Arlington High School
Chemistry Syllabus
September 2002

Introduction:

This syllabus represents the content that will be covered in all Regents Chemistry classes at Arlington High School. It is carefully referenced to the New York State Core Curriculum for the Physical Setting/Chemistry. The Core Curriculum represents the minimum that any Regents Chemistry course should cover and defines the material that can be tested on the NYS Regents Examination.

Schools are expected to supplement the NYS Core Curriculum in their course syllabi. This syllabus includes supplementary material which Arlington teachers view as essential to a complete course in Chemistry. Individual teachers are encouraged to further supplement in their own classes.

Laboratory:

Laboratory work is an integral part of Regents Chemistry and is required before a student may take the Regents Exam. Prior to taking any Regents science exam, New York State requires that each student complete and write satisfactory lab reports for a minimum of twelve hundred minutes of laboratory work. The school is required to keep these reports on file. Any student who fails to meet the laboratory requirement will automatically fail the course because of being barred from the final exam. It is our practice to offer more than the minimum number of labs so that any conscientious student will have no difficulty in meeting the lab requirement. When a student is legally absent from a class lab, the teacher will provide an opportunity for the youngster to make up the lab in a reasonable period of time.

Identifying Core Curriculum material:

Core curriculum material is identified by references to the Physical Setting/Chemistry Core curriculum (e.g., **(3.1kk) "(3.1kk)"** identifies that this concept is listed in the Physical Setting/Chemistry Core, Standard 4, Performance indicator 3.1kk). Process skills, found on pages 12 – 15 in the core curriculum guide, are indicated by Roman numerals (3.1xxiii). **CRT** refers to the fact that the material will be found in the Chemistry Reference Tables and is testable even if it does not appear as a Performance Indicator.

When a topic or vocabulary term which is not in the Chemistry Core is used in the syllabus, it is placed in parentheses. This indicates that it is an important topic or term which should be taught, but can not be tested on the Regents Exam.

Arlington High School
Chemistry Syllabus 2002-2003

I. Introductory Material
  A. Introduction to the lab
    1. Safety (s2.4)
    2. Review of scientific method, controlled experiments, and graphing (m1.1, m2.1, s1.1, s1.2, s1.3, s2.1, s2.2, s2.3, s2.4, s3.1, s3.2, s3.4, s3.5)
  B. Chemical Measurement
    1. Use of SI units – knowledge of SI units, derivation of SI units. (m1.1, **CRT**)
    2. Determination of the amount of certainty in a measured quantity (m1.1, s3.2, s3.3, **CRT**)
      a.) Accuracy and precision
      b.) Reporting measurements to the correct number of significant digits
      c.) Calculating % error
  C. Problem Solving
    1. Use a suitable method for problem solving (m1.1, m3.1)
    2. Dimensional analysis (m1.1)
    3. Scientific notation
  D. Matter **(3.1q)**
    1. Phases **(3.1kk, 3.1xxii)**
      a.) Phase changes **(4.2ii, CRT)**
    2. Substances **(3.1r, 3.1xxxvi, CRT)**
      a.) Elements **(3.1u)**
      b.) Compounds **(3.1cc, 3.1dd)**
    3. Physical changes vs. Chemical changes **(3.2a)**
    4. Mixtures **(3.1s, 3.1t)**
      a.) Homogeneous **(3.1s, 3.1oo)**
      b.) Heterogeneous **(3.1s, 3.1t)**
      c.) Separating mixtures **(3.1nn, 3.1xxiv)**
    5. Solutions **(3.1oo, CRT)**
      a.) Factors affecting solubility **(3.1oo)**
      b.) Interpret and construct solubility curves **(3.1xxv)**
      c.) Saturated, supersaturated and unsaturated solutions **(3.1xxviii)**
      d.) "Like dissolves like" concept **(3.1xxvi)**
      e.) Concentration (molarity,% by mass, % by volume, p.p.m.) **(3.1pp,3.1xxx, 3.1xxix)**
      f.) Boiling point elevation, freezing point depression **(3.1qq)**

E. Energy
1. Forms of energy **(4.1a)**
2. Transfer of heat energy **(4.2a)**
3. Contrast Heat energy and temperature **(4.2i)**
4. Heating and cooling curves **(4.2c, 4.2iii)**
    a.) Kinetic energy **(4.2c)**
    b.) Potential energy **(4.2c)**
    c.) Heat of vaporization - boiling, evaporation, condensation **(4.2iii, CRT)**
    d.) Heat of fusion - melting, freezing, solidification **(4.2iii, CRT)**
    e.) Sublimation, deposition **(4.2c)**
5. Temperature **(4.2b, 4.2i)**
    a.) Phase changes **(4.2ii)**
    b.) Thermometry **(3.4iii)**
6. Measurement of heat in joules **(4.2iv)**
7. Exothermic and endothermic reactions **(4.1b)**
8. Potential energy diagrams **(4.1i, 4.1ii))**

F. Gases
1. Real vs. Ideal gases **(3.4a)**
2. Kinetic molecular theory (KMT) **(3.4b, 3.4c)**
3. Collision theory **(3.4d)**
4. Combined gas law **(3.4ii)**
5. Gas laws and KMT **(3.4i)**
6. Equal volumes contain equal numbers of particles **(3.4e)**

