Name: ____

What You *Need* To Know for the Chemistry Regents Exam

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What is the point of this packet?
This review packet was assembled from NY State's Core Curriculum, which outlines the material to be tested on the Regents exam. This is by no means a thorough review of the entire course. It is designed to be used with review sheets, past Regents exams and your Reference tables to help you prepare for the coming test. Emphasis is placed on key ideas that are stressed by the Core Curriculum. Additional space has been left for you to add your own notes.

You cannot passively prepare for the Chemistry Regents.

There are no shortcuts.

You must study, ask questions, analyze problems and come to review sessions- with questions- to be thoroughly prepared.

Be Prepared
You should come to the exam with a 4-function or scientific calculator (graphing calculators are not allowed), pen, and pencil. Reference Tables will be provided. You are required to stay in the examination room for a minimum of 2 hours from the time the test is distributed.

Exam Date: ____

Chemistry

Tips For Success

1. **Spread out your study time.**
2. **Know your reference tables.**
3. **Attend review sessions.**
4. **Ask questions.**
5. **Get a good night's sleep.**
6. **Have a healthy meal.**
7. **Arrive early. Arrive prepared.**
8. **Relax.**
9. **Read every question carefully.**
10. **Read every question again.**

The Test

This analysis is based on the exams given through January 2019. It is possible new exams will have minor modifications.

The Chemistry Regents Exam is broken down into three sections:

Part A: 30 multiple choice questions from all units covered over the course of the school year.

Part B: 35 questions, 20 multiple choice and 15 short answer. Part B focuses on the Reference Tables, graphing, and laboratory experiments.

Part C: 20 short answer questions. This is often an eclectic mix from various units, and may demand students write short paragraphs, use equations and reference tables, or draw graphs and diagrams to receive credit.

The test is usually 85 questions worth a total of 85 points. Most questions are worth 1 point. The grading scale (i.e. curve) changes with each test, with a raw score of 50/85 typically earning a passing grade of 65, and 74/85 typically earning a mastery score of 85. The actual scoring scale varies with each exam, and there is no way to know what the curve will before the exam begins. A sample scale is included on the last page for your reference.

Topics

There are 12 specific chemistry topics covered on the test, as well as a 13th topic on general lab skills:

<u>**Topic One: The Atom**</u>

1. The atomic model has evolved over many years through the work of many different scientists.

- ✓ *Dalton's Model:*
  - Elements are made of atoms
  - Atoms of an element are the same.
  - Compounds are formed from combinations of atoms.
  - Atoms combine in fixed proportions to form compounds.
- ✓ *Rutherford Experiment*
  - Bombarded gold foil with alpha particles (see Table O).
  - Showed atoms were
    - mostly empty space
    - small, dense nucleus
    - the nucleus has a positive charge
- ✓ *Bohr Model*
  - Small, dense, positively charged nucleus surrounded by electrons in <u>circular</u> orbits
  - Electrons can jump from lower energy levels to higher energy levels and back, depending on the energy lost or gained.
- ✓ *Wave-Mechanical Model (Modern Atomic Theory)*
  - Small, dense, nucleus positively charged nucleus surrounded by electrons moving in "electron cloud".
  - "Orbitals" are areas where an electron with a certain amount of energy is *most likely* to be found.

2. Each atom is made of a positively charged nucleus with one or more orbiting, negatively charged electrons.

3. Protons and neutrons are found in the nucleus.

4. Atoms of the same element all contain the same number of protons.

- ✓ Changing the number of protons changes the atom into a different element.
- ✓ The atomic number is the number of protons in an atom of an element.

5. Protons (Table O)

- ✓ Found in the nucleus
- ✓ Have a charge of +1
- ✓ Mass = 1 amu
- ✓ Number of protons = Atomic number of an atom
- ✓ Number of protons = Nuclear charge (charge of the nucleus)
- ✓ Changing the number of protons changes an atom into a different element (transmutation – see Nuclear Chemistry on p7)

6. Neutrons (Table O)

- ✓ Found in the nucleus
- ✓ Have a charge of zero – they are neutral!
- ✓ Mass = 1 amu
- ✓ Number of neutrons = Mass number of an atom – atomic number
- ✓ Changing the number of neutrons makes a different isotope of the same element (such as changing hydrogen-2 to hydrogen-3)

