1. A The Scientific Revolution occurred in the sixteenth and seventeenth centuries. During this time, many scientific developments were made. [6,2]
2. E Historians are reconsidering the usage of the term "Scientific Revolution" because there is no single event, time, or location that defines it. Also, the changes it brought are less revolutionary than previously thought. [6,3]
3. B Under the Aristotelian view, all matter was composed of some combination of water, fire, earth, or air. This belief persisted until the early 1600s. [6,5]
4. D In the beginning of the seventeenth century, chemistry and alchemy were very similar disciplines. The goal of alchemy was to transform other metals into gold or silver. [6,5]
5. B Empiricism, which became popular in the 1600s, emphasized the use of scientific evidence to develop conclusions. [7,0]
6. C Prior to the rise of scientific societies, scientific discoveries were kept secret. These scientific societies encouraged publishing research, reviewing others' theories, replicating experiments, and exposing mistakes. [7,1]
7. D Atoms, made up of positive, negative, and neutral subatomic particles, are the smallest distinct particles that make up matter. [7,1,2]
8. C An ion is formed when an atom loses or gains one or more electrons. If it loses an electron, it becomes a positive ion (a cation), and if it gains an electron, it becomes a negative ion (an anion). [7,1,2]
9. C By measuring the relative masses of elements in compounds, chemists discovered and categorized elements' atomic masses. [7,2,0]
10. D While its main purpose is to determine relative masses of atoms, mass spectrometry is also used in airports to screen for chemicals that are present in common explosives. [7,2,1]
11. B A mass spectrometer can separate atoms because ions (charged particles) of different mass and energy react differently to magnetic and electric fields. [7,2,1;8,1,0]
12. C A mass analyzer is a component of a mass spectrometer that accelerates ions using magnetic fields and sorts them into different pathways. [8,1,0]
13. C One common type of mass spectrometer uses a curved path. To create this curved path, a curved magnet must be used. [8,1,1]
14. D Samples for mass spectrometers with straight pathways are generally gaseous. [8,1,1]
15. D By either altering the magnetic field and/or the detector position, mass spectrometers can measure substances with multiple different atoms. [8,1,1]

1. D John Dalton was an English scientist who made important contributions to the field of chemistry. [8,1]
2. A While John Dalton's work on atomic theory was one of his greatest contributions to science, the concept of atomism was first proposed by the Greek philosopher Leucippus in the fifth century. [8,1]
3. C John Dalton proposed his atomic theory, which is considered the first modern atomic theory, in the early nineteenth century. [8,1]
4. D While John Dalton was also a chemist and a physicist, he was especially interested in studying meteorology. [8,2]
5. A For 57 years, John Dalton kept daily scientific records. These records led him to develop his theories. [8,2]
6. D John Dalton particularly focused on studying gases. His study of gasses allowed him to propose the law of partial pressures. [9,0]
7. A From his studies, Dalton observed that as long as they were at the same pressure, two liquids subjected to the same temperature change would have the same changes in volume. This discovery led him to propose the law of partial pressures. [9,0]
8. C The law of partial pressures states that the sum of the pressures of each gas in a mixture is equivalent to the total pressure of the mixture. This law is represented by the equation $P_{total} = P_A + P_B$. [9,0]
9. A The French chemist Antoine Lavoisier proposed the law of conservation of mass. By this law, matter is neither created nor destroyed and the mass of the reactants must equal the mass of the products. [9,1]
10. E John Dalton based his law of multiple proportions on the works of Antoine Lavoisier and Joseph Louis Proust. Proust proposed the law of definite proportions. [9,1;9,2]
11. B Ten years separate the publications of the law of conservation of mass and the law of definite proportions, in 1789 and 1799, respectively. [9,1]
12. D Based on his law of multiple proportions, John Dalton proposed his atomic theory, which is considered the first modern atomic theory. [9,2]
13. E John Dalton published *A New System of Chemical Philosophy*, in which he made several major conclusions, including that all matter is composed of atoms, atoms cannot be further divided, atoms of different elements combine in ratios to form compounds, and atoms of different elements have different properties. [9,2]
14. C Though some conclusions in the publication, including that atoms are indivisible, are not true, *A New System of Chemical Philosophy* was notable because unlike works before it, it was rooted in scientific experimentation. [9,3]
15. D By using mass spectroscopy, scientists determined that different isotopes of the same atom have different atomic masses. Isotopes are a type of atom that has the same atomic number as an element, but different numbers of neutrons, thus different masses. [9,3]

1. B Atoms consist of three subatomic particles: the proton, the neutron, and the electron. Electrons have a negative charge, and their symbol is e^-. [10,1,1]
2. E A nucleus contains at least one proton and one or more neutrons. However, the nucleus of hydrogen's most common isotope, is only composed of one proton. [10,1,1]
3. D The symbol "n^0" represents a neutron, as neutrons have no charge and are thus neutral. [10,1,1]
4. C A is the mass number, Z is the atomic number, and X is the elemental symbol, the isotope is represented by $^{A}_{Z}X$. In this example, the mass number is 227, the atomic number is 89, and the elemental symbol is Ac, so the hypothetical element Alpacatinium is represented as $^{227}_{89}Ac$.. [10,1,2]
5. A An element's atomic number is equal to the number of protons present in the nucleus. Relatedly, the mass number is the sum of the number of neutrons and protons. [10,1,2]
6. B A nuclide is defined as an arrangement of protons and/or neutrons that form a nucleus. It is characterized by the mass of its protons and neutrons, the charge of its protons, and the energy content. [10,1,2]
7. B Carbon has 3 naturally occurring isotopes, $^{12}_{6}C$ (the most abundant), $^{13}_{6}C$, and $^{14}_{6}C$. [10,1,2]
8. C An atom that is radioactive is unstable and decays over time to form a more stable nuclei. decays over time to form a more stable nuclei. For example, $^{240}_{94}Pu$ decays to $^{236}_{92}U$. [10,1,2]
9. A At first, hydrogen's atomic mass of 1 was the standard for atomic mass. However, because carbon is solid, it is more convenient to use for comparison than hydrogen, because hydrogen is gaseous. [10,2,1]
10. D Tritium is an isotope of hydrogen that has two neutrons. Its symbol is $^{3}_{1}H$. [11,figure 5]
11. E Deuterium, or $^{2}_{1}H$ (hydrogen with one neutron), is the nuclide that is most commonly used in nuclear fusion reactions. [11,1,0]
12. C $^{14}_{6}C$ forms when cosmic rays hit nitrogen in the earth's upper atmosphere. It then combines with oxygen in the air to form carbon dioxide, mixes with other carbon dioxide, and enters the biosphere. [11,1,1]
13. A $^{60}_{27}Co$ is a radioactive isotope of cobalt. It can be used to kill cancerous cells and examine steel for flaws. [11,2,0]
14. D $^{14}_{6}C$ is a radioactive isotope of carbon that scientists use to trace carbon in chemical reactions, date archaeological artifacts, and study climate change. Scientists even discovered the mechanisms for photosynthesis and cellular respiration using it as a tracer molecule. [11,1,2;11,1,3]
15. E To find a weighted average mass, for each isotope of an element, multiply its abundance (as a decimal) with its mass, then sum each product. .7 x 40 + .2 x 35 + .1 x 38 = 38.8. [11,2,2]

1. B In order to excite electrons, or raise them to a higher energy level, chemists can strongly heat an atom. Alternatively, electrons can become excited when they absorb light. [11,2,3]
2. D A photon is defined as a quantum of light energy. Photons are released when electrons move from an excited state back to their ground state. [11,2,3]
3. B An absorption spectrum is created when atoms absorb specific wavelengths of light. The photons that a substance absorbs are examined to determine their energy. [12,1,0;82,1,2]
4. E In the Bohr model of the atom, electron orbits are stationary and unchanging. The Danish physicist Niels Bohr proposed this model in 1913. [12,1,1]
5. A The nucleus is the dense center part of an atom. Because it contains only neutral neutrons and positive protons, it is always positively charged. [12,1,1]
6. C Because of the distance from each orbit to the center of the atom, electron-nucleus pairs have certain potential energies. The emission of light as a result of the deexcitation of an electron will have an energy value corresponding to the potential energy of the electron-nucleus pair. [12,1,1]
7. D When an electron absorbs light, it gains energy and moves further away from the nucleus. [12,1,1]
8. D The maximum occupancy of one orbit is 2 electrons. Though 2 electrons occupying an orbit are very similar, their properties will vary slightly. [12,1,1]
9. E Scientists proposed the quantum mechanical model on the basis that matter has booth mass and wave properties. This revelation was especially apparent for electrons, which have minimal mass. [12,1,2]
10. A In the Bohr model of the atom, electrons occupy specific locations at fixed distances from the nucleus. Additionally, the Bohr model dictates that electrons rotate in specific orbits. [12,1,2]
11. A Scientists founded the field of quantum physics upon the discovery of matter's wave-particle duality. Scientists first noticed wave-particle duality in the behavior of light. [12,2,1]
12. C Both the quantum mechanical model and the Bohr model dictate that an electron's energy has specific values. However, while the Bohr model proposes that electrons rotate in orbits, the quantum mechanical model does not illustrate the path of electrons. Instead, it hypothesizes that electrons occupy vague clouds. [12,figure 7]
13. E The photoelectric effect occurs when a photon transfers its energy too an electron which is emitted from a metal surface. This effect demonstrated that light acts as a particle. [12,2,2]
14. B In the Davisson-Germer experiment, electrons scattered at specific angles after hitting a crystal. This observation demonstrated the wave properties of electrons. [12,2,2]
15. E In a classic experiment, scientists shot electrons at a nickel crystal to demonstrate the wave properties of electrons. [12,2,2]

1. E A node is the place within an orbital where electrons cannot reside. The more nodes an orbital has, the higher energy it is. [13, figure 9]
2. A Orbitals are electron clouds that are located at specific locations. There are four categories of orbitals: s, p, d, and f. They are differentiated by their shapes. [13,1,1]
3. A In the table of elements, the horizontal rows are called periods and the vertical columns are called rows. [13,2,0]
4. A There are 18 groups. Sometimes they are also labelled into A and B groups with 8 A groups and 10 B groups. [13,2,0]
5. C Atoms in the same group have the same number of valence electrons which in turn, leads them to have very similar chemical properties. [13,2,0]
6. E Nodes are spaces in orbitals where electrons do not exist. A higher number of nodes indicates that the orbital has higher energy. [13, figure 9]
7. C The characteristic of a particular atom's electron orbitals indicates the types of chemical bonding that the atom can perform. [13,2,0]
8. C Main groups consist of 10 of the 18 groups which are labelled with an "A". Elements that belong to the same main group share the same number of valence electrons and have similar chemical and physical properties. [13,2,0]
9. A Orbitals are used as a way to explain the behavior of electrons and their wave-like properties. Orbitals indicate specific areas in which electrons reside around the nucleus. They are given labels depending on their shape. [13,1,1; 13,1,2]
10. C In the periodic table, elements are ordered by their atomic number which is equivalent to the number of protons they have. There are also trends in the properties of each element such as electronegativity and atomic radius. [13,1,2]

