Chemistry Regents Review Survival Guide

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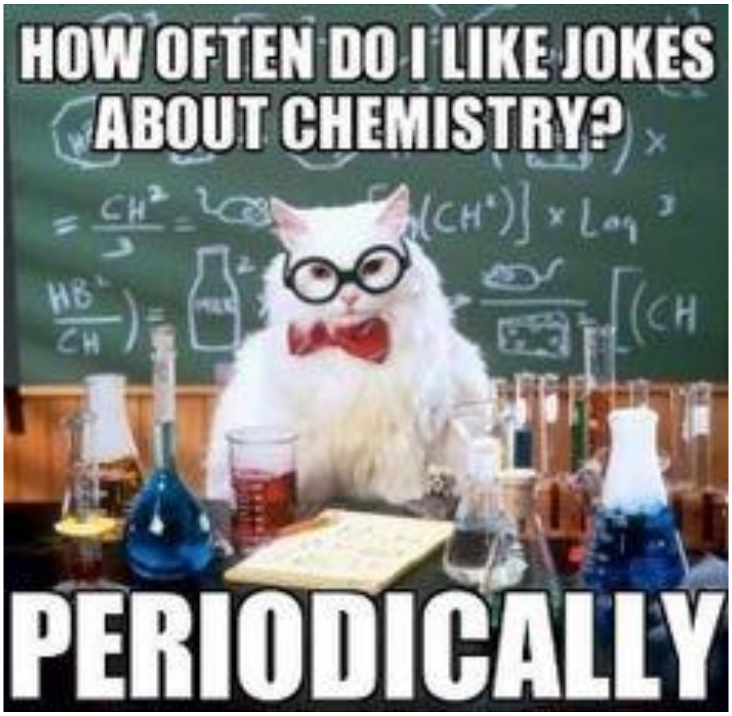

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Unit 1 – Atomic Concept

The Nature of the Atom, Subatomic Particles, Atomic Structure, Energy Levels, Valance Electrons

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 atomic mass 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 "ground state."
- ✓ 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 "excited state".

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 light. 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 valence electrons. 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 orbital notation or Lewis structures 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. Isotopes 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.

Unit 2 – Periodic Table

Development of the Periodic table, Properties of Elements, Chemistry of a Group, Chemistry of a Period, Naming Elements

The Periodic Table

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).

Unit 3 – Moles / Stoichiometry

Formula Writing, Naming & Writing Chemical Compound Formulas, Chemical Equations, Mole Interpretation, Stoichiometry

Moles and Stoichiometry

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 mole ratios 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.

Unit 4 – Chemical Bonding

The Nature of Chemical Bonding, Directional Nature of Covalent Bonds, Intermolecular Forces

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.
- ✓ Non-polar molecules are symmetrical and/or have no polar bonds.
- ✓ REMEMBER! SNAP → Symmetrical Non-Polar, Asymmetrical Polar

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. Electronegativity 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. (BEND → Bond Electronegativity Difference)

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.

- ✓ Hydrogen bonds 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.

Unit 5 – Physical Behavior of Matter

Phases of Matter, Phase Changes, Substances, Mixtures, Solutions, Effect of Solute on Solution, Energy, Kinetics

Matter, Phases and Gas Laws

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

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

9. Specific heat (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 gases. A real gas is most like an ideal gas when it is at HIGH TEMPERATURE AND LOW PRESSURE (think, your ideal vacation)

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.

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

✓ Stored energy is referred to as potential energy.
✓ Energy of motion is kinetic energy.

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

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

17. Temperature is a measure of the average kinetic energy of the particles in a sample. Temperature is NOT a form of energy and should not be confused with heat.

18. 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.

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

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

21. 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.

22. 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.

23. Concentration of a solution can be expressed as molarity (M), percent by volume, percent by mass, or parts per million (ppm).

24. Adding a solute to a solvent causes the boiling point of the solvent to increase and the freezing point to decrease.

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

Unit 6 – Kinetics / Equilibrium

Kinetics, Equilibrium, Spontaneous Reactions

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 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.

Unit 7 – Organic Chemistry

Characteristics of Organic Compounds, Bonding, Homologous Series of Hydrocarbons, Organic Reactions

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.

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.

Unit 8 – Oxidation/Reduction (REDOX)

Reduction, Oxidation, Redox Reactions, Corrosion, Half-Reactions, Electrochemical (Voltaic) Cell, Electrolytic Cell, Electroplating, Reduction of Metals

Oxidation-Reduction (Redox)

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

- ✓ **REMEMBER: LEO the GER** (lose electrons = oxidation & gain electrons = reduction)

2. Reduction is the gain of electrons (GER) and decrease of oxidation number.

- ✓ A half reaction can be written to represent reduction.