7. Electrons (Table O)

- ✓ Found in orbitals around the nucleus
- ✓ Have a charge of -1
- ✓ Have a very, very small mass (essentially 0 amu)
- ✓ Number of electrons = Atomic number – ionic charge
- ✓ Changing the number of electrons creates an <u>ion</u> (a charged atom).
- ✓ Used in the formation of bonds between two atoms.
- ✓ Also called a Beta (-) particle

8. Masses

- ✓ The mass of a proton is 1 amu. The mass of a neutron is 1 amu. The mass of an electron is almost 0 amu.
- ✓ The amu is defined as 1/12 the mass of an atom of Carbon-12.
- ✓ The <u>mass number</u> of an atom is equal to the total number of protons and neutrons.
- ✓ The <u>atomic mass</u> of an atom is a <u>weighted average</u> of the masses of all naturally occurring isotopes.

9. Isotopes

- ✓ <u>Isotopes</u> are atoms with equal numbers of protons but different numbers of neutrons.
- ✓ Isotopes of an element have the same atomic number (protons only), but different atomic masses (protons + neutrons).

9. Charges and Ions

- ✓ Atoms are neutral (no charge) because the number of protons in an atom equals the number of electrons (the positives and negatives cancel each other out). The neutrons have no affect on charge.
- ✓ An atom with a charge is called an <u>ion</u>.
- ✓ Ions are created only when an atom gains or loses an electron. The number of protons do not normally change.
- ✓ Ions have unequal numbers of protons and electrons.
- ✓ A <u>cation</u> is a positive ion. It has lost electrons.
  - Cations have a smaller radius than the original atom.
- ✓ An <u>anion</u> is a negative ion. It has gained electrons.
  - Anions have a larger radius than the original atom.

9. Electrons and Energy

- ✓ <u>Ground State:</u> All electrons are at their lowest possible energy level.
- ✓ Electrons fill in energy levels and orbitals starting with the one that requires the least energy (1s) and progressively fill in those levels and orbitals that require more energy.
- ✓ When an electron gains a specific amount of energy, it moves to a higher orbital and is in the "<u>excited state</u>".
- ✓ When an electron returns from a higher energy state to a lower energy state, it emits energy in the form of <u>light</u>. This can be used to identify an element using a <u>flame test</u> or a <u>bright line spectrum</u>.
- ✓ <u>Spectroscope:</u> An instrument that separates the visible light emitted during a flame test into a bright line spectrum.

10. Valence Electrons

- ✓ **The outermost electrons are called <u>valence electrons</u>.**
- ✓ **Valence electrons determine the chemical properties of the element.**
- ✓ Atoms with a filled valence level are stable (like the noble gases).
- ✓ Most elements can have up to 8 electrons in their valence level. The exceptions are H and He, which can have only 2 valence electrons.
- ✓ Atoms form bonds in order to fill their valence levels.
- ✓ You can use <u>Lewis structures</u> (electron dot diagrams) to show the configuration of the valence electrons.

14. When an atom gains an electron, it becomes a negative ion and its radius increases.

15. When an atom loses an electron, it becomes a positive ion and its radius decreases.

16. Electron-dot diagrams (Lewis structures) represent the valence electron configuration in elements, compounds and ions.

- ✓ Electrons in Lewis structures are arranged by their orbitals.
- ✓ The first two electrons are placed together in the "s" orbital.
- ✓ The remaining electrons are spread among the 3 "p" orbitals.
- ✓ The "s" orbital must be filled first. Then each "p" orbital must have one electron before another "p" orbital gains a second.

17. <u>Electronegativity</u> measures how strongly an atom attracts electrons in a chemical bond. These values are based on an arbitrary scale.

- ✓ Fluorine has the highest electronegativity of all elements (4.00).
- ✓ Electronegativites can be found on Table S.
- ✓ Nonmetals with a high electronegativity are likely to <u>take</u> electrons, and are very reactive.
- ✓ Metals with a low electronegativity are likely to <u>lose</u> electrons and are also very reactive.
- ✓ Elements with mid-range electronegativities are less reactive.
- ✓ Electronegativity is also used to determine the polarity of a covalent bond (see Bonding on p9)

18. First Ionization Energy measures how much energy is needed to remove the first electron from an atom.

- ✓ Nonmetals tend to have high ionization energies, and are unlikely to lose their electrons.
- ✓ Metals tend to have low ionization energies so it is easy for them to lose their electrons.