1. C Periodic trends occur because of variations in attractive forces and electron orbital structures. [14,1,0]
2. D Atomic radius decreases from left to right across a period and increases down a group. [14,1,1]
3. E Carbon has 6 protons, so its atomic number is 6. [14,1,1]
4. E Ionization energy is the energy required to remove a valance electron from an atom. Effectively, it measures the force holding electrons to the nucleus. [14,1,2]
5. C As protons are added moving across a period, the nucleus's positive charge increases in strength. This increase in charge more strongly attracts the electrons of an atom, pulling them closer, and thus decreasing atomic radius. [14,1,2]
6. C Negative ions are formed when an element attracts one or more additional electrons. Fluorine that has attracted another electron is a negative ion, F^-. [14,2,1]
7. D Nitrogen provides an exception to the periodic trend of electron affinity. This exception occurs due to the electronic orbital structure. [14,2,1]
8. E Each added orbital is considered a shell. Once each shell fills, another is added. [14,2,2]
9. C Atomic radius increases down a group on the periodic table because as the number of electrons in an atom increases, they repulse each other more. This repulsion increases the distance between the outermost electrons and the nucleus. This effect is called shielding and causes increases in atomic radius. [14,2,3]
10. D Increased distance between electrons and the nucleus causes a decrease in attraction between the two. Distance and attraction are therefore described as inversely proportional. [15,1,1]
11. A Linus Pauling derived a measurement of electronegativity values. These values are referred to as Pauling electronegativities. [15,1,3]
12. C The most electronegative element is fluorine, with an electronegativity value of approximately 4. [15,figure 12]
13. A Ionic bonds are classified by a difference in electronegativity values of 1.8 or greater. Some examples of ionic compounds are NaBr and LiCl, as their differences in electronegativity are both about 2. [15,1,5;15,2,0]
14. D Dipole moments occur most often in polar covalent molecules. Because the difference in carbon and oxygen's electronegativity values is 1, it is considered a polar covalent compound. [15,2,0]
15. C If electronegativity values are identical, the change in electronegativity between the two elements will be zero. Because both hydrogen atoms have an electronegativity value of about 2.1, the change in electronegativity between the two atoms is 0. [15,2,0]

1. B A molecule is defined as a group of atoms held in a constant ratio by strong covalent bonds. Examples of molecules include nitrogen, N_2, and caffeine, $C_8H_{10}N_4O_2$. [15,2,1]
2. B Silica, which constitutes most sand, exists in the form of a network. [16,1,0]
3. E Silica, or silicon dioxide, has the chemical formula SiO_2. [16,1,0;15,figure 13]
4. D Table salt is an ionic compound, and it is structured as a large lattice. [16,1,0]
5. B Forces between neighboring molecules are called intermolecular forces. These forces produce a substance's physical properties, like boiling point. [16,1,2;16,2,0]
6. B Water, as opposed to methane, has much stronger forces between neighboring molecules. Because of these stronger forces, it has a higher boiling point (100 degrees Celsius) than that of methane (-162 degrees Celsius). [16,2,0]
7. D All types of attractive forces that pertain to atoms and molecules are created by the electrostatic force. The electrostatic force is the force of attraction between opposite charges. [16,2,1]
8. C Ionic bonds occur when one or more electron is transfer from one atom to another, such as in all salts. However, covalent bonds occur when two atoms share an electron pair, such as in water. [16,2,2]
9. D Metallic bonds occur when atoms bond so that many electrons may move freely between the atoms. This body of moving electrons is generally called an electron "sea." [16,2,2]
10. B Moving electrons cause electricity. Because metals contain moving electrons, they are excellent conductors of electricity. [16,2,2]
11. B A metal made of more than one element is called an alloy, and one example of an alloy is brass. When an alloy forms, it can attain more desirable properties, like increased conductivity. [16,2,2]
12. D Brass is a metal made of copper, Cu, and zinc, Zn and it conducts electricity well. [16,2,2]
13. E A negative ion and positive ion pair form an ionic bond. Negative ions are called anions and positive ions are called cations. [17,figure 15]
14. B Each type of bond has a different range of energies, measured in kilojoules per mole. The bond energy range for covalent bonds is 150 kJ/mol to 1100 kJ/mol. [17,figure 15]
15. A Iron (Fe), sodium (Na), gold (Au), and zinc (Zn) are all metals. Bromine (Br) is a halogen that is liquid at room temperature. [16,2,2;17,figure 15]

1. C In a molecule composed of two identical atoms, there is no net charge because each end of the molecule has the same electronegativity value. Because there is no charge, the bond is nonpolar. [17,1,1]
2. A Lowercase delta, δ, indicates a partial charge. A partial charge occurs when there is a change (uppercase delta indicates a change) in electronegativity between two atoms. [18,1,0]
3. A Intramolecular forces are stronger than intermolecular forces. Because covalent bonds are a type of intramolecular force, whereas the others are all intermolecular forces, it is the strongest. [18,1,1]
4. B Physical changes like melting only impact intermolecular forces. Because ionic bonds are an intramolecular force, physical changes do not affect them. [18,1,1;18,figure 17]
5. C Ion-induced dipole and dipole-induced dipole forces are between a polarizable electron cloud and another charge. However, the only intermolecular force that is caused by two polarizable electron clouds is London dispersion forces. [18,figure 17]
6. E Because HCl is a dipole and Cl_2 is nonpolar, dipole-induced dipole forces exist between the two molecules. [18,figure 17]
7. E The only types of intermolecular forces that exist between two nonpolar molecules are London dispersion forces. For example, 2 adjacent molecules of F_2 will exhibit London dispersion forces. [18,2,0]
8. D Hydrogen bonds form when nitrogen, oxygen, or fluorine is bonded with a hydrogen that is covalently bonded to another strongly electronegative atom in a different molecule. [19,1,0]
9. C Hydrogen bonds are a special type of strong dipole-dipole interaction. They are also the strongest type of dipole-dipole interactions. [19,1,0]
10. D In both liquid water and ice, the hydrogen bond is comprised of a weak bond between a hydrogen atom and the very electronegative oxygen atom on a neighboring water molecule. For ice to melt or water to boil, it must be heated strongly because of the strength of these hydrogen bonds. [19,1,0]
11. C The hydrogen bonds between water molecules are extremely rigid because of their strength. This rigidity causes the structure of ice to be open. [19,1,0]
12. A The bases that make up DNA are adenine, thymine, guanine, and cytosine. Uracil only occurs in RNA, where it replaces thymine. [19,figure 18;19,figure 19]
13. D DNA’s back bone is comprised of phosphate groups and a sugar (deoxyribose), to which base pairs attach. [19,figure 19]
14. D Rosalind Franklin and Maurice Wilkins took the X-rays that showed that DNA has a double helix structure. Based on these X-rays, James Watson and Francis Crick established their model of DNA in the 1950s. [19,2,1]
15. D X-ray studies of DNA suggested that it has a double helix structure because of the presence of hydrogen bonds. This finding has since been confirmed in following studies. [19,2,1]

1. D Hydrogen is the most bountiful element in the universe. It is also the simplest, as it has one proton and one electron. [20,1]
2. B While Paracelsus and Robert Boyle had noticed the formation of a flammable gas in their experiments, Henry Cavendish was the first scientist to describe hydrogen as a distinct gas and to describe its properties. [20,2]
3. E Paracelsus, Robert Boyle, and Henry Cavendish all noticed that reactions between metals and acid released hydrogen. [20,1;20,2]
4. E Though it was once known as phlogiston, hydrogen got its name when Antoine Lavoisier called it *hydrogène*. [20,2]
5. D The positive and negative charges in an ionic bond create electrostatic forces, which extend in all directions and attract many neighboring ions to form a lattice network. [20,1,1]
6. E Diamond, which is a form of carbon, forms a covalent network. [21,1,0]
7. C Because each carbon atom in diamond can bond with up to four other carbon atoms, it is one of the strongest known substances. [21,1,0]
8. B In metallic bonds, layers of atoms can slide over one another rather easily, allowing for metals to be bent, hammered into sheets, or stretched into wires. [21,1,1]
9. A All states of matter exhibit vibrational motion, but translational motion determines state of matter. In solids, no translational motion occurs, some occurs in liquids, and it occurs to the highest degree in gases. [21,1,2;21,2,0]
10. C For a solid to melt or a liquid to vaporize, energy must be added in order to sever intermolecular forces. [21,2,0]
11. B Temperature is a measure of a substance's molecular motion, or the average kinetic energy of constituent particles. [21,2,1]
12. D Because of its intermediate strength hydrogen bonds, water is a liquid at room temperature. [22,1,1]
13. D Some compounds only have weak dispersion forces because they don't exhibit permanent dipoles. These compounds are more likely to be gaseous. [22,1,1]
14. E When a solid changes directly to a gas, it has sublimated. The opposite of sublimation is deposition. [22,figure 23]
15. E Water has intramolecular forces in the form of covalent bonds, and intermolecular forces in the form of hydrogen bonds. Each water molecule has 2 covalent O-H bonds and can form up to four hydrogen bonds. [22,1,2]

1. D Though some scientific models are inaccurate or oversimplified, they are still useful in predicting the outcomes of reactions and the properties of new materials. [22,2,0]
2. C Gilbert Lewis was a physical chemist at the University of California, Berkeley in the early and mid-1900s. [22,2,1]
3. B In Lewis diagrams, electrons are represented as dots. Though the modern view of electrons does not visualize them as dots, the Lewis model makes it easier to visualize atoms, ions, and compounds. [22,2,1]
4. C Valence electrons are electrons that may be involved in bonding. They exist in the outermost electron shell. [22,2,1]
5. C Four chlorine atoms can bond to one carbon atom, as the formula for chlorine tetrachloride is CCl_4. The prefix "tetra" means 4. [22,figure 25]
6. E In Lewis structures, a single dash represents a bond. Double bonds are represented by two dashes, and triple bonds are represented by three. [22,figure 25;23,figure 28]
7. A In general, there are less electrons involved in bonding than there are in the whole atom. One exception to this rule, however, is hydrogen, because it only has one electron. [22,2,1]
8. D Electron orbitals are representations of electron density, so when they overlap, the area is dense with electrons and attractive forces form. [23,1,5]
9. A When two electron orbitals overlap, the resultant attractive force forms a single covalent bond. H_2 is one molecule with a single covalent bond. [23,2,0]
10. A Electrons that are not involved in bonding may also be referred to as nonbonding electrons or as lone pairs. [23,1,2]
11. A When an ionic bond between sodium and chlorine forms, one electron is transferred, satisfying the octet rule for both atoms. [23,1,3;23,figure 26]
12. A Each chlorine atom has 7 outer shell electrons, or very close to a full shell. Because it is so close to having a full shell, it attracts electrons easily, making it very reactive. [23,figure 26]
13. B Originally, the Lewis model did not account for the existence of orbitals, so scientists had to adapt the model to explain compound formation and formulas. [23,1,4]
14. E CO_2 has two double bonds, and each double bond is 2 pairs of electrons, so there are 8 electrons that compose the bonds in CO_2. [23,2,1;23,figure 28]
15. C Each molecule of nitrogen, N_2, has a triple bond and 2 pairs of nonbonding electrons. [23,2,1;23,figure 28]