3. Oxidation is the loss of electrons (LEO) 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 (REDCAT → Reduction at cathode ANOX → Oxidation at anode)

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 electrolysis.

- ✓ **Usually a battery is present in order to provide the electricity to start the reaction**

Unit 9 – Acids, Bases, and Salts

Acids & Bases, Acid-Base Reactions, Salts, Normality

Acids, Bases and Salts

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 net ionic equation for all neutralization reactions is the same: $H^+(aq) + OH^-(aq) \rightarrow H_2O\ (l)$

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 tenfold increase in H^+ ion concentration.

Unit 10 – Nuclear Chemistry

Half-Life, Natural Radioactivity, Artificial Radioactivity, Nuclear Energy

Nuclear Chemistry

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 half life.
- ✓ Half-life is a constant that can never be changed.
- ✓ Half life is the measure of the time it takes exactly one half 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 transmutation.

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 Table O.

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.

Lab Skills for the Regents Exam

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.

Last Minute Things to Study

Based on Mistakes Seen

Topic	Point to Remember
Heating Curve	**PE inc. during phase change & KE remains the same (temp stays same); on the L only and G only sections, the KE inc and PE remains the same**
Nuclear Charge	**The charge of the nucleus is equal to the # of protons**
Lewis Dot	**Shows Valence e- only for elements; remember metal ions have no dots- just put the charge; nonmetal ions have gained e- based on charge- use brackets and put the charge outside.**
Lewis dot for compounds	**Ionic compounds 1.7 EN diff – use brackets with charges and put enough e- to show the nonmetal charge gained e-; remember to write the correct formula (ie $AlCl_3$ would have one Al^{+3} and** **3 of the $[Cl]^{-1}$ with the 8 dots around the Cl.** **If Covalent- 1.6 or lower EN – no brackets because e- are shared** **Remember nonpolar covalent BrINCLHOF (double bond on O_2 and triple bond on N_2**
Like dissolves like	**Water is polar and will dissolve other polar molecules** **Nonpolar hydrocarbons will dissolve other nonpolar molecules** **Water and Oil don't mix(don't dissolve)because one is polar and the other is nonpolar!**
Addition reactions aka halogenation (chlorination, bromination, etc)	**Organic reaction between halogen and an unsaturated alkene or alkyne where the double or triple bond gets reduced to a single or double bond and the 2 halogen atoms go where the multiple bond site was(one on each C of the multiple bond)**
Substitution reaction	**Organic reaction between halogen and an saturated alkane where the 1 halogen atom goes on 1 carbon and that H that you take off goes with the left over halogen to make 2 products**
Weighted Ave Mass of combined isotopes	**Multiply the % by the mass of that isotope (move decimal of % value back 2 places) and do that for as many isotopes as you have, then add up all those parts. Think about it. If you have 4 isotopes and one of them is 90% with a mass of 40, then your answer should be close to 40; Do NOT add up and divide by the number of isotopes!!!!!!!!!!**

Color change	**If they ask for an indicator color change then you need to include the starting color and the final color!**
Intermolecular forces	**Weak ones will be seen in substances with low Boiling pts.(low vapor pressure)** **Strong ones – higher BP- like WATER- has Hydrogen bonds as intermolecular forces that attract the molecule polar (-) end to the other molecule nearby (+) end.**
Intermolecular forces	**Van der waal intermolecular forces account for the change you see in Group 17 (gas, gas, liquid, solid) ; get stronger going down group b/c molecular mass increases (so BP is higher at bottom) same is true of noble gas group**
Titration problems- do not just use the Molarity formula	**Read the whole paragraph and label the info. Ma Va Mb Vb** **Remember if they give you initial and final level readings, you need to subtract to find the actual Va or Vb.** **- If a diprotic acid is used then the actual Ma is 2x whatever they say the molarity is. (same idea for a base like $Mg(OH)_2$** **Plug your answer back in to check.**
Formula Mass/mole problems	**Create the chart where you list the Element/# of atoms/ Atomic Mass/ Multiplied amount and then add up all the parts.** **Write out the formula for finding moles and check your answer!**
Moles	**22.4 L of gas = 1 mole; 6.02 x 1023 molecules = 1 mole; formula mass = 1 mole**
Mole equation problems	**Set up the ratio using the coefficients of the balanced equation as the bottom # in the ratio set up. The top set has the unknown and the other given info**
Le Chatelier principle	**You need to say, "Because ________, the reaction will shift to the _______ and thus ________ will __________, in order to reestablish equilibrium."**
Kinetic Molecular Theory	**Use phrases like: "At higher temps, the molecules have greater KE and will move faster and collide more often and the reaction rate will go up." Or "Higher collision rate occurs at higher concentration because there are more molecules present to collide with each other at a higher frequency."**
Nuclear equations	**Use table O carefully- there are Beta positive and Beta negative particles.**