<u>Topic Two: The Periodic Table</u>

1. The placement of an element on the Periodic Table gives an indication of the chemical and physical properties of that element.

2. Elements are arranged in order of increasing atomic number.

3. The number of protons in an atom (atomic number) identifies the element.

- ✓ The number of protons in an atom only changes through nuclear reactions.

4. The mass number is the sum of protons and neutrons in the nucleus.

- ✓ The atomic mass given on the periodic table is a <u>weighted average</u> of the different isotopes of that element (see Atoms p4).
- ✓ Electrons do not significantly add to the atomic mass.

5. Isotopes of an element are identified by the sum of protons and neutrons.

- ✓ Isotopes of the same element have the same number of protons and a different number of neutrons.
- ✓ Isotopic notation shows the mass number (top) and atomic number (bottom) of an isotope: $^{14}_{6}C$

6. Elements can be classified by their properties and their location on the Periodic Table as metals, non-metals, metalloids, and noble gases.

7. Elements may be differentiated by their physical properties.

- ✓ Ex: Density, conductivity, malleability, hardness, ductility, solubility
- ✓ The formula to calculate density is found on Table T.

8. Elements may be differentiated by their chemical properties.

- ✓ Chemical properties describe how an element behaves in a chemical reaction.

9. Elements are arranged into periods and groups.

10. Elements of the same period have the same number of occupied energy levels.

11. Elements of the same group have the same valence configuration and similar chemical properties.

- ✓ Group 1 elements other than H are *alkali metals*.
- ✓ Group 2 elements are *alkali earth metals*.
- ✓ Group 17 elements are *halogens.*
- ✓ Alkali metals, alkali earth metals, and halogens all are highly reactive and do not exist as free elements in nature (they are all found in compounds).
- ✓ Group 18 elements are *noble or inert gases*. These elements have filled valence levels and are do not normally react with other substances.

12. The succession of elements within a group demonstrates characteristic trends in properties. As you progress *down* a group:

- ✓ atomic radius increases.
- ✓ electronegativity decreases.
- ✓ first ionization energy decreases.
- ✓ metallic character increases.

13. The succession of elements within a period demonstrates characteristic trends in properties. As you progress *across* a group from *left to right*:

- ✓ atomic radius decreases.
- ✓ electronegativity increases.
- ✓ first ionization energy increases.
- ✓ metallic character decreases.

<u>**Topic Three: Bonding**</u>

1. Chemical bonds are created by a strong attraction between atoms.

- ✓ The types of bonds formed depends on the types of atoms involved.
- ✓ The way atoms bond will determine the structure of their molecules.
- ✓ The structure of molecules will determine the properties of a substance.

2. Forming or breaking bonds always results in a change of energy.

- ✓ Forming a chemical bond releases energy.
- ✓ Breaking a chemical bond absorbs energy.

3. Two major categories of compounds are ionic and molecular compounds.

4. Compounds can be differentiated by their chemical and physical properties.

- ✓ Ionic substances have high melting and boiling points, form crystals, dissolve in water (<u>dissociation</u>), and conduct electricity in solution and as a liquid.
- ✓ Covalent or molecular substances have lower melting and boiling points, do not conduct electricity.
- ✓ Polar substances are dissolved only by another polar substance. Non-polar substances are dissolved only by other non-polar substances.

5. Atoms gain a stable electron configuration by bonding with other atoms.

- ✓ Atoms are stable when they have a full valence level.
- ✓ Most atoms need <u>8 electrons</u> to fill their valence level.
- ✓ H and He only need <u>2 electrons</u> to fill their valence level.
- ✓ The noble gases (group 18) have filled valence levels. They do not normally bond with other atoms.

6. Chemical bonds are formed when valence electrons are:

- ✓ Transferred from one atom to another – <u>ionic.</u>
- ✓ Shared between atoms – <u>covalent</u>.
- ✓ Mobile in a free moving "sea" of electrons – <u>metallic</u>.

7. Bonding guidelines:

- ✓ Metals form metallic bonds.
- ✓ Metals react with nonmetals to form ionic compounds.
- ✓ Nonmetals bond with nonmetals to form covalent compounds (molecules).
- ✓ Ionic compounds with polyatomic ions have both ionic and covalent bonds.