1. B Chemists formulated the concept of hybridization when the interactions of some atomic orbitals did not predict the correct molecular shape. [23,2,2]
2. B In chemistry, hybridization refers to two or more electron orbitals combining to form new orbitals with a different shape. Hybridization explains the properties of some molecules that were not previously understood. [23,2,2]
3. E When 1 s and 3 p orbitals combine, they form 4 sp^3 orbitals. An s orbital is spherically shaped, and a p orbital looks like a dumbbell. [23,2,2]
4. C Chemists can use mathematical models to trace the positions of bonding electrons, which are also known as "electron waves." [23,2,3]
5. D When 1 s orbital and 1 p orbital combine, they form two sp orbitals. When 1 s and 2 p orbitals combine, they form three sp^2 orbitals. [24,figure 29]
6. D In chemistry, MO stands for "molecular orbitals." Similarly, AO stands for "atomic orbitals." [24,1,0]
7. B In sigma bonds, the electrons are concentrated along an imaginary axis connecting the atoms. However, in pi bonds, electrons are concentrated between the atoms but away from the center line. [24,1,0]
8. E When 2 s orbitals combine, a sigma bond forms. When 2 p orbitals combine, a pi bond forms. [24,figure 30]
9. C For a pi bond to form, the atomic orbitals must align in the x-coordinate direction, as notated by the subscript "x" in each orbital. [24,2,2]
10. C Each F atom has 7 valence electrons, so a molecule of F_2 will have 14 valence electrons. [25,1,1]
11. E Because O_2 must share two pairs of electrons, it forms one sigma bond and one pi bond. [25,1,1]
12. B N_2 shares 3 pairs, or 6 total electrons. Because it shares three pairs, it has a triple bond. [25,1,1]
13. C Between the carbon atom in CO_2 and one of its oxygen atoms, there are two pairs of shared electrons. Since there are two shared pairs, one sigma bond and one pi bond form. [25,2,1]
14. B One molecule of CO_2 has 16 total valence electrons; 6 from each O atom and 4 from the C atom. [25,2,1]
15. C More advanced molecular orbital theory discusses not only bonding and non-bonding orbitals, but also anti-bonding orbitals. An anti-bonding orbital weakens the chemical bond between two atoms. [25,2,2]

1. D Oxidation states are values that represent an atom's loss or gain of one or more electrons. They are essentially the hypothetical charge an atom would have if all its bonds were completely ionic. [25,2,3]
2. A A lithium ion, Li^+, forms when a lithium atom, previously having 3 electrons, loses 1 electron. Because it has lost one electron, its oxidation number is +1. [25,2,4]
3. A Each hydrogen atom in water has an oxidation number of +1, so the oxygen atom must have an oxidation number of -2. [25,2,4]
4. D In neutral molecules, the total oxidation state for all the constituent atoms must add to zero. In charged species, the total sum of the oxidation numbers is equal to the species' charge. [25,2,4]
5. D The nitrogen atom in ammonium (NH_4^+) has an oxidation number of -3. Whereas some elements always have the same oxidation number, nitrogen's oxidation numbers vary widely. [25,2,4]
6. D Each of the four hydrogen atoms in ammonium (NH_4^+) has an oxidation number of +1. Most of the time, hydrogen has the oxidation number +1, but when combined with a metal, it has the oxidation number -1. [25,2,4]
7. D The total of the oxidation numbers in PO_4^{3-} must add to -3. Because each oxygen atom has an oxidation number of -2, the phosphorous atom must have the oxidation number +5 as $5 + 4(-2) = -3$. [25,2,4]
8. E The total of oxidation numbers in H_2S must add to zero. If each hydrogen's oxidation number is +1, then the sulfur atom must have an oxidation number of -2 because $-2 + 2(1) = 0$. [25,2,4]
9. C The VSEPR model predicts molecular geometry as well as other geometrical properties of a molecule like bond length. [26,1,1]
10. B The Valence Shell Electron Pair Repulsion (VSEPR) model depends on the repulsion of electron pairs. Because electron pairs are all negatively charged, they repel each other. [26,2,0]
11. D While the Lewis model is useful in representing the two-dimensional structure of molecules, the VSEPR model can represent molecules in three dimensions, which is more accurate. [26,1,1:26,figure 33]
12. E CH_4, or methane, has tetrahedral geometry. [26,figure 33]
13. D Molecules like BeF_2, beryllium fluoride, and CO_2, carbon dioxide, have linear geometry. [26,figure 33]
14. E In BF_3, each of the 3 fluorine atoms has 6 valence electrons, so the molecule has 18 valence electrons total. [26,figure 33]
15. E SF_6 consists of 1 sulfur atom bonded to 6 fluorine atoms. All six bonds are single bonds, with 2 electrons each, so the total number of bonded electrons in the molecule is 12. [26,figure 33]

1. D The resonance model was introduced because experiments found that bonds had equal properties despite insufficient valence electrons and unsatisfied octet requirements. [26,2,1]
2. B One molecule of ozone, O_3, contains 3 oxygen atoms. One oxygen molecule contains two oxygen atoms, however, as it is a diatom. [26,2,1]
3. B Ozone has 18 valence electrons, but experiments have found that only one type of bond exists between the atoms. [26,2,1]
4. D Ozone has two possible resonance structures, one with a double bond on the left and one with the double bond on the right. [26,2,1]
5. E The term "resonance" arose out of the idea that a molecule oscillated between two or more possible structures. [27,1,1]
6. C In SO_3, each of the three bonds is identical and each bond total is 1 1/3. [27,1,2]
7. D The resonance model was replaced by a more modern theory because it does not explain sigma and pi bond formation or "delocalized" orbitals. [27,1,3]
8. B Whereas ozone molecules were considered to have one double bond and one single bond, it is more accurate to say that on each side, there are 1 ½ bonds. [27,1,1]
9. A Each of the three oxygen atoms in one molecule of SO_3 has 4 nonbonding electrons, so 12 total nonbonding electrons. [27,figure 37]
10. B A more modern theory, molecular orbital theory (MO theory or MOT), has replaced the resonance model. [27,1,3]
11. C A molecule containing cancelling (symmetrical) bond dipoles does not have a dipole moment. Conversely, a molecule whose bond dipoles do not cancel will have a significant dipole moment. [27,1,5]
12. D The four bonds in methane point to the corners of a tetrahedron, making it a tetrahedral molecule. [27,1,4;27,figure 38]
13. D The bond angles in methane are 109.5°. A methane molecule has 4 hydrogen atoms bonded to a central carbon atom. [27,figure 38]
14. C CO_2 has two polar bonds, but no dipole moment, so the polar bonds must be exactly opposite from each other, making the molecule linear. [27,2,1]
15. B Nonpolar molecules have relatively weak intermolecular forces. Because of these weak forces, the molecules are easier to separate, giving them lower melting and boiling points. [27,2,2]

SCIENCE

FOCUSED QUIZ 14
NUCLEAR CHEMISTRY –
FISSION AND FUSION REACTIONS, PP. 28-29

1. E An atom's mass number, A, is equal to the sum of the number of neutrons, N, and the atomic number, Z. [28,1,1]
2. D During alpha decay, a helium nucleus is lost. However, in beta decay, either a neutron transforms into a proton and an electron is emitted, or a proton transforms into a neutron. [28,1,2]
3. A An electron may also be known as a beta minus particle. When a neutron transforms into a proton, an electron is emitted. [28,1,2]
4. E When the ratio of neutrons to protons varies, a nucleus may become unstable. Unstable nuclei become radioactive. [28,1,2]
5. D Alpha decay is most common in nuclei larger than bismuth, with a Z value of 83. [28,1,4]
6. C Neutrinos are subatomic particles that are like electrons, but with no charge and very small (negligible) mass. [28,figure 40]
7. B Plutonium-240 undergoes alpha decay to form uranium-236 and an alpha particle (or a helium-4 nucleus) [28,figure 41]
8. D In beta decay, when there are too few neutrons, positrons are emitted. Conversely, if there are too many neutrons electrons are emitted. [29,1,0]
9. A Sodium-22 undergoes beta decay to form ^{22}Ne and emits an electron, a positron, and a neutrino. [28,figure 40]
10. C The symbol v, the Greek letter nu, represents a neutrino. [28,figure 40]
11. E In beta-plus decay, carbon-10 breaks down to form boron-10, a neutrino, and a positron. [28,figure 42]
12. C In 1919, physicist Ernest Rutherford discovered that bombarding nitrogen with alpha particles creates oxygen atoms. [29,1,1]
13. C In uranium fission, uranium is hit by a free neutron. [29,1,2]
14. E When the products of nuclear fission form, they fly apart with a lot of kinetic energy. This energy can be used to heat water and make steam, among other things. [29,2,0]
15. C Though nuclear fusion can produce electric power, it cannot be done on a large enough scale because a lot of energy is required to fuse the isotopes, as they are both positively charged. [29,2,1]

1. A Elements are distinguished by their number of protons. The number of electrons and neutrons in an element can both change, forming ions and isotopes, respectively. [29,2,3]
2. E Atomic mass is defined relative to carbon-12's mass, which is 12 amus. [30,1,0]
3. C John Dalton proposed his law of partial pressures around 1800. [30,1,2]
4. A The law of partial pressures states that the total pressure of a mixture of gasses is equal to the sum of the pressures of each gas. This law is represented by the equation $P_{Total} = P_A + P_B$. [30,1,2]
5. C Elements are arranged in the period table according to their atomic numbers, which is represented as Z. [30,1,7]
6. C Types of intermolecular forces include ion-dipole forces, London dispersion forces, dipole-induced dipole forces, and ion-induced dipole forces, to name a few. Ionic, covalent, and metallic bonds are type of intramolecular forces. [30,2,1]
7. A When there are large differences in the electronegativity values of two atoms, the bond will be polar. If there is a small change in electronegativity values, the bond is nonpolar. [30,2,2]
8. D Intermolecular forces like ion-dipole forces dictate the physical properties of substances. Physical properties include melting and boiling point, among others. [30,2,3]
9. D Hydrogen bonds are a specific type of dipole-dipole forces. H bonds are stronger than dipole-dipole forces. [30,2,4]
10. B Henry Cavendish was the first to identify hydrogen as a distinct gas and to describe its properties. Its name came from Antoine Lavoisier. [30,2,5]
11. C According to Gilbert Lewis, a covalent bond is made of 2 electrons. Lewis structures are based on this idea. [30,2,6]
12. E Molecular orbitals are orbitals that have hybridized. They include sigma and pi bonds. [30,2,7]
13. B The Valence Shell Electron Pair Repulsion Model (VSEPR) predicts the shapes of molecules depending upon electron pair repulsion. [31,1,2]
14. C The shape of a molecule allows chemists to predict its polarity. Polarity can then predict the physical properties of a molecule. [31,2,1]
15. D Unstable nuclei undergo radioactive decay to become more stable. [31,2,2]

1. B A combination of substances in which each substance remains distinct is a mixture. Mud is a mixture; given time, mud will separate back into dirt and water. [32,1,1]
2. A A combination of substances that looks like one substance is a solution. Drink powder mixed into water is a solution. [32,1,1]
3. C Boyle's law is expressed as PV = C. It indicates that volume of a gas at constant temperature varies inversely with the pressure exerted on it. In other words, the volume will decrease as the pressure increase. [32,2,1]
4. E Early scientists learned to isolate gases and observed that they were highly regular in behavior. [32,2,1]
5. D Boyle's law is expressed as PV = C. It indicates that volume of a gas at constant temperature varies inversely with the pressure exerted on it. In other words, the volume will decrease as the pressure increase. [32,2,1]
6. B Charles's law describes the expanses of gases when they are heated. Volume and temperature can be compared as $V_1/T_1 = V_2/T_2$ [32,2,2]
7. D 0 K is -273 degrees Celsius, the lowest temperature that any system can reach. [33,1,0]
8. C Boyle's law tells us that pressure increases as volume decreases. Therefore, if the volume of a gas in a cylinder with a movable piston has decreased, it is most likely that the pressure has increased. [33,2,3]
9. D The volume and amount of gas is constant, so P/T must stay constant. If the temperature doubles, then the pressure must also double. [33,2,5]
10. D Jacques Charles's investigations into gas stemmed from his fascination with the new technology of hot air balloons. [33,7]
11. D Jacques Charles used hydrogen to fuel his invention of a hot hair balloon. [33,8]
12. C Charles's law relates volume to temperature and in simple terms expresses that fact that gases expand when they are heated. [43,2]
13. A Jacques Charles's experiments established the relationship of volume to temperature that became known as Charles's Law, but he never published his findings. Joseph-Louis Gay-Lussac credited him in his work. [34,2]
14. D (600.0 mL) / (273.0) = (x) / (333.0 K)x = 732 mL [34,2]
15. A Jacques Charles's first hot air balloon was destroyed by suspicious peasants with pitchforks. [34,0]

