About 100 THINGS to REMEMBER to to get a Better Grade on the Regents Exam

1. Protons are positively charged (+). Neutrons have no charge.

2. The number of electrons = the number of protons = the atomic number.

3. Electrons are small and are negatively charged (-).

4. Protons & neutrons are in an atom's nucleus.

5. Electrons are found in "clouds" or orbitals, flying around an atom's nucleus.

6. The mass number is equal to an atom's number of protons & neutrons added together.

7. The atomic number is equal to the number of protons in the nucleus of an atom- which is the same number of electrons.

8. The number of neutrons = mass number minus the atomic number.

9. Isotopes are chemically identical atoms with different numbers of neutrons.

10. Cations are positive (+) ions and form when a neutral atom loses electrons. Cations are smaller than their neutral atoms because they lose all of their valence electrons: whole orbitals

11. Anions are negative ions and form when a neutral atom gains electrons. Anions are larger than their neutral atoms because they end up with full valence orbitals, which contain lots of negative electrons that push on each other and stretch the valence orbital slightly bigger.

12. Dalton's model of the atom was a solid sphere of matter that was uniform throughout, a billiard ball. All property differences between atoms were due to different masses of the atoms (this is not accepted today)

12b. JJ Thompson discovered the electron and developed the "plum-pudding" model of the atom. He "planted" these electron particles in the positive atomic mass.

14. Ernest Rutherford's gold foil experiment showed that an atom is mostly empty space with a small dense, and + charged nucleus. Electrons fly around this nucleus. He shot alpha particles from radioactive Po at gold.

15. The Bohr Model of the atom placed electrons in "planet-like" orbits around the nucleus of an atom. It mathematically "proved" electrons existed in specific energy levels or orbits. Electrons could jump to higher orbits, temporarily, if they gained specific amounts of energy (quanta), and would release this energy as visible light when the electrons return from excited state to ground state.

16. The current, wave-mechanical model of the atom has electrons flying in clouds" (or orbitals) around the nucleus. The electrons are now thought to follow statistical rules, but locating a specific electron is trickier than ever.

17. USE THE REFERENCE TABLES when ever you can. Don't guess at formulas.

18. "STP" means "Standard Temperature and Pressure."
(273 Kelvin or 0°C & 1 atm or 101.3 kPa) which is on reference table "A"

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. Compounds are pure substances composed of two or more different kinds of atoms bonded together (ionic or covalent). Elements and compounds are homogeneous. Mixtures are NOT pure substances. Mixtures can be homogeneous or heterogeneous.

21. Binary compounds are substances made up of only two kinds of atoms. (examples: H_2O, HCl, CO)

22. Diatomic molecules are elements that form two atom molecules in their naturally occurring pure form at STP. Remember HONClBrIF twins. When bonding to other atoms any number of these atoms can bond ex: HCl, or SO_3.

23. Practice significant figures, handout on Arbuiso.com

24. Solutions are the best examples of homogeneous mixtures. They can be aqueous, gases, or even solids. (air, salt water, steel)

25. Heterogeneous mixtures have discernable components and are not uniform throughout. (Chocolate-chip cookies, vegetable soup, soil, muddy water)

26. A solute is the substance being dissolved, while the solvent is the substance that dissolves the solute. (Water is the solvent in saltwater, salt is the solute.)

27. Isotopes are written in a number of ways: C-14 is also Carbon-14, and is also like this:

$${}^{12}_{6}\mathrm{C} \qquad\qquad {}^{14}_{6}\mathbf{C}$$

(non-radioactive carbon 12 and radioactive carbon-14 both have 6 protons.)

Some isotopes are stable because the ratio of neutrons:protons in the nucleus is within the band of stability. When the neutron:proton ratio is outside this band of stability, the isotopes are unstable AKA: radioactive. They change their nuclear ratios by releasing particles and energy called radioactivity.

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

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

30. These configurations show the ground state for the electrons, they are in their lowest energy levels.

31. When they gain some energy (heat or electricity) the electrons move to higher orbitals. Example, sodium at a ground state of 2-8-1 might become 2-7-2. There are not MORE electrons, but they are in a higher than normal orbital

32. Polyatomic ions (Table E) are groups of atoms with an overall charge. Ex: NO_3^{-1}, NH_4^{+}, 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. Ionic compounds are neutral (all compounds are).

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.” KCl is potassium chloride while 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, and, NH_4NO_3 is ammonium nitrate.

