced (otherwise you're breaking the **Law of Conservation of Mass**!).
80. Only manipulate/change **coefficients** when balancing a chemical equation.
81. The **mole** unit is used to count atoms and molecules.
82. **Avogadro's number** is 6.022×10^{23}.
83. The mass of one mole of any substance is the sum of the atomic weights composing the substance.
84. Gram atomic mass is the mass in grams of a mole of atoms (*e.g.* Sulfur is 32.0 amu or 32.0 grams per mole [containing Avogadro's number of atoms]).
85. Gram formula mass ('**molecular weight**") is the sum of the masses of a molecule. (*e.g.* H_2O = 18 amu or 18 grams per mole)
86. One mole of any gas occupies 22.4 liters (when it is at **STP**: 0ºC and 1 atm pressure).
87. **Stoichiometry** is a study of the molar proportions.
88. **Percent composition** is the relative percentages of each element in a formula.
89. There are 5 basic reaction types: **single replacement**, **double replacement**, **synthesis**, **decomposition** and **combustion**.
90. Bonds between atoms when electrons are <u>transferred</u> are ionic bonds (electronegativity difference 1.7 or greater).
91. Some characteristics of ionically bonded substances:
92. High melting points
93. Geometric structure of solid ionic crystals (called "**crystal lattice**")
94. Ionic <u>solids</u> do <u>not</u> conduct electricity.
95. Ionic substances <u>*will*</u> conduct electricity if melted ("**molten**") or dissolved in water ("**aqueous**").
96. Covalent bonded substances share electrons.
97. Some covalent substances share electrons unequally; these bonds are called **polar covalent bonds**.
98. Electronegativity difference in polar covalent substances is > 0.3 but less than 1.7.
99. Electronegativity difference in nonpolar substances is 0.0 to 0.3.
100. Two prime examples of coordinate covalent substances are NH_4^+ ("**ammonium ion**") and H_3O^+ ("**hydronium ion**").
101. **Molecular substances** have discrete particles formed from covalently bonded atoms.
102. Molecular substances can exist as liquids, solids or gases.

103. Also, they are good **insulators** (poor conductors of both heat and electricity) and have low melting points (*e.g.* H_2O, CO_2, $C_6H_{12}O_6$ *etc.*).
104. **Network solids** contain covalent bonds in a giant, repeating "**macromolecule**" network.
105. Also, they are hard, poor electrical and heat conductors and have *very* high melting points (an exception to this is graphite, which is a very good electrical conductor).
106. Diamond (C), Graphite (C), Silicon Carbide (SiC) and silicon dioxide (SiO_2) are all prime examples of network solids.
107. Metallic bonding occurs between atoms of metals that have small numbers of valence electrons (*e.g.* Group 1 and Group 2 metals) and lots of room in their valence shells ("**vacant valence orbitals**").
108. Metallic bonds consist of a regular arrangement of positive ions in a crystal lattice and are immersed in a "**sea of mobile electrons**."
109. Gaining or losing electrons results in a change in an atom's size (increases/decreases, respectively).
110. Metals lose electrons to form positive ions and have smaller radii than they did as neutral atoms.
111. Nonmetals gain electrons to form negative ions and have larger ionic radii than they did as neutral atoms.
112. **Intermolecular forces** (IMF's) are attractions *between* neighboring atoms/molecules; they are not *bonds* (which are found *within* a molecule, holding it together).
113. Examples of intermolecular forces are dipole-dipole attractions, hydrogen bonding, Van der Waals / London forces, and ion-molecule attractions.
114. **Dipole-dipole attractions** are attractions between the negative end of a polar molecule and the positive end of another polar molecule of the same type.
115. **Hydrogen bonds** are formed between molecules when hydrogen is covalently bonded to an element of small radius and high electronegativity, common in compounds of hydrogen and oxygen, fluorine or nitrogen ("FON bonds").
116. The relatively strong attractions between water molecules are due to hydrogen bonding.
117. **Van der Waals** forces are weak attractions between nonpolar molecules.
118. **Molecule-Ion attractions** are between an ionic substance and a molecular substance such as NaCl in water.
119. Matter commonly exists in three phases: solid, liquid and gas.
120. **Vapor pressure** is a measure of how much vapor is present above the surface of a liquid; the more vapor present (in other words, the more **volatile** the liquid is), the weaker the IMF's within the liquid must be.
121. **Boiling point** is the temperature at which a liquid's vapor pressure is equal to the atmospheric pressure.
122. Water normally boils at 100ºC and at 1 atmosphere pressure.
123. Matter in the gaseous phase has high potential energy (the particles are very far apart).
124. A phase change is a physical change only, no chemical change takes place.
125. Melting and freezing points of a substance must be at the same temperature.
126. While ice is melting, the temperature remains the same (KE remains constant).
127. When water boils its temperature remains the same (100ºC), but PE is increasing.
128. Boiling water is an endothermic process; so are melting and sublimation too.
129. Freezing water is an exothermic process; so are condensation and deposition too.
130. Sublimation occurs when a substance goes directly from the solid to the vapor phase, skipping the liquid phase. CO_2 and I_2 are two common substances that sublime.
131. Compounds contain two or more different elements that are *chemically* combined.
132. Mixtures are two or more substances *physically* combined.
133. A mixture that is **homogeneous** (uniform throughout) is a solution like aqueous NaCl.
134. A mixture that is **heterogeneous** (not uniform throughout) is sand and water.
135. Alloys are mixtures of metals.
136. Solutions contain two parts: a **solute** and a **solvent**.
137. The solute is the substance in lesser quantity, the solvent is the greater quantity.
138. Remember that "**Likes dissolve likes**" (polar dissolves polar, nonpolar dissolves nonpolar).
139. **Molarity** is a measure of a solutions strength or *concentration*; it is specified in moles of solute per liter of solution.
140. **Immiscible** means not mixing (literally, not mixable).
141. Dilute solutions are weak solutions (low amounts of solute per unit solvent).
142. Solutions can be saturated, unsaturated or supersaturated, depending on the amount of solute dissolved per given amount of solvent.
143. **Freezing point depression** is the lowering of the freezing point of a solvent by introducing a nonvolatile solute into the pure solvent.
144. **Boiling point elevation** is the raising of the boiling point of a solvent by introducing a nonvolatile solute into the pure solvent.
145. **Temperature** is the average kinetic energy of the particles of a substance.
146. Fixed points on a thermometer are 0ºC and 100ºC.
147. Kelvin is the absolute temperature scale.
148. **Absolute zero** is 0 Kelvin (–273ºC).