SCIENCE

FOCUSED QUIZ 17

PARTIAL PRESSURES AND THE CORRECTION OF GAS VOLUMES COLLECTED OVER WATER – THE RELATION OF PARTICLE SPEED TO MASS, COLLISION FREQUENCY, VOLUME, AND PRESSURE, PP. 35-36

1. C Dalton's Law of Partial Pressure states that the sum of the pressures of each gas in a mixture is equal to the total pressure of the mixture. [35,1,1]
2. E Particles that behave in an ideal manner do not stick together or react with one another. [35,1,2]
3. C The mole fraction is the ratio of moles of one gas to the total moles in a mixture of gases. It can be represented by the expression $\frac{n_A}{n_A+n_B+n_C+\cdots}$. [35,1,2]
4. A The ideal gas law is represented by the equation PV = nRT. This law shows that the number of moles, n, is directly proportional to pressure, P. [35,1,2]
5. E In a jar containing water and water vapor, the liquid and gas are at equilibrium, and the water vapor pressure only depends on temperature of the liquid water. [35,1,3]
6. E MmHg, or millimeters of mercury, is a unit for pressure. 1 mmHg is equal to 1 torr, another unit for pressure. [35,1,4]
7. A In chemistry, T represents temperature, and chemists use the Kelvin scale to measure temperature. To convert from Celsius to Kelvin, add 273 degrees. [35,2,0]
8. A KMT, or Kinetic Molecular Theory, depends upon the assumptions that the volumes of gas molecules are negligible, gas molecules move quickly but not all at the same speed, there are no attractive or repulsive forces between molecules, and pressure is a result of collisions between the gas molecules and the container. [35,2,1]
9. D In the field of kinetics, the average molecular speed is often represented by the variable "u". [36,1,1]
10. B Gas molecules exert forces on their container's walls because they change momentum after each collision. [36,1,2]
11. E The force of a molecule's collisions is represented by Force = $\frac{mu^2}{\ell}$, where m represent mass and represents length. [36,1,2]
12. A Because there are three dimensions, one third of the total molecules, N, travel in each direction. [36,1,3]
13. A Where A represents area and represents length, $A\ell$ represents volume. [36,1,3]
14. A One calculation of pressure uses force and unit area, P = Force/A. [36,1,3]
15. E Both Boyle's law and Charles's law can be justified by the equation PV = 1/3 Nmu^2. [36,13]

SCIENCE

FOCUSED QUIZ 18

AVOGADRO'S LAW - VOLUME AND MASS OF ONE MOLE OF IDEAL GAS AT STANDARD TEMPERATURE AND PRESSURE (STP), PP. 36-38

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1. E When the amount of a gas is varied, the pressure and volume of the gas change in predictable ways. The changes in pressure and volume are most affected by the ratio of total mass of gas to mass of an individual gas molecule. [36,2,1]
2. A 4 grams of gaseous helium has the same volume as 2 grams of gaseous hydrogen, which has the same volume as 16 grams of methane (CH_4). [36,2,2]
3. B It is not the mass of the molecules that determines a substance's properties but the number of molecules present. This concept was further explained by Amedeo Avogadro. [36,2,3]
4. C A mole is the base unit that is used to measure the amount of substance. One mole contains a fixed number of atoms. [36,1]
5. D Jean Baptiste Perrin was the scientist that first estimated the number that would come to be called Avogadro's number in honor of Amedeo Avogadro's contributions. [36,1]
6. E In 1805, Joseph-Louis Gay-Lussac discovered that when gases react, the volumes of the products and reactants are always in integer ratios to each other. [37,1]
7. B Amedeo Avogadro was the first to understand that some gases exist only as diatoms, or two identical molecules bound together. Some examples of diatoms are H_2, N_2, and O_2. [37,2]
8. D In chemistry, "n" refers to the number of moles of a substance. The volume of a gas is represented by the letter "V". [37,2]
9. E Avogadro's law states that at constant temperature and pressure, the volume of a gas is directly proportional to the number of moles of the gas. This law is represented by the equation $\frac{V_1}{n_1} = \frac{V_2}{n_2}$. [37,2]
10. A Standard temperature and pressure, STP, is defined as 273 K and 1 atm. [37,1,1]
11. E While it was historically defined as the number of molecules in a 22.4 L volume at STP, the mole is now defined in terms of the carbon-12 isotope. [37,1,1]
12. C Avogadro's number, 6.022×10^{23}, is the number of atoms in one mole. [36,1;37,1,2]
13. B Helium, He, has a molar mass of 4 g/mol because one mole of helium has a mass of 4 grams. [37,2,0]
14. C C_3H_6O, or propionaldehyde, has a molar mass of 58 g/mol because 12 (atomic mass C) x 3 + 1 (atomic mass H) x 6 + 16 (atomic mass O) x 1 = 58. [37,2,2]
15. D The molar mass of table salt is 58.5 g/mol because Na's atomic mass is 23.0 amu and Cl's atomic mass is 35.5 amu. [38,1,0]

SCIENCE

FOCUSED QUIZ 19

THE DEPENDENCE OF KINETIC ENERGY OF MOLECULES ON TEMPERATURE - THE BEHAVIOR OF GASES UNDER EXTREME CONDITIONS, PP. 38-40

1. B “N” represents Avogadro’s number. There are N, or 6.022 x 10^{23}, molecules in one mole. [38,1,2]
2. A The speed of molecules in one mole of gas can be determined using the expression $\sqrt{\frac{3RT}{M}}$. This expression is derived from the expression for an average molecule’s speed, $\sqrt{\frac{3kT}{m}}$. [38,1,3]
3. B The Boltzmann constant, k, is a universal constant that can be used to determine the energy of a molecule. It is not included in the ideal gas law. [38,2,2]
4. A Boyle’s law, $P_1V_1 = P_2V_2$, and Charles’s law, $\frac{V_1}{T_1} = \frac{V_2}{T_2}$, can be combined to form the equation $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$. [38,2,3]
5. D Though Carl Scheele performed his experiments in the 1770s, he did not publish his results until 1777. [38,1]
6. B After his experiments, Carl Scheele noted that oxygen is odorless and tasteless, and that it supports respiration and combustion better than air. Scheele noted these properties in a letter to Antoine Lavoisier. [38,1]
7. A Both Carl Scheele and Joseph Priestley discovered oxygen through their experiments with burning oxides. Specifically, Priestley heated mercuric oxide. [38,1;39,1]
8. B Because he believed that oxygen was air with its phlogiston removed, Joseph Priestley called oxygen dephlogisticated air. [39,1]
9. C Joseph Priestley is more often credited with the discovery of oxygen than Carl Scheele. Priestley isolated 7 other gases not including oxygen. [39,1]
10. D Antoine Lavoisier named oxygen, debunked phlogiston theory, and understood combustion reactions. He claimed to have independently discovered oxygen, but both Carl Scheele and Joseph Priestley had discussed their findings about oxygen with him in 1774. [39,2;39,3]
11. C Antoine Lavoisier was a prolific French chemist who established the law of conservation of mass thanks to his quantitative experiments. [39,3]
12. E The van der Waals equation corrects for the behavior of non-ideal gasses. The term nb corrects for the volume of molecules and the term corrects for molecular attraction. [40,2,0]
13. E As the values of a and b (two parameters included in the van der Waals equation) decrease, a gas behaves more ideally. Helium has the lowest a and b values of 0.0341 and 0.0237, respectfully. [40,2,1]
14. D As the ratio PV/RT approaches 1, the gas becomes more ideal. Ideal behavior can also be measured with the formula P = nRT/V. [40,1,1]
15. C Chlorofluorocarbons were once used in refrigerators because they easily liquify and cool themselves and their surroundings. However, CFCs create holes in the ozone layer and are therefore no longer used. [40,2,2]

SCIENCE

FOCUSED QUIZ 20

THE DEPENDENCE OF RELATIVE RATES OF DIFFUSION AND EFFUSION ON MOLECULAR MASS AND TEMPERATURE – LIQUIDS, PP. 41-42

1. D Diffusion is the movement of a gas from an area of high concentration to an area of low concentration. Effusion, on the other hand, is the movement of a gas out of a container through a small hole. [41,1,2]
2. D The smell of perfume and the odor of a skunk are both examples of diffusion. [41,1,2]
3. E The simplest way to calculate a gas's rate of diffusion or effusion assumes that molecules travel in an uninterrupted straight line. However, this is not the case as real gas molecules experience collisions, so elaborate computer models are necessary in determining exact diffusion and effusion rates. [41,1,3]
4. D For two gases at the same temperature, $\frac{u_1}{u_2} = \sqrt{\frac{M_2}{M_1}}$. This equation represents the idea that the relative speeds of gas molecules, u, are inversely proportional to the square root of the mass, M, of the gas molecules.. [41,1,4]
5. B Helium atoms have a molar mass of 4 g/mol, whereas hydrogen molecules have a molar mass of 2 g/mol. Therefore, helium atoms are 2 times as massive as hydrogen molecules. [41,1,5]
6. C The speed of diffusion and effusion depends on the mass of gas molecules or atoms. Lighter gases diffuse and effuse faster, so Ne, with the lowest molar mass listed, will have the highest diffusion and effusion rate. [41,2,2]
7. E Acetone, a solvent often used for thinning paint, has the chemical formula C_3H_6O. [41,2,3]
8. E Ammonia has 3 hydrogen atoms, each with a mass of 1, and one nitrogen atom with a mass of 14. 3(1) + 1(14) = 17 g/mol. [41,2,4]
9. E All gases on the periodic table except for the noble gases are diatomic, meaning they are gas molecules consisting of 2 identical atoms. The noble gases exist as atoms, not molecules. [41,2,5]
10. B Solids experience vibrational and rotational motion, whereas liquids experience vibrational, rotational, and translational motion. However, liquids have less translational motion than gases. [41,2,5]
11. D During sublimation, a solid transforms directly into a gas. The opposite of sublimation is deposition, where a gas transforms directly to a solid. [41,2,5]
12. A Liquids have long-range order, meaning that the liquid has a form of ordering over long distances but not short ones. [42,1,0]
13. C A liquid's degree of ordering depends on the liquid's nature and the temperature of the liquid. [42,1,0]
14. E Octane, C_8H_{18}, has a low degree of ordering. It is a component of gasoline. [42,1,0]
15. C The three states of matter, listed in order of increasing density, are gases, liquids, and solids. Density = mass/volume. [42,2,1]