8. In multiple (double or triple) covalent bonds more than 1 pair of electrons are shared between two atoms.

- ✓ oxygen and its family (group 16) form double bonds with each other (O_2)
- ✓ nitrogen and its family (group 15) form triple bonds with each other (N_2)
- ✓ carbon can form double and triple bonds with itself, and group 15 and 16 elements (ex: CO_2)

9. Polarity of a molecule can be determined by its shape and the distribution of the charge. (SNAP rule)

- ✓ Polar molecules must have polar bonds.
- ✓ Polar molecules are asymmetrical.
- ✓ Nonpolar molecules are symmetrical and/or have no polar bonds.

10. The electronegativity difference between two bonded atoms can determine the type of bond and its polarity.

- ✓ The greater the electronegativity difference between two atoms, the more polar the bond is.

11. Intermolecular forces allow different particles to be attracted to each other to form solids and liquids.

- ✓ <u>Hydrogen bonds</u> are an example of a strong IMF between atoms.
- ✓ Hydrogen bonds exist between atoms of hydrogen and oxygen, fluorine, or nitrogen.
- ✓ Substances with hydrogen bonds tend to have much higher melting and boiling points than those without hydrogen bonds.

12. Physical properties of a substance can be explained in terms of chemical bonds and intermolecular forces. These include conductivity, malleability, solubility, ductility, hardness, melting point and boiling point.

13. Some substances may exist in two or more forms. These forms differ in the way they are bonded, which changes their structure and their properties.

- ✓ Ex: Carbon exists as both graphite and diamond (a network solid). Oxygen can exist as $O_2(g)$ (oxygen gas) or $O_3(g)$ (ozone gas).

<u>**Topic Four: Moles, Reactions, and Stoichiometry**</u>

1. A compound is a substance composed of two or more different elements that are chemically combined in a fixed proportion. A chemical compound can only be broken down by chemical means.

2. Chemical compounds can be represented by a specific formula and assigned a name based on the IUPAC system.

3. Types of chemical formulas include empirical, molecular, and structural.

- ✓ Empirical formulas show elements in their simplest whole number ratios. This may or may not be the same as the molecular formula.
- ✓ Molecular formulas show the actual number of atoms per element in a single molecule.
- ✓ Structural formulas show the number of each type of atom as well as their physical arrangement.

4. All chemical reactions conserve mass, atoms, charge and energy.

5. A balanced chemical equation represents conservation of atoms.

6. The coefficients in a balanced chemical equation can be used to determine <u>mole ratios</u> in the reaction.

7. The formula mass of a substance is the sum of the atomic masses of its atoms. The molar mass (gram formula mass) equals the mass of one mole of that substance.

8. The percent composition by mass of each element in a compound can be calculated mathematically.

9. Types of chemical reactions include synthesis, decomposition single replacement, and double replacement.

10. The mole formula can be found on Table T.

Topic Five: Energy

1. Energy can exist in different forms – chemical, electrical, electromagnetic, thermal, mechanical, nuclear.

- ✓ Heat is energy that flows from one substance to another based on a difference in temperature.
- ✓ Kinetic energy is the energy of motion.
- ✓ Potential energy is stored energy. The potential energy stored in chemical bonds is chemical energy.

2. Law of Conservation of Energy: energy cannot be lost or destroyed, only changed from one form to another or transferred between objects.

4. Temperature is a measure of the average kinetic energy of the particles in a sample. Temperature is NOT the same as heat!

5. Processes that are exothermic give off heat energy. This typically causes the surrounding environment to become warmer.

6. Processes that are endothermic absorb energy. This typically causes the surrounding environment to become colder.

7. Heat of reaction (ΔH) tells how much energy is gained or lost in a chemical reaction (Table I)

- ✓ $+\Delta H$ = endothermic
- ✓ $- \Delta H$ = exothermic

8. Heat of fusion (H_f) is the energy needed to change 1.0g of a substance from solid to liquid.

- ✓ The H_f formula can be found on Table T.
- ✓ The heat of fusion of water is 334 J/g (Table B)

9. Heat of vaporization (H_v) is the energy needed to convert 1.0g of a substance from liquid to gas.

- ✓ The H_v formula can be found on Table T.
- ✓ The heat of vaporization of water is 2260 J/g (Table B)

10. Specific heat capacity (C) is the energy needed to raise 1.0g of a substance 1^{O}C.

- ✓ The specific heat capacity of liquid water is 4.18 J/g*^{O}C (Table B).
- ✓ The specific heat formula can be found on Table T.

