

APPENDIX A

CHEMISTRY CORE TOPICS

This section contains ten topic areas in which the major understandings found in the core are sorted by content topic. These ten topic areas may be used for ease in curriculum writing; however, they do not connote a suggested scope and sequence.

I. Atomic Concepts

- I.1 The modern model of the atom has evolved over a long period of time through the work of many scientists. (3.1a)
- I.2 Each atom has a nucleus, with an overall positive charge, surrounded by one or more negatively charged electrons. (3.1b)
- I.3 Subatomic particles contained in the nucleus include protons and neutrons. (3.1c)
- I.4 The proton is positively charged, and the neutron has no charge. The electron is negatively charged. (3.1d)
- I.5 Protons and electrons have equal but opposite charges. The number of protons equals the number of electrons in an atom. (3.1e)
- I.6 The mass of each proton and each neutron is approximately equal to one atomic mass unit. An electron is much less massive than a proton or a neutron. (3.1f)
- I.7 In the wave-mechanical model (electron cloud model), the electrons are in orbitals, which are defined as the regions of the most probable electron location (ground state). (3.1h)
- I.8 Each electron in an atom has its own distinct amount of energy. (3.1i)
- I.9 When an electron in an atom gains a specific amount of energy, the electron is at a higher energy state (excited state). (3.1j)
- I.10 When an electron returns from a higher energy state to a lower energy state, a specific amount of energy is emitted. This emitted energy can be used to identify an element. (3.1k)
- I.11 The outermost electrons in an atom are called the valence electrons. In general, the number of valence electrons affects the chemical properties of an element. (3.1l)
- I.12 Atoms of an element that contain the same number of protons but a different number of neutrons are called isotopes of that element. (3.1m)
- I.13 The average atomic mass of an element is the weighted average of the masses of its naturally occurring isotopes. (3.1n)

II. Periodic Table

- II.1** The placement or location of elements on the Periodic Table gives an indication of physical and chemical properties of that element. The elements on the Periodic Table are arranged in order of increasing atomic number. (3.1y)
- II.2** The number of protons in an atom (atomic number) identifies the element. The sum of the protons and neutrons in an atom (mass number) identifies an isotope. Common notations that represent isotopes include: $^{14}_6\text{C}$, ^{14}C , carbon-14, C-14. (3.1g)
- II.3** Elements can be classified by their properties and located on the Periodic Table as metals, nonmetals, metalloids (B, Si, Ge, As, Sb, Te), and noble gases. (3.1v)
- II.4** Elements can be differentiated by their physical properties. Physical properties of substances, such as density, conductivity, malleability, solubility, and hardness, differ among elements. (3.1w)
- II.5** Elements can be differentiated by chemical properties. Chemical properties describe how an element behaves during a chemical reaction. (3.1x)
- II.6** Some elements exist in two or more forms in the same phase. These forms differ in their molecular or crystal structure, and hence in their properties. (5.2f)
- II.7** For Groups 1, 2, and 13-18 on the Periodic Table, elements within the same group have the same number of valence electrons (helium is an exception) and therefore similar chemical properties. (3.1z)
- II.8** The succession of elements within the same group demonstrates characteristic trends: differences in atomic radius, ionic radius, electronegativity, first ionization energy, metallic/nonmetallic properties. (3.1aa)
- II.9** The succession of elements across the same period demonstrates characteristic trends: differences in atomic radius, ionic radius, electronegativity, first ionization energy, metallic/nonmetallic properties. (3.1bb)

III. Moles/Stoichiometry

- III.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 be broken down by chemical means. A chemical compound can be represented by a specific chemical formula and assigned a name based on the IUPAC system. (3.1cc)
- III.2** Types of chemical formulas include empirical, molecular, and structural. (3.1ee)
- III.3** The empirical formula of a compound is the simplest whole-number ratio of atoms of the elements in a compound. It may be different from the molecular formula, which is the actual ratio of atoms in a molecule of that compound. (3.3d)
- III.4** In all chemical reactions there is a conservation of mass, energy, and charge. (3.3a)
- III.5** A balanced chemical equation represents conservation of atoms. The coefficients in a balanced chemical equation can be used to determine mole ratios in the reaction. (3.3c)
- III.6** The formula mass of a substance is the sum of the atomic masses of its atoms. The molar mass (gram-formula mass) of a substance equals one mole of that substance. (3.3e)