1. C In solids, there are four types of structures: ionic lattice solids, covalent network solids, molecular solids, and metallic solids. [42,2,0]
2. B An allotrope is one of many possible forms of an element. These forms have different properties, like differences in conductivity. [42,2,1]
3. B Carbon is an element that has many allotropes; it is hypothesized that up to 500 allotropes of carbon exist! [42,2,1]
4. C The form of carbon that is the most stable is graphite. Graphite exists in sheets of carbon and is used in pencil lead. [42,2,3;43,1,0]
5. B C_{60} is the chemical formula for buckminsterfullerene, which is a specific type of fullerene. [42,figure 49]
6. A While graphite is a good conductor, diamond, another form of carbon, is an insulator, meaning it does not conduct electricity. [42,2,2]
7. E The arrangement of a solid is dependent upon the size of the particle and the particles' bonding abilities. [43,1,1]
8. A Pure metals have some of the most closely packed structures and are composed of many atoms of one single element. [43,1,2]
9. C The three simplest structures of pure metals are body-centered cubic, face-centered cubic, and close-packed hexagonal. Body-centered cubic has 8 nearest neighbors whereas both face-centered cubic and close-packed hexagonal have 12 nearest neighbors. [43,1,3]
10. A Some metals that have a body-centered cubic structure are iron, tungsten, and chromium. [43,figure 50]
11. E The softness of a metal is determined by how closely packed it is, whereas the structure of the metal is determined by the size of the atoms. [43,1,3]
12. B Metals are conductive because their outer electrons can freely move throughout the solid. In other words, metals have a "sea of electrons". [43,2,0]
13. E Because of their tight structure, metals are lustrous, sonorous, ductile, and malleable. They are also good conductors of heat and electricity. [43,1,4;43,2,0]
14. D Ductility is the ability to be pulled into wires. Relatedly, malleability is the ability to be flattened into thin sheets. [43,2,0]
15. A Alloys increase metal strength because the impurities disrupt smooth layers and decrease layer sliding. Steel, an alloy of iron, is much stronger than iron. [43,2,1]

1. C On a phase diagram, pressure is on the vertical axis and temperature is on the horizontal axis. [44,1,1]
2. E The triple point is the only temperature and pressure at which all three states of matter exist in equilibrium. The critical point is the temperature and pressure at which the gas and liquid phases are no longer distinct. [44,1,1]
3. B Horizontal lines on a phase diagram represent phase changes due to changes in temperature rather than pressure. [44,2,0]
4. D Water's critical point occurs at 374°C and 218 atm. [44,2,2]
5. D The phase diagram for water is unusual because the solid-liquid equilibrium line slopes backward. On most phase diagrams, this line slopes forwards. [44,2,3]
6. E Because of the strong hydrogen bonds in water, ice has a rigid and open structure. [45,1,1]
7. A In most substances, an increase in pressure increases the melting point of the solid phase. However, in water, an increase in pressure decreases the melting point of the solid phase. [45,1,1]
8. C The lowest pressure at which CO_2 will become liquid is 5.11 atm. Its triple point occurs at 5.11 atm and -56.4°C. [45,2,1]
9. E A substance that has reached the supercritical point will behave with the qualities of a liquid and a gas. The substance must surpass the critical point to reach its supercritical state. [45,2,2]
10. B Solvents are most beneficial when they have the density of a liquid and the mobility of a gas. In other words, they are most effective when they behave as both liquid and gas. [45,2,2]
11. D While supercritical CO_2 is an effective solvent, it is impractical to use because it must be maintained at a specific temperature and pressure. [45,2,2]
12. B Supercritical CO_2 is advantageous to use as a solvent because once it is used, the pressure can be reduced below the critical value and it will boil away. [46,1,0]
13. D When he discovered carbon dioxide, Scottish chemist Joseph Black called it "fixed air". [46,1]
14. E Joseph Black was interested in experimenting with alkaline substances because they were thought to treat kidney stones. After he heated calcium carbonate, or limestone, he noticed that carbon dioxide was produced. [46,1]
15. A Joseph Black conducted his experiments with carbon dioxide in 1754 and presented his findings in 1755. [46,1]

1. B A solution is a mixture of substances that is homogenous, meaning that you cannot detect a difference anywhere in the substance. Each sip of your Kool-Aid will taste the same as the last. [46,3]
2. A In a solution, the solvent is the substance that is present in a larger quantity than the solute. [46,4]
3. C "Aqueous" means "relating to water". The solvent in an aqueous solution is water. [46,4]
4. D A solution in which the solvent is a nonpolar substance is an organic solution. [46,4]
5. B The relative amounts of solute and solvent present in a solution is the concentration. A solvent that can dissolve no more solute is saturated. [46,7]
6. C A general rule of solubility is that like dissolves like. In other words, polar solvents will dissolve polar solutes. [47,1,1]
7. E Salt is an ionic compound and will dissolve well in water, which is polar. [47,1,2]
8. E Most ionic salts are soluble, but some have lattice forces that are so strong that the forces of hydration cannot overcome them. [47,1,4]
9. A Most sulfates are soluble, but those of barium, strontium, calcium, lead, silver, and mercury are not. [47,2,d]
10. C All the common compounds of Group I are soluble. [47,1,a]
11. D A solution containing exactly the maximum amount of a solute is saturated. Any more solute will not dissolve. [47,2,2]
12. C If more than one salt ions are added to an aqueous solution and they form a solid after the saturation point is reached, that solid is called a precipitate. [47,2,3]
13. D Because a heated solution will dissolve more solute than a non-heated one, it is likely to form solid crystals when cooled. [47,2,2]
14. D Calcium carbonate is an insoluble compound that forms when carbon dioxide dissolves in water containing calcium. When it precipitates on the roof or floor of a cave, it forms stalactites and stalagmites. [48,1,0]
15. A O-H bonds are hydrogen bonds like those found in water. Because like dissolves like, substances containing O-H bonds will dissolve easily in water. [48,1,2]

1. A Intermolecular forces in nonpolar solvents tend to be weak, meaning that they dissolve easily in other nonpolar substances. [48,1,3]
2. A Bonds between carbon and hydrogen have little polarity because their electronegativity values are similar. Polarity depends on the difference between the charges of bonded atoms. [48,1,3]
3. E Nonpolar solvents are particularly useful in cleaning, because they dissolve also-nonpolar grease and oils. [48,1,4]
4. E Nonpolar solvents are particularly useful in cleaning, but they tend to be quite toxic and are difficult to remove from the environment. [48,2,0]
5. C Nonpolar substances can be removed from the environment by burning, which produces carbon dioxide and water. [48,2,0]
6. B A substance with one relatively polar end attached to one relatively nonpolar end is a soap, which chemists developed to convert nonpolar oils and greases to compounds that can be dissolved by water. [48,2,1]
7. E A solution's composition is most often expressed as a percentage. A 3% hydrogen peroxide solution will contain 3 grams of hydrogen peroxide per 100 grams of solution. [48,2,3]
8. B The concentration of some solutions is expressed by weight. For example, a toothpaste tube may be labeled "0.5% sodium fluoride %w/w"; the "w" is the weight. [49,1,0]
9. D Molarity is expressed as moles of a solute in 1 liter of solution. The unit of molarity is expresses as M. [49,1,1]
10. C Molarity is expressed as moles of a solute in 1 liter of solution, so a solution containing 18.02 grams of water (molar mass = 18.02 g/mol) in 1 liter of a salt solution has a concentration of 0.100 M. [49,1,1]
11. C Colligative properties are physical properties that depend on the relative numbers of moles of a solute and solvent. For example, the boiling point of a substance will change depending on how much solvent is present relative to solute. [49,1,2]
12. A The relative numbers of moles of solute and solvent can be expressed as a mole fraction. Mole fractions indicate a substances colligative properties. [49,1,2]
13. C The number of moles of solute dissolved in 1 kilogram (exactly) of solvent is molality. Molality communicates the way that properties change with concentration. [49,1,2]
14. D Molality is linear, so if a 0.100 molal solution of sugar in water freezes at −0.186 °C, a solution that is 0.300 molal will freeze at 3 x (−0.186), or -0.558 °C. [49,1,2]
15. B Molality expresses the change in a substance's properties as its concentration changes. [49,1,2]

METHODS OF EXPRESSING CONCENTRATION-RAOULT'S LAW AND COLLIGATIVE PROPERTIES, PP. 48-50

1. B Boiling point is a colligative property because it changes based on the concentration of a solution. [49,1,3]
2. D Osmotic pressure is a colligative property because it changes based on the concentration of a solution. [49,1,3]
3. A François Raoult formulated rules to explain colligative properties. He established that the relative number of moles of solute in a solution determined certain physical properties. [49,1,3]
4. A Colligative properties, such as boiling and freezing point, depend upon the amount of substances present. In other words, the concentration of a solution will affects its boiling and freezing point. [49,1,3]
5. A The relative number of moles of solute in a solution affect the solution's properties. If you add 1.80 grams (0.010 mol) of glucose (C6H12O6, molar mass = 180 g/mol) in 1 kilogram of water, you will change the freezing point of the water from 0.0000 °C to –0.0186 °C. If you add ten times that amount (18.0 grams or 0.10 mol), the freezing point will also change by 10 times, to 0.186 °C. [49,2,1]
6. C Raoult's law, which explains colligative properties, works because solute particles prevent solvent molecules from entering the vapor phase. [49,2,3]
7. C Salts cause larger than expected changes in the point at which a solution's phase shifts, because they completely ionize in water. [49,2,4]
8. B The x axis on a phase diagram is temperature, and the y axis is pressure. A phase diagram shows the temperature and pressure at which a solution changes phase. [50,Figure 56]
9. E Adding salts to freezing roadways is an application of colligative properties, because the salts raise the freezing point of the water collecting on the roadways. [50,2,1]
10. C Distillation is possible because vapor does not contain solute. If the vapor can be captured, a pure substance results. [50,2,2]
11. A One mole of $CaCl_2$ will produce three moles of ions when it is added to water, because it completely ionizes. Therefore, its effect on the freezing and boiling points will be three times what would be expected. [50,2,0]
12. A Adding more solute to a solvent will shift the triple point up and right, meaning that a higher temperature and pressure will be required for the solution to change phases. [50,Figure 56]
13. E A Liebig condenser is a standard part of a laboratory water distillation set up. Distillation can be done on a small scale; all it requires is a way to capture vapor. [50,Figure 57]
14. C Compared to reverse osmosis, distillation is easier to scale up, has less need for shut-downs, less needed pretreatment, and less waste. [51,1,2]
15. E Reverse osmosis has lower energy needs, a lower discharge water temperature, purer water output, and smaller plant requirements than distillation. [51,1,2]

1. B Boyle's law states that the volume of gas in inversely proportional to its pressure. Increasing the pressure in a closed container will decrease the volume of the gas. [51,2,1]
2. E Charles's law states that the volume of a gas is directly proportional to its temperature. Increasing the heat of a gas will cause it to expand. [51,2,2]
3. C The Kelvin scale (K) is based on absolute value, the temperature at which all energy ceases in a thermodynamic system. [51,2,4]
4. C Dalton's law sets rules for dealing with mixtures of gases. It states that, in non-reacting gases, the total pressure exerted is equal to the sum of the partial pressures. [52,2,5]
5. A Avogadro's number, 6.022×10^{23}, is the number of particles that are in any one mole of a substance. [51,2,9]
6. C Avogadro proposed that every gas contains the same specific number of small particles under given conditions of temperature and pressure. [51,2,7]
7. B Several scientists, including Antoine Lavoisier, Carl Wilhelm Scheele, and Joseph Priestley, independently discovered oxygen in the 1770s. [51,2,11]
8. C The Van der Waals equation corrects the ideal gas law for attractions between particles and for volume, so it can be used to calculate the properties of gases under non-ideal conditions. [51,2,12]
9. E The particles in a gas have a freedom of motion that is not present in solids and liquids. [51,2,14]
10. B Kinetic-molecular theory rests on the idea that gas consists of very small, rapidly moving particles. This theory helps explain why matter behaves as it does. [51,2,6]
11. E Absolute zero is defined as the temperature at which a thermodynamic system has the lowest energy. It is -273 degrees on the Celsius scale. [51,2,3]
12. B Metals do not typically have a slimy feel, but they are malleable, ductile, lustrous, and conductive. [52,1,4]
13. D Compared to water, carbon dioxide has a more typical phase diagram. Water has unique properties on a phase diagram. [51,2,8]
14. A Joseph Black is credited with the discovery of carbon dioxide in 1754. [52,2,1]
15. B Percent composition, molarity, and molality all express how much solute dissolves in a solvent. [52,2,6]

