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Science - Chemistry

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Options	Standards	Essential Questions	Content	Skills	Suggested Resources	Assessment	Resources
			I. Atomic Concepts 3.1a The modern model of the atom has evolved over a long period of time through the work of many scientists. 3.1b Each atom has a nucleus, with an overall positive charge, surrounded by negatively charged electrons. 3.1c Subatomic particles contained in the nucleus include protons and neutrons. 3.1d The proton is positively charged, and the neutron has no charge. The electron is negatively charged. 3.1e Protons and electrons have equal but opposite charges. The number of protons is equal to the number of	relate experimental evidence (given in the introduction of Key Idea 3) to models of the atom (3.1ii) use models to describe the structure of an atom (3.1i) determine the number of protons or electrons in an atom or ion when given one of these values (3.1iii) calculate the mass of an atom, the number of neutrons or the number of protons, given the other two values (3.1iv) distinguish between ground state and excited state electron configurations, e.g., 2-8-2 vs. 2-7-3 (3.1v)			

			electrons in an atom. 3.1f 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 neutron In the wave-mechanical model (electron cloud), the electrons are in orbitals, which are defined as regions of most probable electron location (ground state). Each electron in an atom has its own distinct amount of energy. When an electron in an atom gains a specific amount of energy, the electron is at a higher energy state (excited state). 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. The outermost electrons in an atom are called the valence electrons. In general, the number of valence	identify an element by comparing its bright-line spectrum to given spectra (3.1vi) draw a Lewis electron-dot structure of an atom (3.1viii) distinguish between valence and non-valence electrons, given an electron configuration, e.g., 2-8-2 (3.1vii) given an atomic mass, determine the most abundant isotope (3.1xi) calculate the atomic mass of an element, given the masses and ratios of naturally occurring isotopes (3.1xii)			

			electrons affects the chemical properties of an element. Atoms of an element that contain the same number of protons but a different number of neutrons are called isotopes of that element. The average atomic mass of an element is the weighted average of the masses of its naturally occurring isotopes.				
			I I . P e r i o d i c T a b l e The placement or location of an element 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. 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.	explain the placement of an unknown element in the Periodic Table based on its properties (3.1xvi) interpret and write isotopic notation (3.1x) classify elements as metals, nonmetals, metalloids, or noble gases by their properties (3.1xiii) describe the states of the elements at STP (3.1xviii) determine the group of an element, given the chemical formula of a compound, e.g., XCl or XCl2 (3.1xv) compare and contrast properties of			

			Common notations that represent isotopes include: 14C, 14C, carbon-14, C-14. 6 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. Elements can be differentiated by their physical properties. Physical properties of substances, such as density, conductivity, malleability, solubility, and hardness, differ among elements. Elements can be differentiated by chemical properties. Chemical properties describe how an element behaves during a chemical reaction. Some elements exist as two or more forms in the same phase. These forms differ in their molecular or crystal structure, and hence in their	elements within a group or a period for Groups 1, 2, 13- 18 on the Periodic Table (3.1xiv)			

			properties. 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. The succession of elements within the same group demonstrates characteristic trends: differences in atomic radius, ionic radius, electronegativity, first ionization energy, metallic/nonmetallic properties. The succession of elements across the same period demonstrates characteristic trends: differences in atomic radius, ionic radius, electronegativity, first ionization energy, metallic/nonmetallic properties.				
			III. Moles and Stoichiometry A compound is a substance composed of two or more different	determine the molecular formula, given the empirical formula and molecular mass			

			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. Types of chemical formulas include: empirical, molecular, and structural. 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. In all chemical reactions there is a conservation of mass, energy, and charge. A balanced chemical equation represents conservation of atoms.	(3.3vii) determine the empirical formula from a molecular formula (3.3v) interpret balanced chemical equations in terms of conservation of matter and energy (3.3ii) balance equations, given the formulas for reactants and products (3.3i) interpret balanced chemical equations in terms of conservation of matter and energy (3.3ii) create and use models of particles to demonstrate balanced equations (3.3iii) calculate simple mole-mole stoichiometry problems, given a balanced equation (3.3iv) calculate the formula mass and the gram-formula mass (3.3viii) determine the number of moles of a substance, given its mass (3.3ix) determine the mass of a given number of moles of a substance (3.3vi) identify types of			