III.7 The percent composition by mass of each element in a compound can be calculated mathematically. (3.3f)

III.8 Types of chemical reactions include synthesis, decomposition, single replacement, and double replacement. (3.2b)

IV. Chemical Bonding

IV.1 Compounds can be differentiated by their chemical and physical properties. (3.1dd)

IV.2 Two major categories of compounds are ionic and molecular (covalent) compounds. (5.2g)

IV.3 Chemical bonds are formed when valence electrons are (5.2a):

- transferred from one atom to another (ionic)
- shared between atoms (covalent)
- mobile within a metal (metallic)

IV.4 In a multiple covalent bond, more than one pair of electrons are shared between two atoms. (5.2e)

IV.5 Molecular polarity can be determined by the shape of the molecule and the distribution of charge. Symmetrical (nonpolar) molecules include CO_2 , CH_4 and diatomic elements. Asymmetrical (polar) molecules include HCl , NH_3 , and H_2O . (5.2l)

IV.6 When an atom gains one or more electrons, it becomes a negative ion and its radius increases. When an atom loses one or more electrons, it becomes a positive ion and its radius decreases. (5.2c)

IV.7 When a bond is broken, energy is absorbed. When a bond is formed, energy is released. (5.2i)

IV.8 Atoms attain a stable valence electron configuration by bonding with other atoms. Noble gases have stable valence configurations and tend not to bond. (5.2b)

IV.9 Physical properties of substances can be explained in terms of chemical bonds and intermolecular forces. These properties include conductivity, malleability, solubility, hardness, melting point, and boiling point. (5.2n)

IV.10 Electron-dot diagrams (Lewis structures) can represent the valence electron arrangement in elements, compounds, and ions. (5.2d)

IV.11 Electronegativity indicates how strongly an atom of an element attracts electrons in a chemical bond. Electronegativity values are assigned according to arbitrary scales. (5.2j)

IV.12 The electronegativity difference between two bonded atoms is used to assess the degree of polarity in the bond. (5.2k)

IV.13 Metals tend to react with nonmetals to form ionic compounds. Nonmetals tend to react with other nonmetals to form molecular (covalent) compounds. Ionic compounds containing polyatomic ions have both ionic and covalent bonding. (5.2h)

V. Physical Behavior of Matter

- V.1 Matter is classified as a pure substance or as a mixture of substances. (3.1q)
- V.2 The three phases of matter (solids, liquids, and gases) have different properties. (3.1kk)
- V.3 A pure substance (element or compound) has a constant composition and constant properties throughout a given sample, and from sample to sample. (3.1r)
- V.4 Elements are substances that are composed of atoms that have the same atomic number. Elements cannot be broken down by chemical change. (3.1u)
- V.5 Mixtures are composed of two or more different substances that can be separated by physical means. When different substances are mixed together, a homogeneous or heterogeneous mixture is formed. (3.1s)
- V.6 The proportions of components in a mixture can be varied. Each component in a mixture retains its original properties. (3.1t)
- V.7 Differences in properties such as density, particle size, molecular polarity, boiling point and freezing point, and solubility permit physical separation of the components of the mixture. (3.1nn)
- V.8 A solution is a homogeneous mixture of a solute dissolved in a solvent. The solubility of a solute in a given amount of solvent is dependent on the temperature, the pressure, and the chemical natures of the solute and solvent. (3.1oo)
- V.9 The concentration of a solution may be expressed as molarity (M), percent by volume, percent by mass, or parts per million (ppm). (3.1pp)
- V.10 The addition of a nonvolatile solute to a solvent causes the boiling point of the solvent to increase and the freezing point of the solvent to decrease. The greater the concentration of particles, the greater the effect. (3.1qq)
- V.11 Energy can exist in different forms, such as chemical, electrical, electromagnetic, thermal, mechanical, and nuclear. (4.1a)
- V.12 Heat is a transfer of energy (usually thermal energy) from a body of higher temperature to a body of lower temperature. Thermal energy is the energy associated with the random motion of atoms and molecules. (4.2a)
- V.13 Temperature is a measurement of the average kinetic energy of the particles in a sample of material. Temperature is not a form of energy. (4.2b)
- V.14 The concept of an ideal gas is a model to explain the behavior of gases. A real gas is most like an ideal gas when the real gas is at low pressure and high temperature. (3.4a)
- V.15 Kinetic molecular theory (KMT) for an ideal gas states that all gas particles (3.4b):
1. are in random, constant, straight-line motion.
 2. are separated by great distances relative to their size; the volume of the gas particles is considered negligible.
 3. have no attractive forces between them.
 4. have collisions that may result in the transfer of energy between gas particles, but the total energy of the system remains constant.

