Two hundred things to know to pass the chem regents exam

Big Chem Outline:
U01: Scientific Skills
U02: Matter
U03: Atomic Structure
U04: Periodic Table
U05: Bonding
U06: Reactions
U07: Energy and phases
U08: Gases
U09: Kinetics + Equilibrium
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U13: Organic
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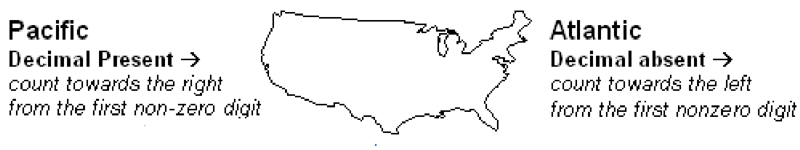
While multiple versions of this document exist, this one is presented with an eye towards **information design**. Moreover, its also designed to present the '200 facts' in a way consistent with the chronology of **Big Chem**.

The Natural Philosophers wish to give credit to this document's original author, however, none has been determined.

Unit 1: Scientific Skills

To determine the number of significant digits in a number:



Unit 2: Matter

*Elements are pure substances composed of only one kind of atom.

* Heterogeneous mixtures have discernable components and are not uniform throughout. Examples include: Chocolate-chip cookies, vegetable soup, soil, and muddy water.

*Physical changes do not form new substances. They change the appearance of the original material. Example: ice melting.

* Chemical changes result in the formation of new substances. Example: burning hydrogen gas to produce water vapor.

*Distillation separates mixtures with different boiling points.

*Filtration separates mixtures of solids and liquids.

*Chromatography can also be used to separate mixtures of liquids and mixtures of gases.

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Unit 3: Atomic Structure

* Protons are positively charged (+).

* Neutrons have no charge.

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

* Protons & neutrons are in an atom’s nucleus (nucleons).

* Electrons are found in “clouds” (orbitals) around an atom’s nucleus.

* The mass number is equal to an atom’s number of protons and neutrons added together.

* The atomic number is equal to the number of protons in the nucleus of an atom.

* The number of neutrons = mass number – atomic number.

* Isotopes are atoms with equal numbers of protons, but differ in their neutron numbers.

* Cations are positive (+) ions and form when a neutral atom loses electrons. They are smaller than their parent atom.

* Anions are negative ions and form when a neutral atom gains electrons. They are larger than their parent atom.

* Ernest Rutherford’s gold foil experiment showed that an atom is mostly empty space with a small, dense, positively-charged nucleus.

* J.J. Thompson discovered the electron and developed the “plum-pudding” model of the atom.

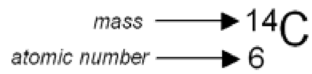
* Dalton’s model of the atom was a solid sphere of matter that was uniform throughout.

* The Bohr Model of the atom placed electrons in “planet-like” orbits around the nucleus of an atom.

* The current, wave-mechanical model of the atom has electrons in “clouds” (orbitals) around the nucleus.

* 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.

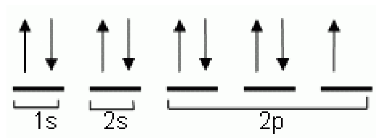
* Isotopes are written in a number of ways: C-14 is also Carbon-14, and is also like this:



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

* Electron configurations are written in the bottom center of an element’s box on the periodic table in your reference tables. Example: Calcium has an electron configuration of 2-8-8-2. This means it has 2 electrons in the first principal energy level, 8 in the second principal energy level, 8 electrons in the third principal energy level, and 2 electrons in the third principal energy level.

* Orbital notation is a way of drawing the electron configuration of an atom. Fluorine's orbital notation looks like this:



Each arrow represents an electron spinning on one direction. The adjecnt arrow represents another electron spinning in the opposite direction. Each spinning electron makes a magnetic field that's opposite the other's, and together they lock together like magnets.

* Polyatomic ions (Table E) are groups of atoms with an overall charge.

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

* The kernel of an atom includes everything in an atom except the atom’s valence electrons.

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Unit 4: Periodic Table

* The Periodic Law states that the properties of elements are periodic functions of their atomic numbers.

* Periods are horizontal rows on the Periodic Table.

* Groups are vertical columns on the Periodic Table.

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

*General properties of metals: malleable, ductile, lustrous, good conductors of heat & electricity, low ionization energy low electronegativity, tend to form positive ions.

General properties of nonmetals: brittle when solid, are mostly gases at STP, dull, good insulators, high ionization energy and electronegativity, tend to form negative ions. Noble gases (Group 18) are inert and stable due to the fact that their valence level of electrons is completely filled.