II. Structure of the atom
A. Atomic Structure
1. Historical perspective of the model of the atom **(3.1a)**
    a.) In 1803, John Dalton proposed the atomic theory which stated that all matter is made of atoms, atoms of the same type of element have the same chemical properties, compounds are formed by two or more different types of atoms, and that a chemical reaction involves either, joining, separating, or rearranging atoms.
    b.) In 1910, Ernest Rutherford passed alpha particles through gold foil which showed that atoms are mostly empty space, and that the positive charge within an atom is centrally located.
    c.) In 1913, Neils Bohr used classic physics to develop a model of an atom where electrons orbit a central nucleus in concentric circles.
    d.) The electron cloud model shows the probability of an electron to be in a specific location in an atom.
2. Current atomic model **(CRT)**
    a.) Atoms are made of protons, neutrons and electrons. **(3.1b, 3.1c, 3.1d, 3.1e, 3.1f)**
        i.) protons have a positive charge, are located in the nucleus, and have a mass of 1.007 amu.

ii.) neutrons have no charge, are also located in the nucleus, and have a mass of 1.009 amu.

iii.) electrons have a negative charge, are located outside of the nucleus, and have a mass of 5.486×10^{-4} amu or 1/1837 amu.

b.) Atomic number is the number of protons in an atom. Mass number is the number of protons plus the number of neutrons. **(3.1g)**

c.) Isoptopes are atoms of the same element having different numbers of neutrons, same number of protons. **(3.1m, 3.1g)**

d.) Average atomic mass is the weighted average mass of all naturally occurring isotopes of an element. **(3.1n)**

3. Electron Configuration **(CRT)**

a.) Electrons fill energy levels in order of lowest energy to highest energy.

b.) The first energy level holds two electrons, the second holds eight, the third holds eighteen, and the fourth holds thirty-two.

c.) No ground state atom has more than thirty-two electrons in any one energy level.

4. The wave mechanical model of the atom is used to show orbitals defined as regions of probable electron location. **(3.1h)**

a.) (Sublevels and orbital diagrams may be included.)

5. Each electron has a distinct amount of energy **(3.1i)**

a.) Electrons can gain energy and jump to an excited state. **(3.1j)**

i.) absorption spectra

b.) When electrons drop back down to the ground state they emit a specific amount of energy, which can be used to identify the atom. **(3.1k)**

i.) emission spectra

6. The valence electrons affect the chemical properties of an atom. **(3.1l)**

III. Physical Behavior of Matter

A. Matter can be either a pure substance or a mixture of substances. **(3.1q)**

1. There are three phases of matter, solid, liquid, and gas, and each phase has its own unique properties. **(3.1kk, 3.1xxii)**

a.) A phase change happens when a substance changes from one phase to another. **(4.2ii, CRT)**

2. Substances are specific types of matter. **(3.1r, 3.1xxxvi, CRT)**

a.) Elements are substances made of only one type of atom. **(3.1u)**

b.) Compounds are substances made of two or more types of elements. **(3.1cc)**

i.) Compounds are represented by chemical formulas. **(3.1cc)**

ii.) Compounds are differentiated by physical and chemical properties. **(3.1dd)**

3. A physical change is a rearrangement of particles, but a chemical change is one that results in new types of particles. **(3.2a)**
4. A mixture is made up of two or more different substances, in various proportions, which retain their original properties. **(3.1s, 3.1t)**
    a.) A mixture is homogeneous if the mixture is the same throughout. A solution is an example of a homogeneous mixture. **(3.1s, 3.1oo)**
    b.) A mixture is heterogeneous if the mixture varies throughout. **(3.1s, 3.1t)**
    c.) A mixture, depending on its density, particle size, molecular polarity, boiling point, freezing point and solubility, can be separated by filtration, distillation, and/or chromatography. **(3.1nn, 3.1xxiv)**
5. A solution is a uniform mixture of a solute and a solvent. **(3.1oo, CRT)**
    a.) Temperature, pressure, and the chemical nature of the solute and solvent all affect solubility **(3.1oo)**
    b.) A solubility curve is a graph of the relationship between solubility and temperature. **(3.1xxv)**
    c.) Solubility curves can be used to distinguish between saturated, supersaturated and unsaturated solutions **(3.1xxviii)**
    d.) The "like dissolves like" concept can be used to predict real life situations. **(3.1xxvi)**
    e.) Solutions can be diluted to create varying concentrations, measured in molarity, percent by mass, percent by volume, parts per million. **(3.1pp, 3.1xxx, 3.1xxix)**
    f.) A solution's boiling point will increase, and its freezing point will decrease with the addition of a solute. **(3.1qq)**

B. Energy
1. Energy can exist in many forms, including chemical, electrical, electromagnetic, thermal, mechanical, and nuclear. **(4.1a)**
2. Energy in the form of heat is transferred from an object of higher temperature to an object of lower temperature. **(4.2a)**
3. Heat is the energy associated with the random motion of atoms. **(4.2i)**
4. Heating and cooling curves **(4.2c, 4.2iii)**
    a.) Kinetic energy is the energy of motion **(4.2c)**
    b.) Potential energy is stored energy. **(4.2c)**
    c.) Heat of vaporization is the energy needed to convert a liquid to a vapor: boiling, evaporation, condensation. **(4.2iii, CRT)**
    d.) Heat of fusion is the energy needed to convert a solid to a liquid: melting, freezing, solidification. **(4.2iii, CRT)**
    e.) Sublimation is the conversion of a solid directly into a gas, while deposition is the conversion of a gas directly into a solid. **(4.2c)**
5. Temperature is not a form of energy, but a measure of average kinetic energy of particles in a substance. **(4.2b, 4.2i)**
    a.) Phase changes involve change in potential energy and change in intermolecular distance. **(4.2ii)**

b.) Temperature read on a Celsius thermometer can be converted to Kelvin by adding 273. **(3.4iii)**