			The coefficients in a balanced chemical equation can be used to determine mole ratios in the reaction. The formula mass of a substance is the sum of the atomic masses of its atoms. The molar mass (gramformula mass) of a substance equals one mole of that substance. The percent composition by mass of each element in a compound can be calculated mathematically. Types of chemical reactions include synthesis, decomposition, single replacement, and double replacement.	chemical reactions (3.2ii)			
			I V . C h e m i c a l B o n d i n g Compounds can be differentiated by their chemical and physical properties.	distinguish among ionic, molecular, and metallic substances, given their properties (3.1xix) demonstrate bonding concepts using Lewis dot			

			Two major categories of compounds are ionic and molecular (covalent) compounds. Chemical bonds are formed when valence electrons are: transferred from one atom to another (ionic); shared between atoms (covalent); mobile within a metal (metallic). 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. Molecular polarity can be determined by the shape and distribution of the charge.	structures representing valence electrons: transferred (ionic bonding); shared (covalent bonding); in a stable octet (5.2i) determine the noble gas configuration an atom will achieve when bonding (5.2iv) demonstrate bonding concepts, using Lewis dot structures representing valence electrons: transferred (ionic bonding); shared (covalent bonding); in a stable octet (5.2i) distinguish between nonpolar covalent bonds (two of the same nonmetals) and polar covalent bonds (5.2v)			

			Symmetrical (nonpolar) molecules include CO_2, CH_4, and diatomic elements. Asymmetrical (polar) molecules include HCl, NH_3, H_2O. 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. When a bond is broken, energy is absorbed. When a bond is formed, energy is released. Atoms attain a stable valence electron configuration by bonding with other atoms. Noble gases have stable valence electron				

			configurations and tend not to bond. 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. Electron-dot diagrams (Lewis structures) can represent the valence electron arrangement in elements, compounds, and ions. Electronegativity indicates how strongly an atom of an element attracts electrons in a chemical bond. Electronegativity values are assigned according to arbitrary scales. The electronegativity difference between two bonded atoms is used to assess the degree of polarity in the bond. 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.					
		V. Physical Behavior of Matter Matter is classified as a pure substance or as a mixture of substances. The three phases of matter (solids, liquids, and gases) have different properties. A pure substance (element or compound) has a constant composition and constant properties throughout a given sample, and from sample to sample. Elements are substances that are composed of atoms that have the same atomic number. Elements cannot be broken down by chemical change. Mixtures are composed of two or more different substances that can be separated by physical means.	use a simple particle model to differentiate among properties of a solid, a liquid, and a gas (3.1xxii) use particle models/diagrams to differentiate among elements, compounds, and mixtures (3.1xxxvi) describe the process and use of filtration, distillation, and chromatography in the separation of a mixture (3.1xxiv) interpret and construct solubility curves (3.1xxv) use solubility curves to distinguish among saturated, supersaturated and unsaturated solutions (3.1xxviii) apply the adage "like dissolves like" to real-world situations (3.1xxvi) describe the preparation of a solution, given the molarity (3.1xxx) interpret solution concentration data (3.1xxx)				

			When different substances are mixed together, a homogeneous or heterogeneous mixture is formed. The proportions of components in a mixture can be varied. Each component in a mixture retains its original properties. 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. 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. The concentration of a solution may be expressed as: molarity (M), percent by volume, percent by mass, or parts per million (ppm). 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.	calculate solution concentrations in molarity (M), percent mass, and parts per million (ppm) (3.1xxix) distinguish between heat energy and temperature in terms of molecular motion and amount of matter (4.2i) qualitatively interpret heating and cooling curves in terms of changes in kinetic and potential energy, heat of vaporization, heat of fusion, and phase changes (4.2iii) distinguish between heat energy and temperature in terms of molecular motion and amount of matter (4.2i) explain phase changes in terms of the changes in energy and intermolecular distance (4.2ii) explain the gas laws in terms of KMT (3.4i) solve problems, using the combined gas law (3.4ii) convert temperatures in Celsius degrees (oC) to kelvins (K), and kelvins to Celsius degrees (3.4iii)			