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

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

* Atomic radii increase as you go down a group.

* Electronegativity is a measure of an element’s attraction for electrons.

* Electronegativity increases as you go up and to the right on the Periodic Table.

* The elements in Group 1 are the alkali metals.

* The elements in Group 2 are the alkaline earth metals.

* The elements in Group 17 are the halogens.

* The elements in Group 18 are the noble gases.

* With the exception of groups 3 - 12, the last digit of an element’s group number is equal to its number of valence electrons.

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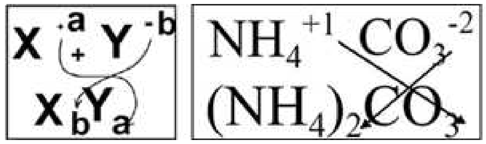
Unit 5: Bonding

* Binary compounds are substances made up of only two kinds of atoms.

*Chemical formulas are written so that the charges of cations and anions neutralize one another. Aka, the crossover method. All oxidation states must add to zero. Example: ammonium carbonate:



* 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: potassium chloride, magnesium sulfide.

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

* Metallic bonds can be thought of as a crystalline lattice of kernels surrounded by a "sea" of mobile valence electrons.

* Atoms are most stable when they have 8 valence electrons (an octet) and tend to form ions to obtain such a configuration of electrons.

* Covalent bonds form when two atoms share a pair of electrons.

* Ionic bonds form when one atom transfers an electron to another atom when forming a bond with it.

* Nonpolar covalent bonds form when two atoms of the same element bond together.

* Polar covalent bonds form when the electronegativity difference between two bonding atoms is between 0.4 and 1.7.

* Ionic bonds form when the electronegativity difference between two bonding atoms is greater than 1.7.

* Substances containing mostly covalent bonds are called molecular substances.

* Substances containing mostly ionic bonds are called ionic compounds.

* Ionic substances tend to be: hard, high melting and boiling points, conduct electricity when molten or when they are in aqueous solutions.

* Covalent (Molecular) substances tend to be: soft, low melting and boiling points, do not conduct electricity (are good insulators).

* Hydrogen bonds form when hydrogen bonds to the elements N, O, or F and gives the compound unusually high melting and boiling points.

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Unit 6: Reactions

* The seven diatomics: they only exist in pairs of themselves at STP. "BrINClHOF" (Br2, I2, N2, CL2, H2, O2, F2).

* 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.

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

*Only coefficients can be changed when balancing chemical equations!

* Synthesis reactions occur when two or more reactants combine to form a single product: A + B -> AB

* Decomposition reactions occur when a single reactant forms two or more products: AB -> A + B

* Single replacement reactions occur when one element replaces another element in a compound: A + BC -> AC + B

* Double replacement reactions occur when two compounds react to form two new compounds: AB + CD -> AD + CB

* Law of Conservation of Mass - the masses of the reactants in a chemical equation is always equal to the masses of the products. Its why equations are balanced.

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

* Know how to calculate the percentage composition of a compound using the following formula. From the reference tables, Table T:

$$\%\text{ composition by mass} = \frac{\text{mass of part}}{\text{mass of whole}} \times 100$$

* Avogadro's number is 6.02 E 23 and is the number of particles in 1 mole of a substance.

*Energy is released when a chemical bond forms. The more energy that is released, the more stable the bond is.

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Unit 7: Energy, Phses of Matter

* Endothermic reactions absorb heat. The energy value is on the left side of the reaction arrow in a forward reaction.

* Exothermic reactions release energy and the energy is a product in the reaction.

* The particles in a solid are rigidly held together.

* Solids have a definite shape and volume.

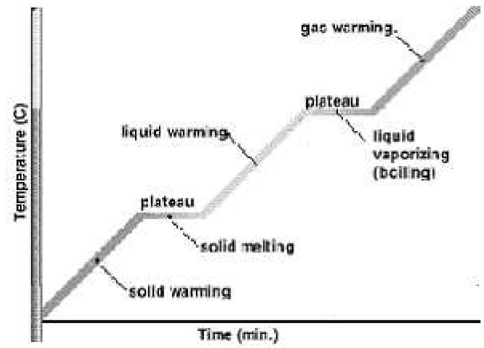
* Liquids have closely-spaced particles that easily slide past one another.

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

* Gases have widely-spaced particles that are in random motion.

* Gases are easily compressed and have no definite shape or volume.