6. Heat, measured in joules, can be calculated from both phase changes and temperature changes **(4.2iv)**
7. A reaction can be exothermic, gives off energy, or endothermic, absorbs energy. **(4.1b)**
8. Potential energy diagrams help differentiate between endothermic and exothermic reactions. Reactants, products, activation energy, and heat of reaction are all key parts of a potential energy diagram. **(4.1i, 4.1ii)**

C. Gases

1. An ideal gas is a model used to describe a real gas. This model most closely fits a real gas when the gas is at low pressure and high temperature: examples are H_2 and He. **(3.4a)**
2. The kinetic molecular theory (KMT) is a model that aids in the understanding of the behavior of gases. **(3.4b, 3.4c)** The premises of the kinetic molecular theory are:
    - a.) Gas molecules move randomly at all speeds and in all directions.
    - b.) The size of a gas molecule is much smaller than the distance between the molecules.
    - c.) The gas molecules do not attract or repel each other.
    - d.) The total energy of the system does not change even when molecules collide.
3. The collision theory says that a reaction will have the greatest likelihood of occurring if the particles collide with sufficient energy and correct orientation. **(3.4d)**
4. The combined gas law allows problems to be solved when the number of moles of gas is constant. **(3.4ii)**
5. The gas laws can be explained in terms of the kinetic molecular theory. **(3.4i)**
6. When two gases with equal volumes are at the same temperature and pressure, they also have an equal number of particles **(3.4e)**

IV. The Periodic Table **(CRT)**

A. Each element has a specific location on the periodic table, which indicates the elements physical and chemical properties **(3.1y)**

1. Elements are arranged in the order of increasing atomic number from left to right across each horizontal row - a.k.a. period
2. Elements in each vertical column - a.k.a. family or group - all form compounds with similar chemical formulas and properties because they have the same number of valance electrons **(3.1z)**
3. Physical properties include density, hardness, conductivity, malleability, ductility and solubility **(3.1w)**
4. Chemical properties describe how an element behaves during a chemical reaction **(3.1x)**

B. Elements can be classified as metals, nonmetals, metalloids or noble gases **(3.1v)**

1. Metals – on the left of the stairs
   a.) lose electrons to form positive ions
   b.) the greater the ability to lose electrons the more metallic character
   c.) have low ionization energy and electronegativity
   d.) are deformable - malleable and ductile
   e.) good conductors of heat and electricity
   f.) mostly solids at room temperature
2. Nonmetals – on the right of the stairs
   a.) gain electrons to form negative ions
   b.) the greater the ability to gain electrons the more non-metallic properties
   c.) have high ionization energies and electronegativities
   d.) are brittle
   e.) are poor conductors of heat and electricity
   f.) tend to be gases or liquids at room temperature
3. Metalloids – on the stairs
   a.) include the elements B, Si, Ge, As, Sb, and Te
   b.) have properties of both metals and nonmetals
4. Noble Gases – last family on the right of the table
   a.) very stable
   b.) have complete valance shells
   c.) unreactive
   d.) gases at room temperature

C. Describe the nature of periods and groups of elements in the periodic table, including trends in atomic radius, ionic radius, electronegativity, ionization energy, metallic/ non-metallic properties and reactivity **(3.1 aa, 3.1 bb)**

1. Atomic radius-
   a.) increases going down a family
   b.) decreases going across a period
   c.) factors include nuclear charge, distance between nucleus and outer electrons, and amount of shielding
2. Ionic radius –
   a.) positive ions are smaller than the parent atom and the greater the positive charge in a period the smaller the ionic size
   b.) negative ions are larger than their parent atom and the greater the negative charge across a period the larger the ion
3. Electronegativity and ionization energy –
   a.)decreases as go down a family and increases across a period
4. Metallic/ non-metallic properties –
   a.) metallic properties increase as go down a family and to the left across a period
   b.) non-metallic properties increase as go up a family and towards the right across a period

5. Reactivity –
   a.) the reactivity of metals increases down a family and decreases to the right across the table
   b.) the reactivity of non-metals decreases down a family and increases to the right across a period
   c.) noble gases are mostly unreactive

D. Describe mass number of atoms **(3.1g)**
   1. The sum of the protons and neutrons in an atom is its mass number
   2. The mass number indicated on the table is the weighted average of all naturally occurring isotopes of the element

E. Some elements exist as two or more forms in the same phase which have different properties. These are called allotropes. Examples include O and C **(5.2 f)**

F. Chemistry of selected elements **(3.1z, 3.1v, 3.1y)**
   1. Alkali metals – Group 1 elements; have silvery appearance; are soft; have low melting points; are too reactive to exist in nature as free elements
   2. Alkaline earth metals – Group 2 elements; are denser, harder and stronger than alkali metals and have higher melting points; are too reactive to exist in nature as free elements
   3. Transition metals – Groups 3 – 12 elements; are good conductors of electricity and have a high luster; less reactive than alkali and alkaline earth metals with palladium, platinum and gold amongst the least reactive of all the metals
   4. Halogens – Group 17 elements; the most reactive of the non-metals; react with metals to form salts
   5. Metalloids - brittle solids with some properties of metals and some of non-metals; have an electrical conductivity intermediate of metals and non-metals