- V16 Collision theory states that a reaction is most likely to occur if reactant particles collide with the proper energy and orientation. (3.4d)
- V17 Kinetic molecular theory describes the relationships of pressure, volume, temperature, velocity, and frequency and force of collisions among gas molecules. (3.4c)
- V18 Equal volumes of different gases at the same temperature and pressure contain an equal number of particles. (3.4e)
- V19 The concepts of kinetic and potential energy can be used to explain physical processes that include: fusion (melting), solidification (freezing), vaporization (boiling, evaporation), condensation, sublimation, and deposition. (4.2c)
- V20 A physical change results in the rearrangement of existing particles in a substance. A chemical change results in the formation of different substances with changed properties. (3.2a)
- V21 Chemical and physical changes can be exothermic or endothermic. (4.1b)
- V22 The structure and arrangement of particles and their interactions determine the physical state of a substance at a given temperature and pressure. (3.1jj)
- V23 Intermolecular forces created by the unequal distribution of charge result in varying degrees of attraction between molecules. Hydrogen bonding is an example of a strong intermolecular force. (5.2m)
- V24 Physical properties of substances can be explained in terms of chemical bonds and intermolecular forces. These properties include conductivity, malleability, solubility, hardness, melting point, and boiling point. (5.2n)

VI. Kinetics/Equilibrium

- VI1 Collision theory states that a reaction is most likely to occur if reactant particles collide with the proper energy and orientation. (3.4d)
- VI2 The rate of a chemical reaction depends on several factors: temperature, concentration, nature of reactants, surface area, and the presence of a catalyst. (3.4f)
- VI3 Some chemical and physical changes can reach equilibrium. (3.4h)
- VI4 At equilibrium the rate of the forward reaction equals the rate of the reverse reaction. The measurable quantities of reactants and products remain constant at equilibrium. (3.4i)
- VI5 LeChatelier's principle can be used to predict the effect of stress (change in pressure, volume, concentration, and temperature) on a system at equilibrium. (3.4j)
- VI6 Energy released or absorbed by a chemical reaction can be represented by a potential energy diagram. (4.1c)
- VI7 Energy released or absorbed during a chemical reaction (heat of reaction) is equal to the difference between the potential energy of the products and the potential energy of the reactants. (4.1d)
- VI8 A catalyst provides an alternate reaction pathway, which has a lower activation energy than an uncatalyzed reaction. (3.4g)

VI.9 Entropy is a measure of the randomness or disorder of a system. A system with greater disorder has greater entropy. (3.1ll)

VI.10 Systems in nature tend to undergo changes toward lower energy and higher entropy. (3.1mm)

VII. Organic Chemistry

VII.1 Organic compounds contain carbon atoms which bond to one another in chains, rings, and networks to form a variety of structures. Organic compounds can be named using the IUPAC system. (3.1ff)

VII.2 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. (3.1gg)

VII.3 Organic acids, alcohols, esters, aldehydes, ketones, ethers, halides, amines, amides, and amino acids are categories of organic molecules that differ in their structures. Functional groups impart distinctive physical and chemical properties to organic compounds. (3.1hh)

VII.4 Isomers of organic compounds have the same molecular formula but different structures and properties. (3.1ii)

VII.5 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. (5.2e)

VII.6 Types of organic reactions include: addition, substitution, polymerization, esterification, fermentation, saponification, and combustion. (3.2c)