Topic Six: Matter

1. Matter is classified as a pure substance or a mixture of substances.

- ✓ A substance has fixed composition and uniform properties throughout the sample. Element and compounds are substances.
- ✓ A substance MUST be homogeneous.

2. A mixture is composed of two or more different substances that may be physically separated.

- ✓ A mixture may be homogeneous (uniform – a solution), or heterogeneous (uneven).
- ✓ Substances in a mixture retain their original properties.
- ✓ Substances in a mixture may be separated by their size, polarity, density, boiling and freezing points, and solubility (among others).
- ✓ Filtration and distillation are examples of processes used to separate mixtures.

2. An element is a substance composed of atoms with the same atomic number. They cannot be broken down by chemical change.

3. A compound is two or more elements bonded together. It can only be broken down by chemical changes.

- ✓ Substances that form a compound gain new properties.
- ✓ The ratio of substances in a compound is constant (e.g. water has a fixed ratio 2:1 ratio of hydrogen to oxygen).

4. A physical change is one that results in the rearrangement of existing particles in a substance (ex: freezing, boiling). A chemical change results in the formation of different substances with different properties.

- ✓ Chemical and physical changes may be endothermic or exothermic.

5. The three phases of matter are solid, liquid and gas. Each has its own properties.

- ✓ Solids have a constant volume and shape. Particles are held in a rigid, crystalline structure.
- ✓ Liquids have a constant volume but a changing shape. Particles are mobile but still held together by strong attraction.
- ✓ Gases have no set volume or shape. They will completely fill any closed contained. Particles have largely broken free of the forces holding them together.

6. A heating curve (or cooling curve) traces the changes in temperature of a substance as it changes from solid to liquid to gas (or gas to liquid to solid).

- ✓ As temperature increases, kinetic energy increases and potential energy remains the same.
- ✓ When the substance undergoes a phase change, there is no change in temperature. The line "flattens" until the phase change is complete.
- ✓ When a phase change is occurring, the potential energy of the substance changes while kinetic energy remains the same.

7. Vapor Pressure is pressure created as particles evaporate

- ✓ Raising temperature increases vapor pressure.
- ✓ When vapor pressure equals air pressure, a substance will boil.
- ✓ Substances with stronger IMFs have low vapor pressures (and vice versa)
- ✓ Table H can be used to find vapor pressure and boiling point

8. The combined gas law states the relationship between pressure, temperature and volume in a sample of gas.

- ✓ Increasing pressure causes a decrease in volume (inverse relationship).
- ✓ Increasing temperature causes an increase in volume (direct relationship).
- ✓ Increasing temperature causes an increase in pressure.(direct relationship).
- ✓ The combined gas law can be found on – you guessed it – Table T!

9. The Kinetic Molecular Theory (KMT) says particles in an ideal gas:

- ✓ are in random motion.
- ✓ have no forces of attraction between them.
- ✓ have a negligible volume compared to the distances between them.
- ✓ have collisions that result in the transfer of energy from one particle to another, with no net loss of energy from the collision.

10. The KMT model is used to explain the behavior of gases.

- ✓ **A real gas is most like an ideal gas when it is at <u>high temperature and low pressure.</u>**
- ✓ The real gases that most resemble ideal gases are H_2 and He

11. Equal volumes of gases at the same temp and pressure have an equal number of particles.

Topic Seven: Solutions

1. A solution is a homogeneous mixture of a solute dissolved in a solvent.

- ✓ Solubility depends on temperature, pressure, and the nature of the solute and solvent.
- ✓ "Like dissolves like" – polar substances dissolve polar substances, and non-polar substances dissolve non-polar substances. Polar and non-polar do not mix.

2. Ionic substances dissolve in polar solvents. The positive ion is attracted to the negative end of the polar molecule, as the negative ion is attracted to its positive end.