V. Chemical reactions and bonding **(CRT)**

A. Elements join to form compounds with different chemical and physical properties **(3.1dd, 5.2iii)**
   1. Dependent on elements involved (metal, nonmetal) and the type of bond formed (ionic, covalent, metallic)

B. Compounds contain elements chemically combined in a fixed proportion (ex/ ammonia always contains one nitrogen atom and three hydrogen atoms) and can only be broken apart by chemical means **(3.1cc)**

C. Compounds have specific chemical formulas, and can be named according to IUPAC rules **(3.1cc)**

D. Chemical formulas can be represented in different forms **(3.1ee, 3.3d)**
   1. Empirical formula is the smallest whole number ratio of the elements in a compound
   2. Molecular formula is the actual ratio of atoms in a molecule of that compound
   3. Structural formula shows the kind, number, arrangement and bonds of the atoms in a molecule

E. There are four basic types of chemical reactions **(3.2b)**
   1. Synthesis – A + B→ AB
   2. Decomposition – AB → A + B
   3. Single replacement – AB + C → AC + B
   4. Double replacement – AB + CD → AD + CB

F. Physical properties are dependent on chemical bonds and intermolecular forces. They include malleability, solubility, conductivity, hardness, melting point and boiling point **(5.2n)**

G. Describe the nature of the different types of chemical bonds and their relationship to valance electrons **(5.2a)**
   1. Ionic – electrons transferred
   2. Covalent – electrons shared
   3. Metallic – mobile sea of electrons shared by all atoms

H. Atoms can gain or lose electrons to become ions **(5.2c)**
   1. An anion is an atom that has gained one or more electrons thus obtaining a negative charge and a larger radius
   2. A cation is an atom that has lost one or more electrons thus obtaining a positive charge and a smaller radius

I. The types of atoms involved in different types of bonds varies **(5.2g, 5.2h)**
   1. Ionic – a metal and a nonmetal - note: ionic compounds which contain polyatomic ions will have both covalent and ionic bonding
   2. Covalent – nonmetals
   3. Metallic – metals

J. Lewis Dot diagrams can be used to represent ionic and covalent bonding by showing the arrangement of valance electrons **(5.2d)**

K. Atoms bond with one another to release energy obtaining a stable valance electron configuration.
   1. The noble gases with a full valance shell are already stable. **(5.2b)**

L. How strongly an atom attracts the electrons in a chemical bond is indicated by its electronegativity **(5.2j)**
   1. Electronegativity is the ability of an atom to attract electrons to itself
   2. Electronegativity is measured on an arbitrary scale
   3. Electronegativity values increase as you move towards the right and up the periodic table

M. The difference in electronegativity between two bonded atoms determines the degree of polarity of the bond. **(5.2k)**
   1. The greater the electronegativity difference the greater the polarity of the bond

N. Molecular polarity can be determined by the shape of the molecule and the distribution of charge. **(5.2l)**
   1. Symmetrical and diatomic molecules are non-polar because electronegativity differences cancel each other out
   2. Asymmetrical molecules are polar because of the difference in electronegativity values

O. Account for the nature and effect of hydrogen bonding **(5.2m)**
  1. Intermolecular force created by the unequal distribution of charge between the hydrogen atom in one molecule with an extremely electronegative element in another molecule

P. Compare the three classes of solids **(5.2a)**
  1. Ionic – brittle, high melting points and boiling points
  2. Molecular – soft, low melting points and boiling points
  3. Metallic – soft, high melting points and boiling points

Q. Explain the role of energy in simple chemical reactions **(4.1b, 4.1c, 4.1d, 5.2i)**
  1. Chemical changes result in the shift of energy
     a.) The formation of bonds is exothermic – energy is released
     b.) The breaking of bonds is endothermic – energy is absorbed
  2. Changes in energy during a reaction can be represented by a potential energy diagram, which shows that the energy absorbed or released is equal to the potential energy difference of the reactants and the products

R. Energy, mass and charge are all conserved in every chemical reaction **(3.3a, 4.1d)**

VI. Math of Chemical Formula and Equations **(CRT)**

A. Balanced chemical equations demonstrate the conservation of matter **(3.3c)**
  1. Chemical equations show a change in the type of substances
  2. Reactants are always on the left side of the arrow and the products are on the right side of the arrow
  3. Conservation of mass –the total mass of reactants is equal to the total mass of the products
  4. Conservation of atoms: for each element, atoms of reactants are equal to atoms of product
  5. The number of molecules multiplied by the number atoms of that element per molecule equals the total number of atoms
  6. Equations can be balanced using the inspection method

B. Mole ratios of a reaction are shown by coefficients in the balanced chemical equation **(3.3c)**
  1. Coefficients indicate the number of moles of individual particles in the equation
  2. Changes in the number of particles are directly proportional to one another.