VIII. Oxidation-Reduction

VIII.1 An oxidation-reduction (redox) reaction involves the transfer of electrons (e^-). (3.2d)

VIII.2 Reduction is the gain of electrons. (3.2e)

VIII.3 A half-reaction can be written to represent reduction. (3.2f)

VIII.4 Oxidation is the loss of electrons. (3.2g)

VIII.5 A half-reaction can be written to represent oxidation. (3.2h)

VIII.6 In a redox reaction the number of electrons lost is equal to the number of electrons gained. (3.3b)

VIII.7 Oxidation numbers (states) can be assigned to atoms and ions. Changes in oxidation numbers indicate that oxidation and reduction have occurred. (3.2i)

VIII.8 An electrochemical cell can be either voltaic or electrolytic. In an electrochemical cell, oxidation occurs at the anode and reduction at the cathode. (3.2j)

VIII.9 A voltaic cell spontaneously converts chemical energy to electrical energy. (3.2k)

VIII.10 An electrolytic cell requires electrical energy to produce chemical change. This process is known as electrolysis. (3.2l)

IX. Acids, Bases, and Salts

- IX.1 Behavior of many acids and bases can be explained by the Arrhenius theory. Arrhenius acids and bases are electrolytes. (3.1uu)
- IX.2 An electrolyte is a substance which, when dissolved in water, forms a solution capable of conducting an electric current. The ability of a solution to conduct an electric current depends on the concentration of ions. (3.1rr)
- IX.3 Arrhenius acids yield H^+ (aq), hydrogen ion as the only positive ion in an aqueous solution. The hydrogen ion may also be written as H_3O^+ (aq), hydronium ion. (3.1vv)
- IX.4 Arrhenius bases yield OH^- (aq), hydroxide ion as the only negative ion in an aqueous solution. (3.1ww)
- IX.5 In the process of neutralization, an Arrhenius acid and an Arrhenius base react to form a salt and water. (3.1xx)
- IX.6 Titration is a laboratory process in which a volume of solution of known concentration is used to determine the concentration of another solution. (3.1zz)
- IX.7 There are alternate acid-base theories. One theory states that an acid is an H^+ donor and a base is an H^+ acceptor. (3.1yy)
- IX.8 The acidity or alkalinity of a solution can be measured by its pH value. The relative level of acidity or alkalinity of a solution can be shown by using indicators. (3.1ss)
- IX.9 On the pH scale, each decrease of one unit of pH represents a tenfold increase in hydronium ion concentration. (3.1tt)

X. Nuclear Chemistry

- X.1 Stability of isotopes is based on the ratio of neutrons and protons in its nucleus. Although most nuclei are stable, some are unstable and spontaneously decay, emitting radiation. (3.1o)
- X.2 Each radioactive isotope has a specific mode and rate of decay (half-life). (4.4a)
- X.3 A change in the nucleus of an atom that converts it from one element to another is called transmutation. This can occur naturally or can be induced by the bombardment of the nucleus by high-energy particles. (5.3a)
- X.4 Spontaneous decay can involve the release of alpha particles, beta particles, positrons and/or gamma radiation from the nucleus of an unstable isotope. These emissions differ in mass, charge, and ionizing power, and penetrating power. (3.1p)
- X.5 Nuclear reactions include natural and artificial transmutation, fission, and fusion. (4.4b)
- X.6 There are benefits and risks associated with fission and fusion reactions. (4.4f)
- X.7 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. (4.4c).

- X.8 Energy released in a nuclear reaction (fission or fusion) comes from the fractional amount of mass converted into energy. Nuclear changes convert matter into energy. (5.3b)
- X.9 Energy released during nuclear reactions is much greater than the energy released during chemical reactions. (5.3c)
- X.10 There are inherent risks associated with radioactivity and the use of radioactive isotopes. Risks can include biological exposure, long-term storage and disposal, and nuclear accidents. (4.4e)
- X.11 Radioactive isotopes have many beneficial uses. Radioactive isotopes are used in medicine and industrial chemistry, e.g., radioactive dating, tracing chemical and biological processes, industrial measurement, nuclear power, and detection and treatment of disease. (4.4d)