			The greater the concentration of solute particles the greater the effect. Energy can exist in different forms, such as chemical, electrical, electromagnetic, thermal, mechanical, and nuclear. Heat is a transfer of energy (usually thermal energy) from a body of higher temperature to a body of lower temperature. Thermal energy is associated with the random motion of atoms and molecules. Temperature is a measure of the average kinetic energy of the particles in a sample of matter. Temperature is not a form of energy. The concept of an ideal gas is a model to explain behavior of gases. A real gas is most like an ideal gas when the real gas is at low pressure and high temperature. Kinetic molecular theory (KMT) for an ideal gas states all gas particles: ? are in random, constant, straight-line motion ? are separated by great distances	qualitatively interpret heating and cooling curves in terms of changes in kinetic and potential energy, heat of vaporization, heat of fusion, and phase changes (4.2iii) calculate the heat involved in a phase or temperature change for a given sample of matter (4.2iv) explain phase change in terms of the changes in energy and intermolecular distances (4.2ii) distinguish between endothermic and exothermic reactions, using energy terms in a reaction equation, ''H, potential energy diagrams or experimental data (4.1i) use a simple particle model to differentiate among properties of solids, liquids, and gases (3.1xxii) explain vapor pressure, evaporation rate, and phase changes in terms of intermolecular forces (5.2iii) compare the physical properties of substances based upon			

			relative to their size; the volume of gas particles is considered negligible ? have no attractive forces between them ? have collisions that may result in a transfer of energy between particles, but the total energy of the system remains constant. Collision theory states that a reaction is most likely to occur if reactant particles collide with the proper energy and orientation. Kinetic molecular theory describes the relationships of pressure, volume, temperature, velocity, and frequency and force of collisions among gas molecules. Equal volumes of gases at the same temperature and pressure contain an equal number of particles. 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.	chemical bonds and intermolecular forces (5.2ii)			

			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. Chemical and physical changes can be exothermic or endothermic. The structure and arrangement of particles and their interactions determine the physical state of a substance at a given temperature and pressure. 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. 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.					
		V I .K i n e t i c s / E q u i l i b r i u m Collision theory states that a reaction is most likely to occur if reactant particles collide with the proper energy and orientation. The rate of a chemical reaction depends on several factors: temperature, concentration, nature of reactants, surface area, and the presence of a catalyst. Some chemical and physical changes can reach equilibrium. 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. LeChatelier's principle can be used to predict the effect of stress (change in pressure, volume,	use collision theory to explain how various factors, such as temperature, surface area, and concentration, influence the rate of reaction (3.4vi) identify examples of physical equilibria as solution equilibrium and phase equilibrium, including the concept that a saturated solution is at equilibrium (3.4 vii) describe the concentration of particles and rates of opposing reactions in an equilibrium system (3.4iv) qualitatively describe the effect of stress on equilibrium, using LeChatelier's principle (3.4v) read and interpret potential energy diagrams: PE of reactants and products, activation energy (with or without a catalyst), heat of reaction (4.1ii) compare the entropy of phases of matter (3.1xxiii)			

			concentration, and temperature) on a system at equilibrium. Energy released or absorbed by a chemical reaction can be represented by a potential energy diagram. Energy released or absorbed by 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. A catalyst provides an alternate reaction pathway which has a lower activation energy than an uncatalyzed reaction. Entropy is a measure of the randomness or disorder of a system. Asystem with greater disorder has greater entropy. Systems in nature tend to undergo changes toward lower energy and higher entropy.				
			VII. Organic Chemistry Organic compounds contain	classify an organic compound based on its structural or condensed structural formula			