37. Physical changes do not form new substances. They merely change the appearance of the original material. example: the melting of ice or the condensing of steam into liquid water

37a. Phase changes to know are melting/freezing, condensing/vaporizing, and sublimation/deposition

38. Chemical changes result in the formation of new substances. (synthesis or combination of H_2 gas and O_2 to produce water)

39. Reactants are on the left side of the reaction arrow and products are on the right.

40. Endothermic reactions absorb heat. The energy is written on the left (reactant) side of the reaction arrow in a forward reaction.

41. Exothermic reactions release energy and the energy is written as a product in the reaction. Energy is most often heat.

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: $2H_{2(g)} + O_{2(g)}$ ---> $2H_2O_{(g)}$

44. Decomposition reactions occur when a single reactant forms two or more products.
Example: $CaCO_{3(s)}$ ---> $CaO_{(s)} + CO_{2(g)}$

45. Single replacement reactions occur when one atom replaces an ion in an aqueous solution.
Example: $Mg_{(s)} + 2HCl_{(aq)}$ ---> $MgCl_{2(s)} + H_{2(g)}$

46. Double replacement reactions occur when two aqueous solutions react to form 2 new ionic compounds which may or many not be aqueous (check table F).
Example: $AgNO_{3(AQ)} + KCl_{(AQ)}$ ---> $AgCl_{(S)} + KNO_{3(AQ)}$
a solid must form or a double replacement reaction did not occur

47. Combustion, hydrocarbons combining with oxygen, the products are always $CO_2 + 2H_2O$
Example: $CH_4 + 2O_2$ ---> $CO_2 + 2H_2O$

47b. And acid base neutralization reactions always have an acid plus base making a salt & water. This is a "fancy" sort of double replacement reaction that neutralizes the acid and base. Sometimes, when a carbonate is acting as a base, there is a third product, carbon dioxide.

47c. The mass of the reactants in a chemical equation is always = to the mass of the products. “Law of Conservation of Mass.” Matter cannot be created or destroyed in any chemical reaction (or during a physical change ie: phase change.)

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 of a compound.

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. They hold their shape.

52. Solids have a definite shape and volume.

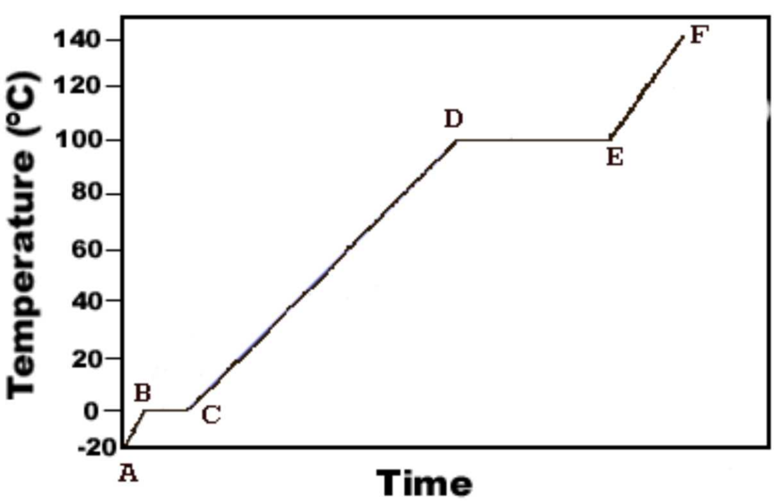
53. Liquids have closely-spaced particles that easily slide past one another.

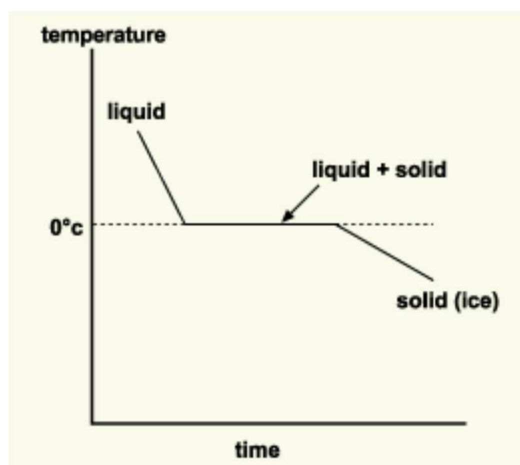
54. Liquids have no definite shape, but have a definite volume.

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 and cooling curves as pictured below.





57b. Temperature changes indicate kinetic energy changes. When heat is being added over time (or removed over time at right), and the temperature does not change, nor does the kinetic energy. During these "flat" areas on the graph, the potential energy is increasing (or decreasing at right). These flat areas of the graph are the PHASE changes of the substances

58. Substances that sublime turn from a solid directly into a gas.
(examples: $CO_{2(S)}$ & $I_{2(S)}$) The reverse is called deposition.

