Name: ________________________________

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

The Test

The Chemisty Regents Exam is broken down into three sections:

Part A: 35 mulitple choice questions from all units covered over the course of the school year.

Part B: Approximately 25 questions, with a mix of short answer and multiple choice. Questions focus on the Reference Tables, graphing, and laboratory experiments.

Part C: Approximately 15 short answer questions, most broken down into smaller parts. This is often an eclectic, unpredictable mix of questions from various units, and may demand students write short paragraphs, use equations and reference tables, or draw graphs and diagrams in order to correctly answer the questions.

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

Exam Date:________________________________

There are 12 specific topics covered on the test. In addition to these you will be required to demonstrate math and graphing skills. The 12 topics covered are:

The Atom	Moles and Stoichiometry
Nuclear Chemistry	Solutions
Bonding	Kinetics and Equilibrium
Matter	Acids, Bases and Salts
Energy	Oxidation-Reduction (Redox)
The Periodic Table	Organic Chemistry

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 have to study, ask questions, analyze problems and come to review sessions to be thoroughly prepared.

Topic One: The Atom

1. The modern model of the atom has evolved over a long period of time through the work of many scientists.

- ✓ *Dalton's Model:*
  - Elements are made of atoms
  - Atoms of an element are the same.
  - Compounds are formed from combinations of atoms.
- ✓ *Rutherford Experiment*
  - Bombarded gold foil with alpha particles. Showed atoms were mostly empty space with small, dense positively charged nucleus.
- ✓ *Bohr Model*
  - Small, dense, positively charged nucleus surrounded by electrons in circular orbits.
- ✓ *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. Protons have a positive charge, neutrons no charge, and electrons a negative charge.

5. The number of protons in an atom equals the number of electrons.

- ✓ The positive charges of the protons are cancelled by the negative charges of the electrons, so overall an *atom* has a neutral charge.

6. 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 a Carbon atom.
- ✓ The <u>atomic mass</u> of an atom is equal to the total number of protons and neutrons.

7. Each electron in an atom has its own distinct amount of energy.

- ✓ When all electrons are at their lowest possible energy, it is called the "<u>ground state</u>."
- ✓ Electrons fill in energy levels and orbitals starting with the one that requires the least energy (1s) and progressively move to those levels and orbitals that require increasing amounts of energy.

8. When the electron gains a specific amount of energy, it moves to a higher orbital and is in the "<u>excited state</u>".

9. When an electron returns from a higher energy state to a lower energy state, it emits a specific amount of energy usually in the form of <u>light</u>. This can be used to identify an element (bright line spectrum).

- ✓ The instrument used to see the bright line spectrum is called a spectroscope.

10. The outermost electrons are called <u>valence electrons</u>. These affect the chemical properties of the element.

- ✓ Atoms with a filled valence level are stable.
- ✓ 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>orbital notation</u> or <u>Lewis structures</u> to show the configuration of the valence electrons.

11. 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.

12. <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).

13. The average atomic mass of an element is the weighted average of its naturally occurring isotopes.

<u>Topic Two: 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 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. (see 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>.

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 use. 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.

Topic Three: Bonding

1. Chemical compounds are formed when atoms are bonded together.

- ✓ Breaking a chemical bond is an endothermic process.
- ✓ Forming a chemical bond is an exothermic process.
- ✓ Compounds have less potential energy than the individual atoms they are formed from.

2. Two major categories of compounds are ionic and molecular (covalent) compounds.

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

- ✓ Ionic substances have high melting and boiling points, form crystals, dissolve in water (dissociation), 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.

4. Chemical bonds are formed when valence electrons are:

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

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

6. Polarity of a molecule can be determined by its shape and the distribution of the charge.

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

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

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

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

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

10. Electron-dot diagrams (Lewis structures) represent the valence electron arrangement 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.

11. <u>Electronegativity</u> indicates how strongly an atom of an element attracts electrons in a chemical bond. These values are based on an arbitrary scale.

12. The electronegativity difference between two bonded atoms can determine the type of bond and its polarity.
0.0 - 0.4 = non-polar covalent
0.4-1.7 = polar covalent
1.7+ = ionic

13. Bonding guidelines:
- ✓ 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.

14. 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.

15. 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.

<u>Topic Four: Matter, Phases and Gas Laws</u>

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.