C. Determining the atomic mass of elements and the formula mass of Compounds **(3.3e)**
  1. Use the periodic table to determine the atomic mass of an atom
  2. Use proper unit, the amu, to label an atom's mass
  3. Use subscripts and atomic mass to calculate the formula mass of compounds (including hydrates)

D. Calculate the mass of a given number of moles of a substance **(3.3e, 3.3vi)**
  1. Perform calculations using units and significant digits to justify a reasonable answer
  2. Use scientific notation in calculations

E. Calculate the number of moles of a given mass **(3.3ix)**
    1. Calculations using units and significant digits to justify a reasonable answer (dimensional analysis)
F. Calculate the number of molecules in a given no. of moles
    1. (Introduce Avogadro's Number "6.02×10^{23}" particles per mole)
    2. (Convert from moles to atoms, ions or molecules)
G. (Calculate the volume of a given mass of gas at STP)
    1. (Introduce molar volume for gases at STP)
    2. (A value of 22.4 liters per mole at STP is an approximate value for most gases)
    3. (The exact molar volume for different gas substances at STP do vary slightly)
    4. (Use complex mole conversions of mass to moles to liters)
H. Determine % Composition, empirical and molecular formulas **(3.3f, 3.1ee)**
    1. % Composition for mixtures may vary
    2. % Composition for a compound is constant
    3. % Composition may be calculated using mass relationships or volume relationships.
    4. An empirical formula is the lowest whole number relationship between atoms of various elements in a compound
    5. A molecular formula is a whole number multiple of the empirical formula
I. Derive quantitative information about reactants and products in a chemical reaction
    1. Coefficients in an equation may be interpreted with different units atoms -molecules or moles
J. (Solve mass – mass, mass – vol., and vol. – vol. Problems)
    1. (Recognize the given amount of substance in a word problem)
    2. (Recognize the information requested -the find)
    3. (Relate the given and find to a balanced chemical equation)
    4. (Use dimensional analysis to calculate the requested information.)

VII. Energy of Chemical Reactions **(CRT)**
A. Chemical Kinetics and Thermodynamics
    1. Chemical Kinetics relates to movement of reacting particles
    2. Thermodynamics relates to transfer of energy between kinetic energy and potential energy.
    3. Explain the collision theory of reactions **(3.4d)**
        a.) Chemical change may occur when reactant particles collide.
        b.) The relative particle velocity must be large to produce an effective collision.
        c.) The particles must collide with proper geometry to product a chemical change

4. Account for the factors that affect reaction rates: nature of reactants, surface area, concentration, temperature of reacting system, presence of a catalyst **(3.4f, 3.4g)**
    - a.)Nature of reactants
        - i.) Stable molecules have low enthalpies
        - ii) Stable molecules require greater molecular velocities to react
        - iii) Unstable molecules have large enthalpies
        - iv.) Unstable molecules require lower molecular velocities to react.
        - v.) Translational kinetic energy is responsible for particle collisions.
        - vi.) Solutions, liquids, and gases possess translational kinetic energy
    - b.) Surface area
        - i.) Solids only react on their surface
        - ii.) Smaller solid particles have more surface area per gram
    - c.) Concentration
        - i.) Solid and liquid reactants have a constant concentration
        - ii.) Solutions and gases have variable concentrations
        - iii.) The greater the concentration of mobile particles the greater the probability of collision.
    - d.) Temperature
        - i.) Temperature is directly related to particle velocity (translational kinetic energy)
        - ii.) Temperature equals the Average kinetic energy of the system
        - iii.) An increase in temperature increases the rate of ***ALL*** chemical reactions
        - iv.) Higher temperature causes a greater frequency of collisions
        - v.) Higher temperature causes greater percentage of collisions occurring with sufficient energy
    - e.) Catalyst
        - i.) A catalyst lowers the activation energy required to start the reaction
        - ii.) A catalyst increases the rate of all chemical reactions
        - iii.) A catalyst is not consumed during a chemical reaction
5. Interpret potential energy diagrams and energy distribution diagrams **(4.1c, 4.1d)**
    - a.) Potential energy diagrams
        - i.) A transfer of energy between kinetic and potential occurs during chemical reactions
        - ii.) A transfer from kinetic energy to potential energy is described as an absorption of energy and the reaction is labeled an endothermic reaction with respect to energy

iii.) A transfer from potential energy to kinetic energy is described as a release of energy and the reaction is labeled an exothermic reaction with respect to energy

iv.) The symbol "ΔH" the change in enthalpy is the scientific notation for changes in energy

v.) ΔH is positive for endothermic reactions

vi.) ΔH is negative for exothermic reactions

vii.) $\Delta H_{reaction} = [H_{products}] - [H_{reactants}]$

viii.) A graph may be used to show energy changes during a chemical reaction. The axis are labeled potential energy and reaction pathway

b.) Energy distribution diagrams

i.) At a uniform temperature molecules in the same sample have different kinetic energies

ii.) At lower temperatures only a few molecules have high kinetic energy

iii.) As temperature increases a greater number of molecules in the sample have high kinetic energy

iv.) (A graph is used to show the distribution of kinetic energy its axis are labeled fraction of molecule and kinetic energy)

6. Interpret significance of changes of heat in chemical or physical changes **(4.1d)**

a.) If the temperature of a system increases during a chemical or physical change ΔH is negative – heat is released

b.) If the temperature of a system decreases during a chemical or physical change ΔH is positive –heat is absorbed

7. Describe the role of changes in entropy on chemical and physical changes **(3.1ll, 3.1mm)**

a.) Entropy is randomness or disorder

b.)ΔS is the symbol for the change in entropy

c.) High temperature favors large entropies

d.) Low temperature favors low entropies

e.) Pure substances have low entropy, mixtures have high entropy

f.) Entropy is related to the phase of matter-Solids have low entropy –Gases have high entropy

g.) An increase in total entropy of a system favors a spontaneous change

B. Chemical Equilibrium

1. (Equilibria may be static or dynamic)
2. Chemical Equilibria are dynamic
3. During chemical equilibrium the rate of a forward reaction is equal to the rate of a reverse reaction