3. Concentration of a solution can be expressed as molarity (M) or parts per million (ppm).

- ✓ Formulas for calculating concentration can be found in Table T.

4. Adding a solute to a solvent causes the boiling point of the solvent to increase and the melting point to decrease.

The more particles (greater concentration) of solute in a solution, the greater the effect on melting and boiling point.

5. A saturated solution exists in equilibrium – the rate of crystallization equals the rate of dissolving.

6. Table F can be used to find the solubility of a compound in water.

7. Table G shows the solubility of various substances in 100g of water at different temperatures.

- ✓ If the amount of dissolved solute is below the curve, it is unsaturated.
- ✓ If the amount of dissolved solute is exactly on the curve, it is saturated.
- ✓ If the amount of dissolved solute is above the curve, it is unsaturated.

Topic Eight: Acids, Bases and Salts

1. Behavior of many acids and bases can be explained by the Arrhenius theory. Arrhenius acids and bases are electrolytes.

- ✓ Common acids and bases are listed on Tables K and L
- ✓ Acid-Base indicators are listed on Table M

2. An electrolyte is a substance which, when dissolved in water, forms a solution capable of conducting electricity. |

3. Arrhenius acids yield H+(aq) ions as the only positive ion in solution.

- ✓ H^+(aq) ions may also be written as H_3O^+(aq) ions (hydronium ions).

4. Arrhenius bases yield OH^-(aq) ions as the only negative ion in solution.

- ✓ Organic compounds with OH^- are not bases.
- ✓ Ammonia (NH_3) is a base.

5. In neutralization reactions an Arrhenius acid and an Arrhenius base react to form salt and water.

- ✓ The <u>net ionic equation</u> for all neutralization reactions is the same: H^+(aq) + OH^- (aq) → H_2O (l)

6. Titration is a lab process in which a volume of a solution of known concentration is used to determine the concentration of another solution. Titration is a practical application of a neutralization reaction.

7. There are other acid-base theories besides the Arrhenius theory. The Bronsted-Lowry theory states an acid is a H^+ donor and a base a H^+ acceptor.

8. The acidity or alkalinity of a solution can be measured by pH.

- ✓ A low pH (0-6) indicates a higher concentration of H^+ ions than OH^- ions.
- ✓ A high pH (8-14) indicates a lower concentration of H^+ ions than OH^- ions.
- ✓ A neutral pH (7) indicates an equal concentration of H^+ ions than OH^- ions.
- ✓ Pure water has a neutral pH.

9. On the pH scale, each decrease of one pH unit represents a <u>tenfold</u> increase in H^+ ion concentration.

Topic Nine: Kinetics and Equilibrium

1. Collision theory states that a reaction is most likely to occur if reactant particles collide with the proper energy and orientation.

2. The rate of a chemical reaction depends on temperature, concentration, nature of the reactants, surface area and the presence of a catalyst.

3. Energy absorbed or released by a chemical reaction can be represented by a potential energy diagram.

4. The amount of energy released or absorbed during a chemical reaction is the heat of reaction (table I).

- ✓ Heat of reaction equals the PE of the products – PE of reactants.
- ✓ Positive heat of reaction implies an endothermic reaction.
- ✓ Negative heat of reaction implies an exothermic reaction.

5. A catalyst provides an alternative pathway for a chemical reaction. The catalyzed reaction requires a lower activation energy than the uncatalyzed reaction.

- ✓ Adding a catalyst increases the rate of the forward and reverse reactions equally, so there is no shift in equilibrium.

6. Entropy is a measure of the randomness or disorder in a system. A system with greater disorder has greater entropy.

7. Systems in nature tend to undergo changes towards lower energy and higher entropy.

8. Exothermic reactions that result in increased entropy are spontaneous.

10. At equilibrium the rate of the forward reaction equals the rate of the reverse reaction.

11. The measurable quantities of reactants and products remain constant at equilibrium.

12. LeChatelier's principle can be used to predict the effect of stress on a system in equilibrium.

- ✓ Stresses include a change in pressure, volume, concentration, and temperature.

Topic Ten: Oxidation-Reduction (Redox)

1. An oxidation-reduction (redox) reaction involves the transfer of electrons (e^-).

2. Reduction is the gain of electrons and decrease of oxidation number.
- ✓ A half reaction can be written to represent reduction.