1. A The five most basic types of chemical reactions are synthesis reactions, decomposition reactions, single replacement reactions, double replacement reactions, and combustion reactions. [53,1,2]
2. D In a chemical equation, the reactants, which are sometimes called the reagents, are on the left and the products are on the right. [53,2,0]
3. A In a synthesis reaction, two or more reactants combine to form one product. [53,2,2]
4. C Potassium reacts with chlorine to form potassium chloride. The chemical equation for this reaction is $2K(s) + Cl_2(g) \rightarrow 2KCl(s)$. [53,figure 58]
5. E Hydrogen peroxide, H_2O_2, decomposes to form water and oxygen. This reaction is represented by the equation $2H_2O_2 \rightarrow 2H_2O + O_2$. [54,1,1]
6. D Oxygen gas was first identified in 1774 when chemists observed the decomposition of mercury (II) oxide. [54,1,1]
7. B Double replacement reactions take the form $AB + CD \rightarrow AD + CB$. These reactions involve two ionic compounds in solution trading ions. [54,1,2]
8. D There are two possible formats for a single replacement reaction, dependent upon whether the atom is a metal or a nonmetal. One example of a single replacement reaction with a nonmetal is $Cl_2 + 2KBr \rightarrow 2KCl + Br_2$. [54,1,3]
9. D Single replacement reactions only occur if the atom is more reactive than the ion. The relative reactivities of different metals are shown in the activity series. [54,2,1]
10. C In this group of metals, barium is most reactive. Following barium in order of most to least reactive are magnesium, chromium, iron, and cobalt. [55,figure 63]
11. B Mercury, Hg, cannot displace H_2 in any source. Some compounds can displace H_2 from acid, some from steam, some from water, and some not at all. [55,figure 63]
12. B In combustion reactions with organic hydrocarbons, the hydrocarbon reacts with oxygen to form water and carbon dioxide. [54,2,2]
13. C Potassium can displace H_2 from water. Sodium and a few other metals can also displace H_2 from water. [55,figure 63]
14. C Silver is not reactive enough to displace H_2, so $Ag(s) + 2H^+(aq) \rightarrow$ no reaction. [55,figure 63]
15. E The chemical formula for silver nitrate is $AgNO_3$. Silver nitrate will react with copper to form copper nitrate and silver in the equation $2AgNO_3(aq) + Cu(s) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$. [55,figure 61]

FOCUSED QUIZ 28

SCIENCE

ANTOINE LAVOISIER, THE CONSERVATION OF MASS, AND THE BIRTH OF MODERN CHEMISTRY - TITRATIONS -- MIXING ACIDIC AND BASIC SOLUTIONS TO MEASURE THE CONCENTRATION OF AN UNKNOWN, PP. 56-59

1. B The goal of alchemy was to turn less valuable substances into substances of more value, like gold. While this pursuit was unsuccessful, alchemists did lay foundations for modern chemistry by focusing on observation and experimentation, and pioneering the processes of distillation, percolation, and extraction. [56,1]
2. D Antoine Lavoisier refuted phlogiston theory when he developed a scientific explanation for combustion. [56,4]
3. D The law of conservation of mass was founded on the observation that after heating mercury calx, the combined masses of mercury and gas were equal to the original mass of the calx. In the reverse of this experiment, the combined masses of the products and the combined masses of the reactants were also equal. [56,4]
4. E Antoine Lavoisier made many important contributions to the field of chemistry. He determined that air is composed of mostly nitrogen and oxygen, founded the law of conservation of mass, created a system of chemical nomenclature that is still used, and determined the role of oxygen in combustion. He also gave oxygen its current name. [57,3]
5. D The first theory of acid-base reactions was proposed by the Swedish scientist Svante Arrhenius. [57,1,2]
6. B In the current theory of acid-base reactions, a basic solution has an excess of hydroxide ions, OH^-. Conversely, acidic solutions have an excess of hydronium ions, H_3O^+. [57,2,0]
7. A A substance that is amphoteric can act as an acid or as a base. Water and ammonia are examples of amphoteric substances. [57,2,2]
8. A The pH scale is a logarithmic scale. The equation for pH is $pH = -\log[H_3O^+]$. [57,2,3]
9. E All acid and base solutions contain both acid and base. An acidic solution, or a solution with a low pH, contains mostly acid but some base. [58,1,1;58,1,2]
10. D Titrations are used to measure the concentration of an unknown. Most of the time, a base is added to an acid, but an acid may also be added to a base. [58,1,3;59,1,1;89,2,4]
11. C Basic compounds like NaOH or KOH often contain an OH^- ion. Acidic compounds like HNO_3 or HBr often have formulas beginning with hydrogen. [58,1,4]
12. B When a strong acid reacts with a strong base, a neutral salt and water form. [58,1,5]
13. D In a titration, a titrant is added using a buret to a solution called an analyte contained by an Erlenmeyer flask. The pH can be measured with a pH indicator, a compound that has different colors at different pH, or a pH meter. [58,2,0;58,2,1]
14. B In $M_aV_a = M_bV_b$, $_a$ represents the total acid and $_b$ represents the total base. The equation must be written for the total acid content because of the existence of polyprotic (diprotic or triprotic) substances, or substances that can donate more than one H^+ ion. [58,2,2]
15. E In $M_aV_a = M_bV_b$, M represents the molarity of each substance and V represents the volume of each substance. For the equation to be accurate, the volume must be measured in liters, which often means it is necessary to convert from mL to L. [59,1,1]

1. E Hydrolysis is the process in which a salt dissolves in water and its ions interact with the water molecules. The anion, the cation, both ions, or no ions can react with the water. [59,1,2]
2. A In a basic salt reaction, the anion of the salt reacts with the water molecules. This reaction produces the conjugate acid of the salt and hydroxide ions, OH^-. [59,1,2]
3. E Acetic acid is the conjugate acid of acetate. The acetate ion has the formula $C_2H_3O_2^-$ (often written as CH_3COO^-), and acetic acid has the formula CH_3COOH. [59,1,3]
4. D Acid-base reactions form a salt; either basic, acidic, or neutral, and water. Anions of weak acids form basic salts, cations of weak bases form acidic salts, and the anion of a strong acid or the cation of a strong base form neutral salts. [59,1,3;59,1,4;59,2,1]
5. C HCO_3^-, the bicarbonate ion, is an anion that will react to form a basic salt. Other anions that form basic salts are CH_3COO^-, F^-, and SO_4^{2-}. [59,2,2]
6. D Cu^{2+} is a cation that will react to form an acidic salt. Other cations that form acidic salts are NH_4^+ and $CH_3NH_3^+$. [59,2,2]
7. C Precipitation reactions are a specific type of double replacement reaction where one or both products are insoluble and precipitate out of solution. [59,2,3]
8. A One sign that a precipitate has formed is a solution becoming cloudy. After the solution becomes cloudy, the precipitate may settle to the bottom. [59,2,3]
9. C Silver nitrate and sodium chloride react in the equation $AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$ to form a precipitate of AgCl and aqueous $NaNO_3$. The net ionic equation for this reaction is $Ag^+ + Cl^- \rightarrow AgCl$. [60,1,0]
10. D Silver nitrate and sodium chromate react in the equation $2AgNO_3 + Na_2CrO_4 \rightarrow Ag_2CrO_4 + 2NaNO_3$ to form solid silver chromate and aqueous sodium nitrate. Na^+ and NO_3^- are the spectator ions in this reaction as they do not change. [60,figure 65]
11. A According to the solubility rules, phosphates are insoluble. Since $Zn_3(PO_4)_2$ contains phosphate, PO_4^{3-}, it is insoluble. [60,1,2]
12. A All sulfates are soluble, except for sulfate compounds with barium, strontium, calcium, lead (II), silver, and mercury (I). [60,2,0]
13. E The solubility of metal hydroxides depends on the OH^- concentration. When the OH^- concentration, and therefore pH, is high, OH^- ions react with the metal hydroxide and forms a soluble oxide ion. [60,2,2]
14. E Mercury (II) turns into CH_3Hg in the aquatic environment through a process called methylation. Methylation is a specific type of alkylation. [61,figure 67]
15. B Solubility principles are applied to control soil pH, evaluate the aquatic precipitation of mercury, and recover silver waste from photographic chemicals. Mine run-off is a source of harmful precipitates, as are some atmospheric particles. [61,1,0;61,2,0;62,1,0;62,1,2]

1. A An atom that has oxidized has lost one or more electrons. On the other hand, atoms that have been reduced have gained one or more electrons. [62,1,3]
2. B Batteries are powered by the oxidation-reduction reactions that occur in an electrochemical cell. [62,2,1]
3. D An oxidation number is the number of electrons that must be added or subtracted from an atom in its combined state to reach the number of electrons in its neutral state. Oxidation numbers may also be known as oxidation states. [62,2,2]
4. B Hydrogen always has an oxidation number of +1 except in metal hydrides, where it has an oxidation number of -1. [62,2,3]
5. A Oxygen's oxidation number is usually -2. However, in peroxides, it is -1 instead. [62,2,3]
6. B Oxidation numbers have some general rules, including that the oxidation number of a monatomic ion is equal to the ion's charge, neutral atoms have oxidation numbers of 0, the oxidation numbers in a neutral compound add up to 0, and the oxidation numbers in polyatomic ions add up to the charge of the ion. Oxidation numbers help track the movement of electrons in reactions. [62,2,5]
7. B Because K is a group one element, its oxidation number is +1. This compound is not an exception to the rule that O's oxidation number is -2, so 2(1) + 4(-2) +1(x) = 0, and x = +6. [63,1,0]
8. C Positive ions are called cations and negative ions are called anions. [63,2,1]
9. D To form a complete equation, two half-reactions are added, but the equations must first be multiplied so they contain the same number of electrons. 4 x $(Al \rightarrow Al^{3+} + 3e^-)$ = $4Al \rightarrow 4Al^{3+} + 12e^-$ 3 x $(O_2 + 4e^- \rightarrow 2O^{2-})$ = $3O_2 + 12e^- \rightarrow 6O^{2-}$ $4Al \rightarrow 4Al^{3+} + 12e^-$ +$3O_2 + 12e^- \rightarrow 6O^{2-}$ = $4Al + 3O_2 \rightarrow 4Al^{3+} + 6O_2^-$ [64,1,2;64,2,1]
10. E Chrome plating, short for chromium plating, is one of the most common types of electroplating. [65,1,0]
11. E The tendency of an electron to leave or join an atom is measured in volts, a unit of electrical energy. This unit was named after Alessandro Volta. [65,2,1]
12. C All half-cell potentials are measured relative to the reduction reaction of hydrogen, $2H^+ + 2e^- \rightarrow H_2$. [66,1,1]
13. A A reaction's possible voltage is equal to $E°_{red}$ - $E°_{ox}$, or the difference between the reduction potential of the reduction reaction ($E°_{red}$) and the reduction potential of the oxidation reaction ($E°_{ox}$). [66,1,3]
14. E $Li^+ + Fe^{3+} \rightarrow Li + Fe^{2+}$ is a nonspontaneous reaction, meaning that it does not naturally occur. [66,2,1]
15. C J. F. Daniell invented the galvanic cell that was powered by the reaction $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$ in 1836. This cell was the earliest reliable battery. [66,2,3]