4. Distinguish between a reversible reaction at equilibrium and a nonreversible reaction **(3.4h, 3.4i)**
   - a.) Nonreversible reaction
     - i.) If products are very stable or if a product is removed from the reaction the reaction is not reversible
     - ii.) A nonreversible reaction is described as "going to completion"
     - iii.) Reaction goes to completion because: water forms; gas forms; or an insoluble product forms
     - iv.) When one reactant is consumed the reaction stops
   - b.) Reversible reactions
     - i.) Products change back into reactants during reversible reactions
     - ii.) A double arrow is used to show reversible reactions
     - iii.) At equilibrium the two arrows of the double arrow have equal lengths
     - iv.) (Equilibrium constants K_{eq} are used to show the position of equilibrium)
     - v.) (Large K_{eq} values favor products)
     - vi.) (Small K_{eq} values favor reactants)
     - vii.) (K_{eq} values are $\frac{[products]}{[reactants]}$)

5. Explain Le Chatelier's Principle in terms of pressure, volume, concentration and temperature **(3.4j)**
   - a.) Pressure
     - i.) For solid and liquid systems changes in pressure have negligible effect
     - ii.) Pressure affects some systems containing gases
     - iii.) Systems where molar volume of gas reactants is not equal to molar volume of gas products are affected
     - iv.) An increase in pressure favors the side with smaller molar volume
     - v.) A decrease in pressure favors the side with the larger molar volume
   - b.) Volume
     - i.) Changes in volume only affect gas systems where the molar volume of gas reactants is not equal to the molar volume of gas products

- c.) Concentration
  - i.) A change in concentration of one item -a stress- in the equilibrium destroys the equilibrium
  - ii.) A stressed system will readjust to re-establish a new equilibrium
  - iii.) (K_{eq} will remain the same when a concentration change is made)
  - iv.) Explain changes in terms of forward rate and reverse rate
- d.) Temperature
  - i.) (A temperature change causes the K_{eq} value to change)
  - ii.) Increases in temperature favor endothermic reactions
  - iii.) Decreases in temperature favor exothermic reaction

VIII. Acids, Bases, and Salts **(CRT)**

A. Defining Acids, Bases, and Salts

1. Arrhenius acids and bases are electrolytes. **(3.1 uu)**
2. An electrolyte, dissolved in water, has the ability to conduct electricity. The ability to conduct electricity depends on the concentration of the ions. **(3.1 rr)**
3. Arrhenius acids yield hydronium ions a form of hydrogen ion, , as the only positive ion in an aqueous solution. **(3.1 vv)**
4. Arrhenius bases yield hydroxide ions as the only negative ion in an aqueous solution. **(3.1 ww)**
5. Describe properties of Acids, Bases, and Salts.
   - a.) Given properties, identify substances as Arrhenius acid or Arrhenius base. **(3.1xxxi)**
6. Bronted-Lowry theory states that an acid is a proton donor and a base is a proton acceptor. **(3.1 yy)**
7. The acidity or alkalinity of a solution can be measured by its pH value. **(3.1 ss)**
8. On the pH scale, each decrease of one unit of pH indicates a tenfold increase in the hydronium ion concentration. **(3.1 tt)**
   - a.) Identify solution as acid, base, or neutral based upon the pH. **(3.1xxxii)**
9. The relative level of acidity in a solution can be shown with indicators. **(3.1 ss)**
   - a.) Interpret changes in acid-base indicator color. **(3.1xxxiii)**
10. (An amphoteric substance is one that can act either as an acid or a base, depending on its chemical environment.)

B. Reactions of Acids and Bases

1. In the process of neutralization, an acid and a base react to form a salt and water. **(3.1 xx)**
   - a.) Write simple neutralization reactions when given the reactants. **(3.1xxxiv)**

2. Titration is a process in which a volume of solution of known concentration is used to determine the concentration of another solution. **(3.1 zz)**
3. Calculate the molarity of an acid or base.
4. Solve titration problems.
    a.) Calculate the concentration, or volume of a solution, using titration data. **(3.1xxxv)**

IX. Oxidation-Reduction **(CRT)**

A. Oxidation-Reduction Reactions (REDOX)

1. Oxidation numbers can be assigned to atoms and ions. Changes in oxidation numbers indicate that oxidation and reduction have occurred. **(3.2 i)**
2. An oxidation-reduction reaction involves the transfer of electrons. **(3.2d)**
    a.) Determine a missing reactant or product in a balance equation. **(3.2iii)**
3. Oxidation is the loss of electrons and Reduction is the gain of electrons. **(3.2 e 3.2 g)**
4. The particle oxidized is the reducing agent, and the particle reduced is the oxidizing agent.
5. Half reactions can be written to represent both oxidation and reduction. **(3.2 f 3.2 h)**
    a.) Write and balance half reactions for oxidation and reduction of free elements and their monatomic ions. **(3.2vi)**
6. In a redox reaction the number of electrons lost must equal the number of electrons gained. **(3.3 b)**
7. (Balance redox reactions by identifying changes in oxidation number writing half reactions.)