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.
- ✓ Gasses 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).

- ✓ 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.
- ✓ As temperature increases, kinetic energy increases.

7. <u>Heat of fusion</u> (H_f)is the energy needed to convert one gram of a substance from solid to liquid.

8. <u>Heat of vaporization</u> (H_v) is the energy needed to convert one gram of a substance from liquid to gas.

9. <u>Specific heat</u> (C) is the energy required to raise one gram of a substance 1 degree (Celcius or Kelvin).

- ✓ The specific heat of liquid water is 1 cal/g*J or 4.2 J/g*K.

10. 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).

11. An ideal gas model is used to explain the behavior of gasses. A real gas is most like an ideal gas when it is at <u>low temperature and high pressure.</u>

12. The Kinetic Molecular Theory (KMT) for an ideal gas states that all gas particles:

- ✓ 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, but there is no net loss of energy from the collision.

13. Equal volumes of gasses at the same temp and pressure have an equal number of particles.

<u>Topic Five: Energy</u>

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

- ✓ Stored energy is referred to as <u>potential energy</u>.
- ✓ Energy of motion is <u>kinetic energy.</u>

2. The Law of Conservation of Energy states that energy can not be lost or destroyed, only changed from one form to another.

3. Heat is a transfer of energy (often but not always thermal energy) from a body of higher temperature to a body of lower temperature.

4. Temperature is a <u>measure</u> of the average kinetic energy of the particles in a sample. <u>Temperature is NOT a form of energy</u> and should not be confused with heat.

5. The concepts of kinetic and potential energy can be used to explain physical processes such as fusion (melting), solidification (freezing), vaporization (boiling, evaporation), condensation, sublimation, and deposition.

6. Processes that are <u>exothermic</u> give off heat energy. This typically causes the surrounding environment to become warmer.

7. Processes that are <u>endothermic</u> absorb energy. This typically causes the surrounding environment to become colder.

<u>Topic Six: 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 atomic mass is the sum of protons and neutrons in the nucleus.
- ✓ The mass number given on the periodic table is a weighted average of the different isotopes of that element.
- ✓ 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.
- ✓ Examples of isotopic notation are: ${}^{14}_{6}C$, ${}^{14}C$, carbon-14, C-14

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

7. Elements may be differentiated by their physical properties.
- ✓ Ex: Density, conductivity, malleability, hardness, ductility, solubility

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 gasses*. 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.

14. Some elements may exist in two or more forms in the same phase. These forms differ in their molecular or crystal structure, hence their different properties.

- ✓ Ex: Carbon exists as both graphite and diamond (a network solid).

<u>Topic Seven: Moles 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 show a conservation of mass, energy and charge.

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.

<u>Topic Eight: Solutions</u>

1. A solution is a homogeneous mixture of a <u>solute</u> dissolved in a <u>solvent.</u>

- ✓ 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 <u>molarity</u> (M), percent by volume, percent by mass, or parts per million (ppm).

4. Adding a solute to a solvent causes the boiling point of the solvent to <u>increase</u> and the freezing point to <u>decrease</u>.

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

<u>Topic Nine: Kinetics and Equilibrium</u>

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 several factors: temperature, concentration, nature of the reactants, surface area and the presence of a catalyst.

3. Some chemical and physical changes can reach equilibrium.

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

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

6. 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.

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

8. The amount of energy released or absorbed during a chemical reaction is the heat of reaction.

- ✓ 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.

9. 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.

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

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

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

<u>Topic Ten: Acids, Bases and Salts</u>

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

2. An electrolyte is a substance which, when dissolved in water, forms a solution capable of conducting electricity. The ability to conduct electricity depends on the concentration of ions.

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. One states that an acid is an H^+ donor and a base an H^+ acceptor.

8. The acidity or alkalinity of a solution can be measured by pH.
- ✓ A low pH indicates a higher concentration of H^+ ions than OH^- ions.
- ✓ A high pH 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.

<u>Topic Eleven: Oxidation-Reduction (Redox)</u>

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.

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>.

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

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

- ✓

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).

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

5. Functional groups give organic molecules distinct physical and chemical properties.

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.

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.
- ✓ Use rules for significant figures.
- ✓ Identification and use of lab equipment.
- ✓ Lab safety.