3. Oxidation is the loss of electrons and increase of oxidation number.
- ✓ A half reaction can be written to represent oxidation.

4. In redox the number of electrons lost is equal to the number of electrons gained.

5. Oxidation numbers/states can be assigned to atoms and ions. Changes in oxidation numbers indicate that a redox reaction has occurred.
- ✓ Double replacement reactions are not redox reactions.
- ✓ A reaction in which an element is alone on one side of a reaction, and part of a compound on the other side is always a redox reaction.

6. An electrochemical cell can be either voltaic or electrolytic.

7. In an electrochemical cell oxidation occurs at the anode and reduction at the cathode.
- ✓ You can use Table J to determine the anode and the cathode. The anode is higher in a voltaic cell, and lower in an electrolytic cell.

8. A voltaic cell spontaneously converts chemical energy to electrical energy.

9. An electrolytic cell requires energy to produce a chemical change. This is called <u>electrolysis</u>.

Topic Eleven: Organic Chemistry

1. Organic compounds consist of carbon atoms which bond to each other in chains, rings and networks to form a variety of structures.

- ✓ **Names of organic compounds are based in part on the number of carbon atoms they contain. The root of the name is listed on table P.**

2. Organic compounds can be named with the IUPAC system.

3. Hydrocarbons are compounds that contain only carbon and hydrogen.

- ✓ Saturated hydrocarbons contain only single carbon-carbon bonds.
- ✓ Unsaturated hydrocarbons contain at least one multiple carbon-carbon bond (double or triple bond)
- ✓ Families of hydrocarbons are listed on Table Q.

4. Organic acids, alcohols, esters, aldehydes, ketones, ethers, halides, amines, amides, and amino acids are categories of organic molecules that differ in their structures (Table R).

5. Functional groups give organic molecules distinct physical and chemical properties (Table R).

6. Isomers of organic compounds have the same molecular formula but different structures and properties.

7. In a multiple covalent bond, more than one pair of electrons are shared between two atoms. Unsaturated organic compounds contain at least one double or triple bond.

8. Types of organic reactions include: addition, substitution, polymerization, esterification, fermentation, saponification, and combustion.

- ✓ Addition – An atom is added to an unsaturated molecule in place of a double or triple bond.
- ✓ Substitution – A new atom replaces another atom on an saturated molecule.
- ✓ Esterification – Organic acid + alcohol → ester
- ✓ Fermentation – Sugar decomposes forming alcohol and CO_2 (by yeast).
- ✓ Saponification – Makes soap from a fat and a base.
- ✓ Combustion – Hydrocarbon + O_2 → CO_2 + H_2O
- ✓ Polymerization – Smaller molecules (monomers) join to form a large polymer.

<u>Topic Twelve: Nuclear Chemistry</u>

1. The stability of an isotope depends on the ratio of protons to neutrons in the nucleus.

- ✓ Most nuclei are stable, but some are unstable. These nuclei will spontaneously decay, emitting radiation.
- ✓ Stable isotopes of small atoms have a 1:1 ratio of protons and neutrons. Most radioactive isotopes have twice as many neutrons as protons.
- ✓ All elements with an atomic number higher than 83 are radioactive.

2. Each isotope has a specific mode and rate of decay. (Table N)

- ✓ The rate of decay is called <u>half-life</u>.
- ✓ Half-life is a <u>constant</u> that can <u>never</u> be changed.
- ✓ Half-life is the measure of the time it takes exactly <u>one half</u> of an amount of isotope to decay.
- ✓ The amount of substance will never decay to zero.

3. A change in the nucleus of an atom changes it to a new type of atom (i.e. a new element). This is called <u>transmutation</u>.

- ✓ Transmutation can occur naturally or artificially.
- ✓ Artificial transmutation requires the bombardment of a nucleus by high energy particles.

4. Spontaneous decay can involve the release of different particles from the nucleus.

- ✓ The types of particles, as well as their masses and charges, can be found on <u>Table O</u>.
- ✓ <u>Decay Mode</u>: How a nucleus decays, determined by the particle emitted. Can be found on <u>Table N.</u>

5. Nuclear reactions include natural and artificial decay, nuclear fission and nuclear fusion.

- ✓ Nuclear fission occurs when the nucleus of an atom is split. This can be caused artificially by "shooting" the nucleus with a neutron.
- ✓ Nuclear fusion combines two light nuclei to form heavier nuclei. Nuclear fusion is the process that powers the sun.
- ✓ Nuclear fusion requires very high temperatures and is not yet ready for practical energy production. The main advantage it offers is that the products are not radioactive wastes (as with fission).