A. Electrochemistry

1. An electrochemical cell can be either voltaic or electrolytic. **(3.2 j)**
    a.) Compare and contrast voltaic and electrolytic cells. **(3.2ix)**
2. Describe the assembly and operation of voltaic cell.
    a.) Cathode, anode, salt bridge, and direction of electron flow given the reaction equation. **(3.2vii)**
    b.) Use an activity series to determine whether a redox reaction is spontaneous. **(3.2x)**
3. A voltaic cell spontaneously converts chemical energy into electrical energy. **(3.2 k)**
4. In a voltaic cell, oxidation occurs at the anode and reduction occurs at the cathode. **(3.2 j)**
5. An electrolytic cell requires electrical energy to produce a chemical change. This process is electrolysis. **(3.2 l)**
6. Describe the assembly and operation of an electrolytic cell.
    a.) Cathode, anode, and direction of electron flow given the reaction. **(3.2viii)**

X. Nuclear Chemistry **(CRT)**

A. Nuclear reactions – the energy released during nuclear reactions is much greater than the energy released during chemical reactions **(5.3c)**

1. Radioactive decay **(3.1o)**
    - a.) The stability of an isotope is based on the ratio of neutrons to protons in its nucleus or the binding energy per nucleon.
    - b.) Nuclei that are unstable spontaneously decay emitting radiation.
2. Natural and artificial transmutations or conversion from one element to another **(4.4a, 5.3b)**
    - a.) Natural transmutation occurs because of an unstable ratio of protons to neurons in the nucleus. Modes include: alpha decay, beta decay, positron emissions and gamma radiation
    - b.) Artificial transmutation occurs by bombarding the nucleus with high energy particles
3. Determining decay mode **(3.1p, 3.1ix)**
    - a.) Spontaneous decay can involve the release of alpha particles, beta particles, positrons, and/or gamma radiation with these emissions differing in mass, charge, ionizing power and penetrating power
4. Write nuclear equations showing alpha and beta decay **(4.4c, 4.4iii)**
    - a.) Nuclear reactions can be represented by equations that include symbols which represent atomic nuclei - with the mass number and atomic number, subatomic particles - with mass number and charge, and/or emissions such as gamma radiation
    - b.) By completing nuclear equations missing particles can be predicted
5. Fission **(4.4b, 4.4ii, 5.3b, 4.4f)**
    - a.) Involves the splitting of a heavy nucleus with a neutron to produce two lighter nuclei, one or more neutrons, and a conversion of mass to energy
    - b.) Products are radioactive with long half-lives
6. Fusion **(4.4b, 4.4ii, 5.3b, 4.4f)**
    - a.) Involves the combining of two lighter weight nuclei to form one heavy nuclei and a release of energy
    - b.) Limitation is that it requires extremely high pressure and temperature - ex/ sun
7. Half-life and calculations **(4.4a, 4.4i)**
    - a.) Each radioisotope has its own rate of decay - half-life
    - b.) Any of the following can be determined given two of the three variables: initial amount of isotope present, the fraction of isotope remaining after a given amount of time, or the half-life of the isotope

B. The risks associated with radioactivity **(4.4e, 4.4f)**

1. Biological exposure – can damage normal tissue and cause mutations that can be passed from generation to generation
2. Long term storage and disposal– Fission reactions create radioactive byproducts with very long half-lives that are difficult to dispose of – most are buried and there is the possibility of leakage
3. Nuclear accidents – possibility of power plant accidents like Chernobyl that poses a threat by releasing radioactive material into the air or water

C. Uses of radioactive isotopes **(4.4d, 4.4iv, 4.4f)**

1. Medicine – treatment and diagnostic purposes
    - a.) I-131 – used for diagnosing and treating thyroid disorders
    - b.) Co-60 – used for killing cancer cells
    - c.) (Tc-99 – used to diagnose cancer cells)
    - d.) (Co-60 and Cs-137 – used to kill bacteria like anthrax bacilli medically and other bacteria on food to extend shelf life)
2. Radioactive dating – age of a specimen can be determine by the isotopes half-life
    - a.) C-14/C-12 ratio used to date previously living material
    - b.) U-238/ Pb-206 ratio used to date rocks and other geological formations
3. Biological and chemical tracers – a radioisotope used to follow the path of a material in a system
    - a.) (Biological - P-31and C-14 used to trace pathways in plants)
    - b.) (Chemical – Radiation products can be used to determine the thickness of materials such a plastic wrap or to test the strength of a weld)
4. Nuclear power – used to produce electricity with fission reactions

XI. Organic Chemistry **(CRT)**

A. Organic

1. The nature of carbon and its ability to bond to itself in chains, rings and networks. **(3.1ff)**
    - a.) Carbon is a nonmetal that forms four covalent bonds
    - b.) Carbon may bond to itself to form long molecules
    - c.) Carbon may bond to itself to form rings for example benzene and cycloalkanes
    - d.) Carbon may form network solids (diamond)
2. The IUPAC system of nomenclature. **(3.1ff)**
    - a.) The IUPAC system uses suffixes to classify types of molecules
    - b.) The IUPAC system uses prefixes to indicate the number of attached groups
    - c.) The IUPAC system uses Arabic numbers to indicate the point of attachment of a subgroup

3. Structural and condensed structural formulas. **(3.1xvii)**
    - a.) A structural formula shows the exact position of all atoms in a molecule
    - b.) A structural formula shows the type of bond between the atoms of a molecule
    - c.) A condensed structural formula uses subgroups and functional groups to describe a molecule – $CH_3(CH_2)_5COOH$