			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. 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. Organic acids, alcohols, esters, aldehydes, ketones, ethers, halides, amines, amides, and amino acids are types of organic compounds that differ in their structures. Functional groups impart distinctive physical and chemical properties to organic compounds. Isomers of organic compounds have the same molecular formula, but different structures and properties. In a multiple covalent bond, more than one pair	(3.1xvii) draw structural formulas for alkanes, alkenes, and alkynes containing a maximum of ten carbon atoms (3.1xxi) classify an organic compound based on its structural or condensed structural formula (3.1xvii) draw a structural formula with the functional group (s) on a straight chain hydrocarbon backbone, when given the correct IUPAC name for the compound (3.1xx) identify types of organic reactions (3.2iv) determine a missing reactant or product in a balanced equation (3.2iii)				

			of electrons are shared between two atoms. Unsaturated organic compounds contain at least one double or triple bond. Types of organic reactions include: addition, substitution, polymerization, esterification, fermentation, saponification, and combustion.				
			VIII.Oxidation-Reduction An oxidation-reduction (redox) reaction involves transfer of electrons (e-). Reduction is the gain of electrons. A half-reaction can be written to represent	determine a missing reactant or product in a balanced equation (3.2iii) write and balance half-reactions for oxidation and reduction of free elements and their monatomic ions (3.2vi) compare and contrast voltaic and electrolytic cells (3.2ix) identify and label			

			reduction. Oxidation is the loss of electrons A half-reaction can be written to represent oxidation. In a redox reaction the number of electrons lost is equal to the number of electrons gained. Oxidation numbers (states) can be assigned to atoms and ions. Changes in oxidation numbers indicate that oxidation and reduction have occurred. An electrochemical cell can be either voltaic or electrolytic. In an electrochemical cell, oxidation occurs at the anode and reduction at the cathode. A voltaic cell spontaneously converts chemical energy to electrical energy. An electrolytic cell requires electrical energy to produce chemical change. This process is known as electrolysis.	the parts of a voltaic cell (cathode, anode, salt bridge) and direction of electron flow, given the reaction equation (3.2vii) use an activity series to determine whether a redox reaction is spontaneous (3.2x) identify and label the parts of an electrolytic cell (anode, cathode) and direction of electron flow, given the reaction equation (3.2viii)			
			IX. Acids, Bases, and					

			Salts Behavior of many acids and bases can be explained by the Arrhenius theory. Arrhenius acids and bases are electrolytes. 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. Arrhenius acids yield H+ (aq), hydrogen ion as the only positive ion in aqueous solution. The hydrogen ion may also be written as H3O+(aq),	given properties, identify substances as Arrhenius acids or Arrhenius bases (3.1xxxi) write simple neutralization reactions when given the reactants (3.1xxxiv) calculate the concentration or volume of a solution, using titration data (3.1xxxv) interpret changes in acid-base indicator color (3.1xxxiii) identify solutions as acid, base, or neutral based upon the pH (3.1xxxii)				

			hydronium ion. Arrhenius bases yield OH- (aq), hydroxide ion as the only negative ion in an aqueous solution. In the process of neutralization, an Arrhenius acid and an Arrhenius base react to form salt and water. Titration is a laboratory process in which a volume of solution of known concentration is used to determine the concentration of another solution. There are alternate acid-base theories. One such theory states that an acid is an H+ donor and a base is an H+ acceptor. The acidity and alkalinity of an aqueous solution can be measured by its pH value. The relative level of acidity or alkalinity of a solution can be shown by using indicators. On the pH scale, each decrease of one unit of pH represents a tenfold increase in hydronium ion concentration.				
			X. Nuclear Chemistry	calculate the			

			Stability of isotopes is based on the ratio of the neutrons and protons in its nucleus. Although most nuclei are stable, some are unstable and spontaneously decay emitting radiation. Each radioactive isotope has a specific mode and rate of decay (half-life). 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. 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, ionizing power, and penetrating power. Nuclear reactions include natural and artificial	initial amount, the fraction remaining, or the halflife of a radioactive isotope, given two of the three variables (4.4i) determine decay mode and write nuclear equations showing alpha and beta decay (3.1ix) and contrast fission and fusion reactions (4.4ii) complete nuclear equations; predict missing particles from nuclear equations (4.4iii) identify specific uses of some common radioisotopes, such as: I-131 in diagnosing and treating thyroid disorders; C-14 to C-12 ratio in dating living organisms; U-238 to Pb-206 ratio in dating geological formations; Co-60 in treating cancer (4.4iv)			

			transmutation, fission, and fusion. There are benefits and risks associated with fission and fusion reactions. 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. 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. Energy released during nuclear reactions is much greater than the energy released during chemical reactions. 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. 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 diseases.				

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