59. Kelvin = °C + 273

60. Use this formula to calculate heat absorbed/released by substances. $q = mC\Delta T$

q = heat absorbed or released (in Joules)

m = mass of substance in grams

C = specific heat capacity of substance (J/g·°C) … for water it's 4.18 J/g·°C

ΔT = temperature change in °C

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

63. As the pressure on a gas increases, the volume decreases. This is inversely proportional.

64. As the pressure on a gas increases, temperature increases. This is directly proportional.

65. As the temperature of a gas increases, volume increases. This is directly proportional.

66. Always use Kelvin for temperature when using the combined gas law. This law is on back page of the reference tables. Zero degrees, or negative degrees will wreck the math, so we use the absolute scale, when zero really means zero, and does not exist in reality.

67. Real gas particles have volume & are attracted to each other, and don't always behave ideally.

68. Real gases behave more like ideal gases at <u>low pressures & high temps</u>. They are least likely to clump into a liquid with these conditions.

68b. When comparing, the gas with the smallest particles would act most ideal

69. Distillation separates mixtures that have different boiling points. Fractional distillation is used to separate crude oil into various hydrocarbons.

70. Mixtures are not chemically combined, so physical means will separate the components from each other. Filtration separates mixtures of solids & liquids.

71. Chromatography can also be used to separate mixtures of liquids, or also mixtures of gases. Always use differences in physical properties to separate mixtures from each other.

72. The Periodic Law states that the properties of elements repeat themselves PERIODICALLY in the groups, which means similar properties show up in the vertical groups.

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 to the right of it, and metalloids border this staircase line. Al & Po are exceptions to the metalloid line, they are pure metals. Hydrogen is a non metal, located on the "wrong side" of the stairs too.

76. Memorize this chart.

metals are...	malleable	ductile	lustrous	good conductors of heat & electricity	form (+) ions
non-metals are...	brittle solids	often gases at STP	dull	good insulators	form (–) ions

77. Noble gases (Group 18) are almost all inert (make no compounds) because their valence orbital of electrons is completely filled.

78. First 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 with the same number of orbitals. Nuclear charge is always +, and always the atomic number. Ex: the net nuclear charge of mercury is +80.

80. Atomic radii increase as you go down a group (more orbitals in each period).

81. Electro-negativity is a measure of an element's attraction for electrons.

82. Electro-negativity increases as you go to the right on the Periodic Table (F has the highest 4.0)

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.

86b. ALL of the inner transitional metals are in GROUP 3 only, periods 6+7. Look at the periodic table hanging in the class for a better view of this.

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. When bonds form, energy is released. (exothermic)

89. When bonds are broken it uses energy (endothermic).

90. Draw one dot for each valence electron when drawing an element's or ion's Lewis diagram.

91. Lewis dot diagrams for IONS require BRACKETS $[\ X]^{+2}$ with charges outside the brackets.

91b. There are NO unknown elements. If a question says an unknown element with 5 total electrons needs to be drawn, only one element in the universe has 5 total electrons, look it up on the periodic table and draw it correctly (three dots, not five. Boron has a 2-3 electron configuration)

92. Metallic bonds can be imagined as a crystalline lattice of cations surrounded by a "sea" of mobile valence electrons. These loose electrons give metals all of their metallic characteristics (ability to conduct electricity, malleable, ductility).

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 <u>share a pair of electrons</u>. One pair of electrons being shared is a single covalent bond.

95. Ionic bonds form when cations and anions are attracted to each other because of the <u>transfer of electrons</u> from cations to anions. Opposite charges attract.

96. Non-polar covalent bonds form when two atoms of the same element bond together.

97. Polar covalent bonds form when the electro-negativity difference between two bonding atoms is more than zero.

98. Ionic bonds form when a metal cation bonds to a non-metal anion (or polyatomic ions).

99. Substances containing mostly covalent bonds are called molecular substances.

99b. Covalent bonds can be single pair sharing, double pair sharing, or even triple pair sharing. NO QUADRUPLE BONDS, the atoms literally can't bend that way. Ionic bonds do not share any electrons, the electrons are transferred, so they can never be single, double, or triple.

100. Substances containing ionic bonds are called ionic compounds.

101. NH_4Cl is an exceptional ionic compound. It might be one of the only ones you will see without a metal cation. Ammonium cations are ionic even though they are made up of only non-metals.

102. K_3PO_4 is a compound that is ionic, but also has covalent bonds holding the phosphate together. It has BOTH ionic and covalent bonds.

103. When electrons return from the excited state to the ground state they emit the energy they gained to jump up in the first place. This energy is given off as visible light called spectra.

104. Every atom and every molecule and every formula unit gives off a unique spectra, and these spectra can be used to compare unknown substances against known spectra to determine what substances are present.