149. **Specific heat** is the measure of the amount of energy required to raise the temp of 1 gram of a substance by 1 degree.
150. **Real gases** have many variations from **ideal gases** (because gas particles actually do take up volume in space, and they do exhibit electrical interactions with each other).
151. H_2 and He are the two real gases closest to ideal (typically, the smaller a gas' molecular mass, the more ideally it will behave).
152. **Boyle's Law** states that when temperature remains constant, the volume of a gas varies inversely with its pressure.
153. **Charles' Law** states that when pressure remains constant, the volume of a mass of gas varies directly with its Kelvin temperature.
154. **Gay-Lussac's Law** states that when volume remains constant, the pressure of a mass of gas varies directly with its Kelvin temperature.
155. The **Combined Gas Laws** rolls Boyle's, Charles', and Gay-Lussac's Law all into one convenient and easy to use form (see table T).
156. Standard Temperature and Pressure are 0ºC and 1 atmosphere pressure (see table A).
157. **Dalton's Law of Partial Pressures**: total pressure in a mixture of gases is the sum of the pressures of each gas in the mixture.
158. **Graham's Law of Effusion**: the relative rates at which two gases will diffuse/effuse are proportional to the square roots of their molar masses.
159. **Avogadro's Hypothesis**: equal volumes of gases at the same conditions of temperature and pressure contain equal numbers of molecules.
160. **Kinetics** is the branch of chemistry dealing with the rates of chemical reactions and their mechanisms.
161. Products of exothermic reactions are <u>more</u> stable than the reactants from which they formed (negative "heat of formation" values) (see Table I).
162. Products of endothermic reactions are <u>less</u> stable than the reactants from which they formed (positive "heat of formation" values) (see Table I).
163. Reactions occur because of **effective collisions** between particles.
164. Potential energy diagrams graphically show the results of a chemical reaction.
165. **Activation energy** is the amount of energy required to start any reaction (whether it is endothermic or exothermic).
166. A **catalyst** lowers the activation energy needed to initiate a reaction.
167. Change in **enthalpy** (heat content) during a chemical reaction is called the **heat of reaction** (ΔH).
168. An increase in the concentration of a reactant increases the rate of reaction.
169. An increase in the temperature at which a reaction occurs increases the rate of the reaction.
170. Greater surface area of reactants increases the rate of reaction.
171. **Equilibrium** is a balance between the forward and reverse reactions in a closed chemical system.
172. Phase equilibrium is an equilibrium that exists between two different phases (*e.g*, between ice and liquid water in a system at 0ºC and a 1 atm).
173. **LeChatelier's Principle** explains the effects of changes in pressure, concentration, and temperature on a system already at equilibrium.
174. Reactions occur spontaneously when the heat of reaction is negative and randomness ("**entropy**") is positive; overall, therefore, a reaction is spontaneous when the value of ΔG is negative.
175. Organic chemistry is the study of compounds of carbon.
176. Organic compounds typically have low melting points.
177. Organic compounds are typically non-electrolytes.
178. Organic compounds are generally soluble with other organic (*i.e.* nonpolar) solvents (thus, <u>not</u> in water).
179. Carbon atoms form a total of 4 covalent bonds each, whether as single, double, or triple bonds.
180. Organic compounds that have only single bonds between carbons are said to be **saturated** compounds ("**alkanes**").
181. Organic compounds that have at least one double or triple bond between carbons are said to be **unsaturated** compounds ("alkenes," "alkynes," or dienes and "diynes," respectively).
182. **Benzene** is a cyclic organic compound; this molecule is the basis for a whole series of compounds called **aromatics**.
183. **Hydrocarbons** contain only carbon and hydrogen.
184. **Alcohols** are Organic compounds that contain at least one hydroxyl functional group (OH) attached to a carbon atom.
185. Alcohols do not behave like inorganic bases!
186. **Organic acids** contain an OH and an O attached to an end/terminal carbon (COOH / "carboxyl" functional group).
187. **Aldehydes** have an oxygen attached to an end/terminal carbon (C=O).
188. **Ketones** have an oxygen atom bonded, similar to an aldehyde, but attached to an <u>interior</u> rather than a terminal carbon.
189. **Ethers** have oxygen within the carbon chain (C-O-C) (form an "oxygen bridge").
190. **Esters** have an oxygen within the chain and an oxygen attached to the carbon on the chain (COO).
191. **Amines** have NH_2 attached to an end carbon.
192. Amino acids have an amine and an organic acid attached to the carbon chain.