* Be able to read and interpret the heating/cooling curve as pictured below:



* Substances that sublime turn from a solid directly into a gas.

* Degrees Kelvin = ℃ + 273

*Be able to calculate heat absorbed/released by substance using: q = m * c * (final temp - initial temp)
q = heat absorbed or released (Joules), m = mass of substance in grams, c = specific heat (4.18 for water) , temperature must be in degrees Celsius

* 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, on reference table).

* 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, on reference table).

* Liquids boil when their vapor pressure is equal to the atmospheric pressure.

* The normal boiling point of a substance is the temperature at which it boils at 1 atm of pressure. (Take note of Table H)

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Unit 8: Gases

* “STP” means “Standard Temperature and Pressure.” (273 Kelvin & 1 atm).

* As the pressure on a gas increases, the volume decreases proportionally.

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

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

* Always use Kelvins for temperature when using the combined gas law: (P1V1)/T1 = (P2V2)/T2

* Real gas particles have volume and are attracted to one another, and thus do not always behave like ideal gases.

* Real gases behave more like ideal gases at low pressures and high temperatures.

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Unit 9: Kinetics and Equilibrium

* Covalently bonded substances tend to react more slowly than ionic compounds.

* Increasing the concentration of reactants will increase reaction rate.

* Increasing the surface areas of the reactants will increase reaction rate.

* Increasing the pressure on gases increases reaction rate.

* Catalysts speed up reactions by lowering their activation energies. They are not changed themselves and can be reused many times over.

* Increasing temperature increases reaction rate.

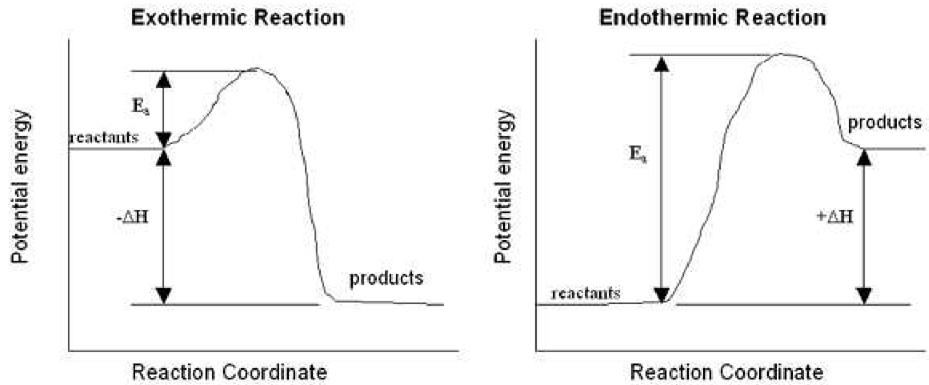
* DH is (+) for endothermic reactions and is (-) for exothermic reactions.

* The rates of the forward and reverse reactions are equal at equilibrium.

* Adding any reactant or product to a system at equilibrium will shift the equilibrium away from the added substance.

* Removing any reactant or product from a system at equilibrium will shift the equilibrium point toward that removed substance.

* Be able to recognize and read potential energy diagrams:



* An increase in temperature shifts an equilibrium system in the endothermic direction.

* A decrease in temperature shifts an equilibrium system in the exothermic direction.

* Increasing the pressure on a gaseous equilibrium will shift the equilibrium point toward the side with fewer moles of gas.

* Decreasing the pressure on a gaseous equilibrium will shift the equilibrium point toward the side with more moles of gas.

* Catalysts have no effect on an equilibrium. It just establishes itself quicker.

* Enthalpy (H) is the heat energy gained or lost in a reaction.

* Entropy (S) is high in a highly unorganized system, such as a gas, a messy room, etc.

* For the hypothetical reaction: wA + xB ->yC + zD, Keq is expressed by:

$$K_{eq} = \frac{[C]^y[D]^z}{[A]^W[B]^x}$$

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Unit 10: Solutions

* Solutions are the best examples of homogeneous mixtures. (Air, salt water, etc.)

* 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.)

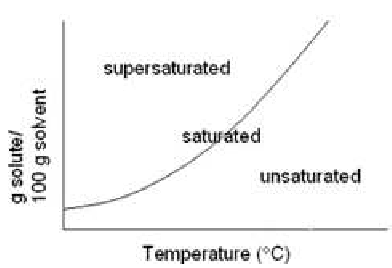
* Use Table F to predict the solubilites of compounds.

* Remember substances tend to be soluble in solvents with similar bond types. Polar solutes dissolve in polar solvents. Nonpolar solutes dissolve in nonpolar solvents. “Like dissolves like”

* As temperature increases, solubility increases for most solids.