1. C Luigi Brugnatelli is most often credited with developing the modern process of electroplating. In 1805, he noticed that if he attached metal to the negative end of a battery and submerged it in a gold salt solution, the metal would be coated with a thin layer of gold. [66,1]
2. E Luigi Galvani noticed when dissecting frogs that the frogs hung on metal hooks twitched if touched with a metal object, and he believed this was because of "animal electricity". In 1799, Alessandro Volta conducted experiments that showed that the electricity was generated by the metals, disproving Galvani's "animal electricity" theory. [67,0;67,1]
3. B In electrochemical reduction reactions, positively charged metal ions are attracted to a battery's negative electrode, which donates electrons, turning the metal into its neutral species, causing it to precipitate out of the solution and form a coating. Electroplating was discovered in the early 1800s but was not widely used until the 1840s. [67,1]
4. C Electroplating has numerous benefits in a myriad of applications. In the auto industry, plating provides added hardness, corrosion protection, wear resistance, and enhanced appearance, but it does not increase magnetism. [67,2]
5. A Electroplating is often used in the aerospace industry for its cooling properties. For example, it was used in the construction of the shuttle *Columbia*, which featured 41 kg of gold. [67,3]
6. A Gold, silver, nickel, copper, platinum, and palladium are all used in electroplating. Sodium is not suitable for electroplating because it forms hydrogen gas when used as an electrolyte. [67,4]
7. A The hexagonal component mirrors of NASA's James Webb Space Telescope are plated with gold. This plating helps the mirrors to reflect infrared light optimally. [67,1,image]
8. D Palladium coating results in a finish that looks like white gold and is often used in plating expensive watches. [68,0]
9. A Platinum finishes resist corrosion, so they are used to plate sterling silver and surgical instruments. [68,0]
10. E Because one mole of Cr^{3+} has lost three moles of electrons, it takes 3 Faradays to convert it to its elemental form. 1 Faraday is the electrical charge on one mole of electrons. [68,1]
11. D German chemist Walther Nernst developed the Nernst equation to predict the results of oxidation-reduction reactions. [68,1,1]
12. D The most fundamental form of the Nernst equation is $E = E° - RT \ln Q / nF$, but the most useful form is $E = E° - (0.0592 \log Q) / n$. In this version, the natural logarithm is changed to a base ten logarithm, temperature is assumed to be 298K, and R is included in correct units. [68,2,2]
13. E The symbol ° denotes a value that is under standard state conditions. For example, E represents cell potential and E° represents cell potential under standard state conditions. [68,2,1]
14. E When the concentration of reactants and products are equal, $E = E°$. If the concentrations are equal, $Q = 1$ and $\log 1 = 0$, so $E = E°$. [69,1,0]
15. C A concentration cell is an electrochemical cell that has different concentrations of the same solution in each side. In a concentration cell, $E° = 0$ even though there will be a cell voltage. [69,1,0]

1. E Stoichiometry tracks the movement of atoms in a chemical reaction and establishes and uses mass-mole relationships in these reactions. Doubling a recipe is analogous to the process of stoichiometry. [69,1,1]
2. D When hydrogen burns in air, water forms. This reaction is represented by the equation $2H_2 + O_2 \rightarrow 2H_2O$. [69,1,2]
3. D Net ionic equations exclude spectator ions, or ions that do not change or form precipitates during a reaction. [69,2,0]
4. E The reaction between silver nitrate and sodium sulfide is an example of a double replacement reaction. In this reaction, nitrate and sulfide switch places. [69,2,1]
5. B The net ionic reaction for silver nitrate and sodium sulfide is $2Ag^+(aq) + S^{2-}(aq) \rightarrow Ag_2S(s)$. The complete reaction is $2AgNO_3(aq) + Na_2S(aq) \rightarrow 2NaNO_3(aq) + Ag_2S(s)$. [69,2,2]
6. E The spectator ions in the reaction between silver nitrate and sodium sulfide are NO_3^- and Na^+. [69,2,2]
7. A An empirical formula is the lowest mole ratio of atoms in a formula. Acetylene's empirical formula is CH, but its molecular formula is C_2H_2, as CH is an unstable molecule. [69,2,3]
8. D The molar mass of water is 18 g/mol because each hydrogen has a molar mass of 1 g/mol and oxygen has a molar mass of 16 g/mol. $2(1) + 16 = 18$ g/mol. [70,1,0]
9. C In the reaction between hydrogen and oxygen, two moles of hydrogen react with one mole of oxygen. So for 6 moles of hydrogen to react fully, 3 moles (half of 6) of oxygen are needed. [70,2,1]
10. B The molar mass of carbon dioxide is 44 g/mol. Each oxygen atom has a molar mass of 16 g/mol and carbon has a molar mass of 12 g/mol, $12 + 2(16) = 44$ g/mol. [70,1]
11. E Proust's law is also known as the law of definite proportions. It states that elements combine in fixed ratios to form compounds. [70,2]
12. A Proust's law is based on his studies of the composition of water and copper carbonate in the 1790s and early 1800s. [70,2]
13. A When Joseph Proust published his findings on the theory that became known as his law, several scientists disagreed with it. One of these scientists was the French chemist Claude Louis Berthollet. [70,2]
14. E Though some scientists disagreed with Proust's theory, others, like John Dalton, agreed with it. Dalton's theory both supports and is supported by Proust's law. [70,2]
15. E John Dalton proposed the first modern theory of atomic behavior in the early 1800s. [70,2]

SCIENCE

FOCUSED QUIZ 33

QUILIBRIUM - SOLUBILITY PRODUCT CONSTANTS AND THEIR APPLICATION TO PRECIPITATION AND THE DISSOLUTION OF SLIGHTLY SOLUBLE COMPOUNDS, PP. 71-72

DemiDec® We do our best, so you can do yours.

1. C For the reaction A + B ↔ C + D, $K_c = \frac{[C][D]}{[A][B]}$. More generally, the reaction aA + bB ↔ cC + dD kas the equilibrium constant $\frac{[C]^c[D]^d}{[A]^a[B]^b}$. [71,1,5]
2. E [A] represents the concentration of compound A in moles per liter. [71,2,0]
3. A The reaction A + B ↔ 2C has the equilibrium constant $\frac{[C][C]}{[A][B]}$, which simplifies to $\frac{[C]^2}{[A][B]}$. [71,2,1]
4. A K_p is the equilibrium constant used for a reaction that involves gas pressures. K_c is the equilibrium constant for a reaction involving solutions. [71,2,3]
5. B As a K value increases, the reaction increasingly favors the products. The largest K value in this list is 1.6×10^{12}, so it belongs to the reaction in which products are favored most. [71,2,4]
6. C Molarity, M, is a measure of concentration measured in moles per liter. [71,2,5]
7. A Pure liquids and pure solids are never included in equilibrium expressions. Since water is a pure liquid, it is not included in K values. [72,1,0]
8. D A concentration value of a solution can also be called the solution's activity. Activity is defined as the ratio of the concentration of a substance to its standard state concentration. [72,1,1]
9. D K values are unitless because they come from a calculation where the units cancel out. [72,1,1]
10. E Standard state aqueous solutions have a concentration of 1 M. In gases, the standard state is at 1 atmosphere of pressure. [72,1,1]
11. C Activity is calculated by dividing concentration by standard state concentration. In pure substances, concentration equals standard state concentration, so activity equals 1. [72,1,1]
12. A To convert between K_p and K_c, chemists use the equation $K_p = K_c(RT)^{\Delta n}$. When the change in moles (Δn) is zero, $K_p = K_c$. [72,1,2]
13. B Vinegar is aqueous acetic acid, $HC_2H_3O_2$. It is considered a weak acid. [72,2,1]
14. A As the value of K_a or K_b increases, so does the strength of the acid or base, respectively. The stronger the acid or base, the more of it reacts. [72,2,1;72,2,2]
15. E The smaller the K_{sp} value, the less soluble a compound is, and 1×10^{-10} is the smallest non-negative number of these values. K_{sp} values are never negative because they are calculated with concentration, which are also never negative. [72,2,3]

1. D Kinetics is the study of reaction rates and which factors, like temperature and catalysts, affect them. [73,1,1]
2. D Rate laws show the dependence of the reaction rate on concentrations. Each reaction has a different rate law. [73,1,1]
3. B The speed of an object is the distance traveled divided by the time elapsed. The speed is then the rate of change of an object's position. [73,1,2]
4. D The collision model is the best model for chemical reactions, and it is similar to the Kinetic Theory. The collision model states that reaction rates depend on the number of collisions and the energy at which they happen. [73,1,3]
5. B One way to find an instantaneous rate of reaction is to graph the concentration of a substance over time and calculate the slope of the tangent line at the desired time. [73,1,2]
6. A Alka-Seltzer tablets dissolve faster in hot rather than cold water and broken rather than intact. Alka-Seltzer tablets are a type of antacid tablet. [73,2,1;74,1,0]
7. E E_a stands for activation energy, and it is one factor that affects the rate of reactions. [74,1,2]
8. B Activation energy is the energy needed for reactants to transform into products. The rate of reaction depends in part on activation energy. [74,1,2]
9. B Catalysts affect reaction rates by providing a new pathway for the reaction to take place and therefore changing the necessary activation energy. [74,1,4]
10. E The vertical axis on a potential energy diagram is free energy and the horizontal axis is reaction progress. [73,figure 73]
11. C Potential energy diagrams can show whether a reaction is exothermic or endothermic and how quickly a reaction will take place. [88,1,5]
12. A ΔG represents a change in free energy or Gibbs free energy. On potential energy diagrams, it represents the change from the reactants' energy to the products' energy. [73,figure 73]
13. C Reaction rate decreases as activation energy increases, as collisions are less likely to occur successfully when only a few collisions have the required activation energy. [74,1,3]
14. C In catalyzed reactions, activation energy changes, but free energy does not. [73,figure 74]
15. E Enzymes are catalysts in the human body, and they increase reaction rates at body temperature. [74,1,5]

1. B In chemistry, a system is the specific substance being studied, and the surroundings are everything else. In this example, the solution is being studied, so it is the system. [74,2,1]
2. B Endothermic reactions are reactions where heat is absorbed. System + surroundings = universe, and the universe's energy is constant, so for a system's energy to increase, the surroundings' energy must decrease. [74,2,2]
3. C In addition to his work in thermochemistry, Joseph Black discovered carbon dioxide in 1754. [75,1]
4. B Joseph Black noticed that even after air temperature is higher than freezing, snow doesn't all melt immediately. This observation led him to his experiments in the 1750s. [75,1]
5. B Joseph Black observed the temperatures of 5 ounces of water and 5 ounces of ice over time in a room with a temperature of 47°F. [75,2]
6. D Joseph Black performed experiments with 5 ounces of water and 5 ounces of ice. The water started at 33°F and ended at 40°F. [75,2]
7. B In Joseph Black's experiments on thermochemistry, 5 ounces of water took 30 minutes to reach its final temperature, whereas 5 ounces of ice took ten hours and thirty minutes to reach the same temperature. [75,2]
8. A In Joseph Black's experiments, by the time the ice reached its final temperature, it had absorbed 147 units of heat. (40 – 33) x 21 = 147. [75,2]
9. E Latent heat is the heat released or absorbed by a substance when it undergoes a phase change without a change in temperature. Joseph Black understood this concept and lectured on it. [86,1,12]
10. A A substance's heat of vaporization is also known as the latent heat involved in vaporizing a liquid or condensing a gas. [75,3]
11. D James Watt, a friend and student of Joseph Black, invented the steam engine based on the concept of latent heat. [75,3]
12. E The invention of the steam engine drove the Industrial Revolution and was based on the concept of latent heat. [75,3]
13. C The first ice-calorimeter was used by Antoine Lavoisier and Pierre-Simon Laplace in 1782 and 1783. [75,image]
14. A A state function is a quantity determined by its final and initial states, independent of the path taken to reach the final state. The opposite of a state function is a path or process function, where the path taken does matter. [76,1,1]
15. D Some examples of state functions are volume, pressure, and energy. ΔV, ΔP, and ΔE are also state functions. [75,1,1]