6. Nuclear reactions can be represented by equations that include symbols which represent atomic nuclei (with mass number and atomic number), subatomic particles (with mass and charge) and emitted particles.

7. Energy from nuclear reactions comes from the very small fraction of mass that is lost – the reaction converts matter into energy.

- ✓ Einstein's $E=mc^2$ describes the relationship between energy and matter.

8. The energy released from nuclear reactions is much greater than that released from chemical reactions.

9. The risks associated with using radioactive isotopes include biological exposure (which may cause radiation poisoning and cancer), long-term storage and disposal, and nuclear accidents.

10. Radioactive isotopes may be used in medicine (tracing chemical and biological processes), radioactive dating, industrial measurement, nuclear power, and detection and treatment of disease.

- ✓ Carbon 14 is used for radioactive dating of organic material.
- ✓ Uranium-238 is used for radioactive dating of rocks and the Earth.
- ✓ Iodine-131 is used to treat thyroid disorders.
- ✓ Cobalt-60 is used in radiation therapy to treat cancer.

Topic Thirteen: Lab skills

1. Any standard chemistry lab procedure is fair game for the Regents. Specific skills that may be tested include:

- ✓ Using the scientific method for a controlled experiment.
- ✓ Construct a graph.
- ✓ Use proper units of measurement.
- ✓ Making accurate and precise measurements.
- ✓ Calculate % error (Table T).
- ✓ Use rules for significant figures.
    - Significant figures are used when doing math with measurements and reflect the precision of your measurements.
    - Significant figures include all digits that you are certain of, plus one that is an estimate.
    - All non-zero digits are significant.
    - All leading zeros (zeros before the 1^{st} non-zero digit) are not significant, whether or not there is a decimal.
    - All zeros between non-zero numbers are significant.
    - Trailing zeros (zeros after the last non-zero number) are significant *if and only* if there is a decimal point.
    - Constants (such as the freezing point of water at STP) and certain numbers (the number of people in the room) are considered to have an infinite number of sig figs.
    - For multiplication and division, the answer should have the same number of digits as the measurement with the fewest sig figs.
    - For addition and subtraction, the answer must have the same number of *decimal places* as the measurement with the fewest number of decimal places. The total number of digits does not matter.
    - When problems have a mix of multiplication/division and addition/subtraction, use the multiplication/division rules.
- ✓ Identification and use of lab equipment.
- ✓ Lab safety.

Regents Examination in Physical Setting/Chemistry – June 2017

Chart for Converting Total Test Raw Scores to Final Examination Scores (Scale Scores)

Raw Score	Scale Score
85	**100**
84	**98**
83	**97**
82	**95**
81	**94**
80	**92**
79	**91**
78	**90**
77	**88**
76	**87**
75	**86**
74	**85**
73	**84**
72	**83**
71	**82**
70	**81**
69	**80**
68	**79**
67	**78**
66	**77**
65	**76**
64	**75**

Raw Score	Scale Score
63	**74**
62	**74**
61	**73**
60	**72**
59	**71**
58	**71**
57	**70**
56	**69**
55	**68**
54	**68**
53	**67**
52	**66**
51	**66**
50	**65**
49	**64**
48	**64**
47	**63**
46	**62**
45	**62**
44	**61**
43	**60**
42	**60**

Raw Score	Scale Score
41	**59**
40	**58**
39	**57**
38	**57**
37	**56**
36	**55**
35	**54**
34	**53**
33	**53**
32	**52**
31	**51**
30	**50**
29	**49**
28	**48**
27	**47**
26	**46**
25	**45**
24	**44**
23	**43**
22	**42**
21	**40**
20	**39**

Raw Score	Scale Score
19	**38**
18	**36**
17	**35**
16	**33**
15	**32**
14	**30**
13	**29**
12	**27**
11	**25**
10	**23**
9	**21**
8	**19**
7	**17**
6	**15**
5	**13**
4	**10**
3	**8**
2	**5**
1	**3**
0	**0**