½ life	**Make the chart, start time at zero and multiply ½ life time amt by 1, 2, 3, 4, as you complete the time side; the mass side gets decreased by ½ each time.**
Nuclear reactions	**A half life problem is a type of *natural radioactive decay or a natural transmutation.*** ***Fission* is the splitting of a uranium nucleus by hitting it with a neutron and a small bit of mass of the reactants is converted into a large amount of energy ($E=mc^2$**) exothermic reaction ***Fusion*** is the joining of 2 light H nuclei that have to be traveling really fast to smash those similarly charged nuclei together- so it has to be really hot! – problem on earth- happens in sun- really exothermic- same idea of mass converts to energy.
Conservation of mass (Matter)	**A balanced equation has the same # of each type of element (atom) on each side of the arrow.** ***There are not the same # of molecules on each side!***
REDOX	**Remember the LEO says GER to help write the correct half reactions.** **e- are on the right side of arrow if its oxidation and on the left side if its reduction. Charges go up for OX and go down for Red.** **Put the charge of 0 above the element written by itself.** **Table J- higher metal is anode- more reactive metal gets oxidized** **Salt bridge is for the flow of IONS – maintain neutral solutions.** **E- flow from anode to cathode** **Voltaic cells or electrochemical cells are spontaneous and convert chemical energy to electrical energy** **Electrolytic cells use electrical energy (battery or voltage supply) to force(drive) a chemical reaction and are not spontaneous- ex:** **Electroplating a spoon with silver ions- the spoon is attached to the negative terminal of the electric supply and will attract the Ag +1 ions to it and reduction will occur there at the spoon.**
Kelvin temp	**Look up and write out the equation and use your calculator and then check your answer!!!!!**
Gas Laws	**See above! Convert temp to Kelvin too. Read the paragraph info and label P_1 ,V_1 ,T_1 as you go.**

Graphs	Make sure the axes are both scaled uniformly(count by some even number (or equal interval) for both axes. Do not extend line past data unless told to.
Wave mechanical model	This is the same as the electron cloud model and has the electrons found in (sub)orbital areas- regions of space based on probability.
conductivity	Metals conduct due to sea of mobile valence e-; loosely held e- Solutions that are electrolytes include acids, bases and salt solutions and conduct due to the free mobile ions in solution.
Saturated solution	Is at equilibrium (rate of dissolving = rate of crystallization)
Exothermic/endothermic Use Table I	EXo – get hot- combustion - negative Delta H Endo- get cold- + delta H remember the NaCl in ice H20 got really cold- wonderful ice cream
Forming Ions	Metals lose e- and the radius gets smaller Nonmetals gain e- and radius gets a bit bigger
Allotropes	O2 and O3 are different forms of same element and have different physical and chemical properties; same for graphite and diamonds- both made of carbon
Organic reactions	Use your ref table to identify the product by diagram Alcohol- fermentation (sugar and enzymes- also makes CO2 Ester- esterification (alcohol and organic acid) Saponification= lye (strong base (NaOH) and a fatty acid (lard, oil,etc) – makes soap Polymerization- monomer units make up the polymer--- Ex: addition polymerization (add Cn H2n – units together) or condensation polymerization- joining molecules that lose a water in the joining process
Water	Boils at 100 C and freezes at 0 C- has polar covalent bonds and is a bent polar molecule- has Hydrogen bonds that make it have a High BP.
Stock system of naming	Use roman numeral to tell the charge of the metal or positive nonmetal atom in the formula

PPM	Don’t forget to write out the 1000000 in the set up. The denominator must include the water and the solute. If the wording says a 250. g sample of the water is taken and analyzed, that implies that the solute is already in the sample of water. If the wording says that .003 g of $CuSO_4$ is found in 155 ml of water, then you need to add the 155 and the .003 for the denominator.
Sig fig	When you add up or subtract numbers, go by the # of decimal points (least) when you multiply or divide, count the total # of sig fig in the original measurements and then round off to the least # of sig figs. 1000 has 1 sig fig 1000. has 4 .0003 has 1sig fig .000300 has 3 0.90 has 2 .9 has 1 40.00 has 4 40. has 2 40 has 1
General Test taking strategies	Relax- take your time- read the question several times, label info as you read-try to use reference table as much as possible- look up the names/symbols on Table S just to make sure you don’t mess up S for Sodium instead of sulfur or P for Potassium/Phosphorus or put a Fl for Fluorine or a C for chlorine- or B for Bromine (ALL of those mistakes have been made)- use scrap areas to jot down things you remember- remember units- write out the formula for problem solving- put your answer back in and check the work. Triple check your work to get the better grade. DO NOT RUSH just to go home. Remember, the grading on this particular regents is the harshest one of all exams. They do not cut any breaks for the upper end grades. You have to work really hard to avoid the traps set up. Be on the lookout and re-read everything!!!!!