1. C Enthalpy is defined as the energy content of a system and it is directly proportional to the amount of chemical present. [76,1,2]
2. D In calorimetry, a reaction's temperature change is compared to that of 1 gram of water that has been raised by 1 degree Celsius. This value for water is 4.18J. [76,1,4]
3. E Hess's law states that the heat change of a reaction is the same whether it takes place in one or more steps. His law is also known as the law of constant heat summation. [76,2,0]
4. C Subtracting $CO + 1/2O_2 \rightarrow CO_2$ from the equation $C + O_2 \rightarrow CO_2$ yields the equation $C + 1/2O_2 \rightarrow CO$. The first equation's enthalpy value is -393 kJ/mol and the second equation's enthalpy is -283 kJ/mol, so the enthalpy of the combine reaction is -111 kJ/mol as -393 – (-283) = -111. [76,2,2]
5. C Negative enthalpy values indicate that a reaction is exothermic, or that it has released energy. Positive enthalpy values indicate an endothermic reaction, where energy is absorbed. [76,2,2]
6. E The symbol for entropy is S, the symbol for enthalpy is H, and the symbol for Gibbs free energy is G. [76,2,3]
7. D All natural processes increase the entropy of the universe. In other words, nature favors disorder. [76,2,4]
8. D A decrease in a system's entropy occurs if the surroundings experience a matching increase in entropy, because system + surroundings = universe. [76,2,4]
9. D In order of lowest to highest entropy, or increasing disorder, the states of matter are solids, liquids, solutions, and gases. [76,2,4]
10. C For the Gibbs free energy equation to be correct, temperature must be measured in Kelvin. A temperature in Kelvin is equal to the corresponding temperature in Celsius plus 273. [77,1,1]
11. A One way of calculating ΔG is through the equation $\Delta G = \Delta H - T\Delta S$. Additionally, under standard conditions, $\Delta G = -RT\ln K = -nFE°$. [77,1,1]
12. E Reactions that have a positive ΔH value and a negative ΔS value are always nonspontaneous in the forward direction but spontaneous in the reverse direction. [77,1,2;77,figure 76]
13. B When ΔH is negative and ΔS is positive, a reaction is spontaneous in the forward direction but nonspontaneous in the reverse direction at all temperatures. [77,1,2;77,figure 76]
14. E When ΔG has a negative sign, the reaction is always spontaneous. Conversely, if ΔG has a positive sign, the reaction is nonspontaneous. [77,1,2;77,figure 76]
15. E Reactions that are spontaneous at only low temperatures have negative ΔH and ΔS values, like the reaction $NH_3 + HCl \rightarrow NH_4Cl$. Reactions with negative ΔH and ΔS values are also nonspontaneous at high temperatures. [77,1,2;77,figure 76]

1. E Gibbs free energy is directly related to both equilibrium constants and electrode potentials. The signs of ΔG and E reveal which way a reaction proceeds. [77,2,1]
2. C The E and ΔG values of a reaction approach zero as the reaction approaches equilibrium. Therefore, at equilibrium, $\Delta G = E = 0$. [77,2,1]
3. C If a reaction has a negative ΔG value, its galvanic cell voltage is positive. Gibbs free energy and galvanic cell potential have opposite signs. [77,2,1]
4. C A positive galvanic cell voltage value indicates that a reaction proceeds spontaneously in the forward direction. In other words, the reaction favors the formation of products. [77,2,1]
5. E In order to compare reactions effectively, chemists must perform calculations using standard state values. The symbol "°" denotes a substance in its standard state. [77,2,2]
6. C The equilibrium constant, K, can be determined directly using the values of ΔG and E. These values also allow for the determination of the sign of the cell potential. [77,2,2]
7. D An equilibrium constant, K, is considered large if it is greater than one. Conversely, it is small if it is less than one. [77,2,2]
8. C Chemists use the equation $\Delta G° = -RT\ln K$ to convert between free energy, $\Delta G°$, and equilibrium constant, K. This equation can be rewritten as $K = e^{-\Delta G°/RT}$. [77,2,3]
9. A In order to change between free energy to standard cell potential, scientists use the equation $\Delta G° = -nFE°_{cell}$. In this equation, n represents the number of moles. [77,2,3]
10. D In the equations used for changing from free energy to equilibrium constants and standard cell potential, in order for the signs of K and E to be correct, one side of the equation must have a negative sign. [77,2,3]
11. A Galvanic cells always spontaneously proceed in the forward direction, so their E values are always positive. K values are positive in all reactions because they're computed using concentrations, which are also always positive. [77,2,3]
12. C In a reaction where K is greater than 1, the reaction proceeds spontaneously in the forward direction, or favors the products. [78,figure 77]
13. A In a reaction where $E°_{cell}$ is negative, reactant formation is favored. Additionally, when $E°_{cell}$ is negative, K is greater than zero but less than one. [78,figure 77]
14. B If the $\Delta G°$ is negative, the reaction favors the formation of the reactants and proceeds in the reverse direction. [78,figure 77]
15. B Whenever K = 1, the reaction is at equilibrium. In other words, the formation of reactants and products are equally favored. [78,figure 77]

1. C Combustion reactions are reactions between a substance and oxygen. If the other reactant is an organic molecule, water and carbon dioxide will be the products. [78,1,6]
2. B Antoine Lavoisier's experiments led to more understanding of combustion reactions, the law of conservation of mass, and the naming of oxygen. [78,1,7]
3. D The pH scale ranges roughly from 0 – 14, and substances with a pH of 7 are neutral. Values less than 7 are acidic (acidity increases as pH decreases) and values greater than 7 are basic (as pH increases, the substance is more basic). [78,2,1]
4. A H_3O^+ is the hydronium ion and pH = -log[H_3O^+]. H_3O^+ is sometimes replaced by the hydrogen ion, H^+. [78,2,2]
5. C A substance that can act as either an acid or a base is amphoteric. Water and ammonia are examples of amphoteric substances. [78,2,3]
6. D In a titration, a buret is used to measure how much titrant (the solution contained by the buret) is added to the analyte (the solution with unknown pH). [78,2,4]
7. E In titrations, when the amount of acid equals the amount of base, the equivalence point has been reached. The equivalence point is also known as the endpoint. [78,2,5]
8. E Acid-base reactions produce a salt and water. The salt produced can be acidic, basic, or neutral, depending on the strength of the acid or base that reacted. [78,2,6]
9. A Electrochemical cells rely on electrochemical reduction reactions, which are a specific type of oxidation-reduction (redox) reaction. [79,1,3]
10. D The Nernst equation has a few forms, but the most useful one is $E = E° - (0.0592 \log Q)/n$. This equation connects cell potentials to changes in free energy. [79,1,8]
11. D For a generic reaction, $aA + bB \rightarrow cC + dD$, $K = \frac{[C]^c[D]^d}{[A]^a[B]^b}$. K is the reaction's equilibrium constant. [79,2,4]
12. E K_c, K_p, K_a, K_b, and K_{sp} are the equilibrium constants for reactions involving concentration, gas pressures, acids, bases, and precipitates, respectively. [79,2,6]
13. B If a reaction's K value is less than 1, it favors the reactants. If K is greater than 1, it favors the products and if K is equal to one, it favors neither the reactants nor the products. [79,2,6]
14. A State functions are functions where the pathways in irrelevant (they only depend on the initial and final values). Some examples of state functions are volume, pressure, temperature, entropy, enthalpy, and free energy. [80,1,5]
15. D If $\Delta G° = 0$, a reaction is at equilibrium. If it is negative, the reaction is spontaneous, but if it is positive, the reaction is nonspontaneous. [80,2,4]

1. C Acidic solutions have an excess of H^+, the hydrogen ion, whereas basic solutions have an excess of OH^-, the hydroxide ion. [82,1,3]
2. E Alchemy, or the practice of turning metals into a more valuable metals like gold or silver, was particularly popular during the Middle Ages. [82,1,6]
3. C An alpha particle is a helium nucleus, represented as . [82,1,10]
4. D An atomic mass unit may also be called a Dalton. 1 amu (1 Dalton) is 1/12 the mass of a carbon-12 atom. [82,2,7]
5. B Avogadro's number is 6.022×10^{23}. It represents the number of particles in one mole. [82,2,11]
6. C A binary substance is a compound composed of 2 elements. NaCl is an example of a binary substance. [83,1,4]
7. C Ductility is the property of metals that allows them to be pulled into wires. Relatedly, malleability is the ability of metals to be flattened into sheets. [84,1,12]
8. E J. J. Thompson discovered the electron in 1896. [84,2,3]
9. C Enthalpy is a measure of a substance's heat content under constant pressure, represented by the symbol H. [85,1,1]
10. E The ideal gas law equation is $PV = nRT$, where P = pressure, V = volume, n = number of moles, R = the gas constant, and T = temperature. [85,2,9]
11. E Water boils at 373 Kelvin, 100°C, or 212°F. [86,1,8]
12. E Molarity is measured in moles/liter. Molality measures concentration in moles/kilogram. [87,1,2]
13. A Steel is an alloy composed of mainly iron and carbon, but may also contain nickel, chromium, cobalt, molybdenum, or zirconium. [89,1,10]
14. E A triprotic acid is an acid that can donate three hydrogen ions. With 3 hydrogens, H_3PO_4 is triprotic. [89,2,9]
15. C Volts are a unit of electrical potential. The unit was named after Alessandro Volta. [90,2,1]

1. A Elements in the same group have similar properties. Because silicon and germanium are both in group 14, their properties will be similar. [91]
2. A Elements in the periodic table are arranged in order of their numbers of protons, which often corresponds to their atomic mass. [91]
3. B Group 18 or VIIA elements are known as the noble or inert gases because they have low reactivity. [91]
4. A The diatomic elements include nitrogen, iodine, and bromine, to name a few. [91]
5. E Tantalum, Ta, and tungsten, W, are located directly besides each other on the periodic table. [91]
6. B Between the alkaline earth metals and the post-transition metals are the transition metals, the largest family on the periodic table. [91]
7. D There are a total of 118 elements on the periodic table. Out of these, 24 are synthetic. [91]
8. A Alkali metals are group 1 on the periodic table. Some alkali metals are lithium, potassium, and cesium. [91]
9. D Phosphorous is considered a polyatomic nonmetal. Sulfur, selenium, and carbon are also polyatomic nonmetals. [91]
10. B Manganese has the elemental symbol Mn. Magnesium has the elemental symbol Mg. [91]
11. A Rubidium, Rb, is part of group 1. Group 1 is the lowest group on the periodic table. [91]
12. B Period 7 and period 6 each include 32 elements. Copernicium is a group 7 element. [91]
13. C The most massive element on the periodic table is Oganesson, Og, with a mass number of 294. [91]
14. C Period 7 is the last period on the periodic table. It contains all the synthetic elements and most of the rare natural elements that are usually produced artificially. [91]
15. A Hydrogen is a diatomic nonmetal, meaning hydrogen molecules consist of two hydrogen atoms. [91]