B. Hydrocarbons

1. General characteristics
    - a.) Compounds that contain only hydrogen and carbon
    - b.) Nonpolar molecules
    - c.) Main source is crude oil
2. Saturated hydrocarbons; alkanes. **(3.1gg)**
    - a.) Contain only single covalent bonds between carbon atoms
    - b.) The name of the molecule ends in "ane"
    - c.) The number of carbon atoms determines the root name
    - d.) This group is the least reactive organic family
3. Unsaturated hydrocarbons; alkenes, alkynes. **(3.1gg, 5.2e)**
    - a.) Each molecule has one multiple bond between carbon atoms
    - b.) Alkenes are named to end in "ene" which shows one double bond between carbon atoms
    - c.) Alkynes are named to end in "yne" which shows one triple bond between carbon atoms
    - d.) Chemical activity increases as the number of multiple bonds between two carbon atoms increases.
4. Structural formulas for alkanes, alkenes and alkynes containing a maximum of ten carbon atoms. **(3.1xxi)**
    - a.) Slash structural formulas illustrate the shape of each molecule

C. Other organic compounds **(3.1hh, 3.1xvii, 3.1xx)**

1. Organic acids **(3.1hh, 3.1xvii, 3.1xx)**
    - a.) Functional groups is the carboxyl R-COOH
    - b.) IUPAC ending is "oic"
    - c.) Some organic acids are electrolytes
2. Alcohols **(3.1hh, 3.1xvii, 3.1xx)**
    - a.) Functional group is the hydroxyl R-OH
    - b.) IUPAC ending is "ol"
    - c.) Alcohols are used as anti freeze
3. Esters **(3.1hh, 3.1xvii, 3.1xx)**
    - a.) Functional group is the ester R-COOC-R
    - b.) IUPAC ending is "oate"
    - c.) Esters are named from the alcohol and the acid of their synthesis, an alcohol akyl plus acid ion
    - d.) Esters are used in perfumes and as food flavors
4. Aldehydes **(3.1hh, 3.1xvii, 3.1xx)**
    - a.) Functional group is the aldehyde R-CHO
    - b.) IUPAC ending is "al"

5. Ketones **(3.1hh, 3.1xvii, 3.1xx)**
    a.) Functional group is the R-CO-R
    b.) The IUPAC ending is "one"
    c.) Acetone is a major component of finger nail polish remover
6. Ethers **(3.1hh, 3.1xvii, 3.1xx)**
    a.) The functional group is the R-O-R
    b.) The IUPAC naming uses the two akyl groups plus ether example methyl ethyl ether CH_3-O-CH_2CH_3
    c.) (Ether is used in spray paint solvents and anesthesia agents)
7. Halides **(3.1hh, 3.1xvii, 3.1xx)**
    a.) Organic molecules that contain halogens
    b.) They are named as prefixes on a carbon chain
8. Amines **(3.1hh, 3.1xvii, 3.1xx)**
9. Amides **(3.1hh, 3.1xvii, 3.1xx)**
10. Amino Acids **(3.1hh, 3.1xvii, 3.1xx)**

D. Isomers **(3.1ii)**
1. Compounds that have the same molecular formula but different structural formulas
2. Isomers will have different IUPAC names
3. Isomers will have different physical properties and different chemical properties

E. Types of organic reactions **(3.2c, 3.2iv, 3.2iii)**
1. Addition **(3.2c, 3.2iv, 3.2iii)**
    a.) Alkene and Alkyne molecules undergo addition reactions
    b.) One of the multiple bonds break during addition
    c.) Two atoms are attached to the molecule, one to each of the adjacent carbons involved with the multiple bond
    d.) Only one product is formed
    e.) The name of the parent molecule changes
2. Substitution **(3.2c, 3.2iv, 3.2iii)**
    a.) Alkanes undergo substitution reactions
    b.) One hydrogen is removed and is replaced by a different type of atom
    c.) Two products result
3. Polymerization **(3.2c, 3.2iv, 3.2iii)**
    a.) Uses monomers to build large molecules
    b.) A monomer repetitively links to itself forming a long chain
    c.) A condensed formula uses the subscript "n" to show repetition
    d.) (Natural polymers are: starch, protein, cellulose)
    e.) (Synthetic polymers are plastics: polyethylene, polystrene, polypropylene, polytetrafluoroethylene and polyacrilonitrile)
    f.) (Two types of polymerization are: condensation and addition)

4. Esterification **(3.2c, 3.2iv, 3.2iii)**
    a.) Making an ester
    b.) Alcohol + Acid → Ester + Water
    c.) This is a dehydration synthesis reaction
    d.) Esters occur naturally or synthetically
5. Fermentation **(3.2c, 3.2iv, 3.2iii)**
    a.) A natural process
    b.) Yeast provides the enzyme to allow the process to be spontaneous
    c.) Carbohydrate $\xrightarrow{\text{Zymase}}$ ethyl alcohol + carbon dioxide
    d.) Fermentation must occur in the absence of molecular oxygen if alcohol is to be produced
    e.) Alcohol may be used as a fuel
6. Saponification **(3.2c, 3.2iv, 3.2iii)**
    a.) Soap making
    b.) Fat + strong base ⟶ 3 Soaps + glycerol
    c.) Soaps are long molecules, one end is polar the other end is nonpolar
    d.) Soaps bind nonpolar dirt molecules to polar water molecules
7. Combustion **(3.2c, 3.2iv, 3.2iii)**
    a.) Is burning
    b.) Is an exothermic process releasing light energy and heat energy
    c.) Complete combustion of hydrocarbons produce CO_2 and H_2O
    d.) Carbon monoxide is produced if limited amounts of oxygen occur