193. **Substitution reactions**: a hydrogen atom is substituted for by another element [typically a halogen]; only possible in a saturated hydrocarbon.
194. **Addition reaction**: occur between an alkene or alkyne and another element [typically hydrogen or a halogen]. A multiple bond is broken and the hydrogen or halogen is added to the carbon(s); only possible in an unsaturated hydrocarbon.
195. **Fermentation** is a reaction between a sugar in the presence of an enzyme from yeast to produce an alcohol (*e.g.* C_2H_5OH) and carbon dioxide (CO_2).
196. **Combustion** reaction is the reaction between a hydrocarbon and oxygen to produce carbon dioxide and water, and releases a whole lot of energy!
197. **Esterification** is a reaction between an organic acid and an alcohol to produce an ester and water.
198. **Saponification** is a reaction between a fat and a base to produce soap and glycerol.
199. **Polymerization** is the creation of large molecules from several small molecules called **monomers** (starches, cellulose, nylon and proteins are examples of products from polymerization).
200. **Redox** reactions are the result of oxidation and reduction processes.
201. Losing electrons is **oxidation** (**LEO**).
202. Gaining electrons is **reduction** (**GER**).
203. The atom that *loses electrons* is called the **reducing agent** since it causes the other atom to gain electrons.
204. The atom that *gains electrons* is called the **oxidizing agent** since it causes the other atom to lose electrons.
205. Changes in an atom's oxidation number (charge) indicates the gain or loss of electrons by that atom.
206. A positive change (an increase) in oxidation numbers means *electrons were lost*.
207. A negative change (a decrease or **reduction**) in oxidation numbers means *electrons were gained*.
208. The number of electrons lost must be equal to the number gained.
209. Half-reactions show the changes between the atoms that are oxidized and reduced.
210. In a chemical cell the electrons spontaneously transfer from one substance to another, and electrical power (voltage) is produced.
211. Reduction occurs at the **cathode** (becomes more negative). **RED-CAT !**
212. Oxidation occurs at the **anode** (becomes more positive). **AN-OX !**
213. Electrolysis/electroplating occurs with the help of an outside energy source because the reaction is **nonspontaneous**.
214. Acids contain hydrogen (H^+) as the positive ion.
215. Bases contain hydroxide (OH^-) as the negative ion.
216. Acids can also be defined as **proton donors**; bases can also be defined as **proton acceptors**.
217. Acids turn litmus Red and have a pH less than 7; bases turn litmus blue and have a pH of greater than 7.
218. **Neutralization** is a reaction between and acid and a base to produce an ionic compound (a "salt") and water.
219. Salts are ionic compounds that contain positive ions other than hydrogen and negative ions other than hydroxide.
220. **pH** is the measure of the hydrogen ion concentration.
221. Nuclear chemistry is the study of reactions involving changes to the nucleus of an atom.
222. **Half-life** is the time required for a radioactive element isotope to decay to half its original value (see table N).
223. **Transmutation** is the change that occurs in the nucleus of an atom as it is turning into a new element.
224. **Alpha decay** is when an alpha particle is given off; alpha particles are the same as helium nuclei.
225. **Beta decay** is when a high-speed electron (beta particle) is given off.
226. **Gamma rays** are not particles, but are radiation similar to X-rays (they have the most penetrating power).
227. **Artificial transmutation** occurs when a nucleus is bombarded by some type of particles in order to force a transmutation to occur.
228. Fission is the splitting of atomic nuclei; smaller particles result and energy is released.
229. Fusion is the joining of atomic nuclei; larger particles result and energy is released.
230. Radioactive isotopes are used in medicine to help detect diseases, among other applications (*e.g.* radioactive iodine can help detect thyroid disorders *etc.*).

Study hard!

***Do not let the Chem* OWN YOU!**