* At low temperatures and high pressures solubility increases for most gases.

* Use Table G to determine whether a solution is saturated, unsaturated, or supersaturated. Here's how to interpret it:



* 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.

* Percent by mass = mass of the part / mass of the whole x 100%

* Parts per million (ppm) = grams of solute / grams of solution x 1,000,000

* Solutes raise the boiling points and lower the melting points of solvents.

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Unit 11: Acids and Bases

* Acids and bases are both good electrolytes. Their solutions conduct electricity well.

* Weak acids taste sour.

* Weak bases taste bitter.

* Acids and bases turn indicators different colors. They're listed on Table M.

* Acids have a pH < 7.

* Bases have a pH > 7.

* Tables K & L list names and formulas of common acids and bases asked about on the Regents.

* The metals above H_2 on Table J will react with acids to make H_2 gas bubbles.

* Arrhenius says: "Acids give off H^+ or H_3O^+ ions in solution." "Bases give off OH^- ions in solution."

* Brønsted says: "Acids donate protons." "Bases accept protons."

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

* Titrations are controlled neutralization reactions used to find the concentration of an acid or base sample. Note the formula for it on Table T.

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Unit 12: Oxidation and Reduction

* 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.

* 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.

* Redox reactions always involve the exchange of electrons.

* Remember…. "LEO says GER!", Lose electrons oxidation, gain electrons reduction.

* Identify redox reactions by seeking an uncombined element on one side of a reaction that is in a compound on the other side: $Zn + 2HCl \rightarrow ZnCl_2 + H_2$, Uncombined Zn is combined with Cl.

* Oxidizing agents are what get reduced in a redox reaction. Reducing agents are what get oxidized in a redox reaction.

* Electrochemical cells produce electricity with a spontaneous redox reaction.

* The left electrode is usually the site of oxidation in an electrochemical cell diagram.

* Memorize this saying… "I have AN OX and a RED CAT."

* In electrochemical cells, the ANode gets OXidized and REDuction occurs at the CAThode.

* Electrolytic cells use an applied electrical current to force a nonspontaneous redox reaction to take place.

* Electrolytic cells are usually used for metal plating of objects.

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Unit 13: Organic Chemistry

* All organic compounds contain the element carbon.

* Carbon ALWAYS makes four bonds in molecules.

* Saturated hydrocarbons have all single bonds within them (alkanes).

* Unsaturated hydrocarbons have double or triple bonds in them (alkenes & alkynes).

* Hydrocarbons contain ONLY the elements hydrogen and carbon.

* The homologous series of hydrocarbons' formulas are on Reference Table Q.

* The functional groups on organic molecules are listed on Reference Table R.

* Structural isomers of organic compounds have different structural formulas but the same molecular formula.

* Number the parent carbon chain in an organic molecule from the end closest to the alkyl group(s).

* Combustion reactions occur when a hydrocarbon reacts with oxygen to make CO_2 and H_2O.

* 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.

* Organic addition reactions occur when an alkene or alkyne combine with a halogen to make one product (halide).

* Esterification occurs when an organic acid and an alcohol react to make water and an ester.

* Saponification occurs when an ester reacts with a base to make alcohol and a soap.

* Fermentation reactions occur when yeast catalyze a sugar ($C_6H_{12}O_6$) to make carbon dioxide and ethanol.

* Polymers are long chains of repeating units called monomers.

* Polymers form by polymerization reactions.

* Addition polymerization occurs when unsaturated monomers join in a long polymer chain.

* Condensation polymerization occurs when monomers join to form a polymer by removing water. Water is a product!

* Natural polymers include starch, cellulose, and proteins. Synthetic polymers include plastics such as nylon, rayon, and polyester.

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Unit 14: Nuclear Chemistry

* Unstable atoms that are radioactive are called radioisotopes. (Table N)

* Radioisotopes can decay by giving off any of the particles/emanations listed in Table J.

* Alpha particles (see Table J) are positively charged (+).

* Beta particles (see Table J) are negatively charged (-).

* 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)

* C-14 is used to determine the ages of organic material up to 23,000 years old.

* U-238 is used to determine the ages of rocks.

* I-131 is used to treat thyroid disorders.

* Co-60 is used to treat cancer tumors.

* Radiation can be used to kill bacteria on foods to slow the spoilage process.

* Disposal of radioactive waste is a problem associated with nuclear reactors.

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

* Fission reactions split heavy nuclei into smaller ones.

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

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