SCIENCE

FOCUSED QUIZ 01

THE STRUCTURE OF MATTER –
THE DETERMINATION OF ATOMIC MASSES, PP. 6-8

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1. Generally, the term "Scientific Revolution" refers to advancements from the
    a. sixteenth and seventeenth centuries
    b. fourteenth and fifteenth centuries
    c. eighteenth and nineteenth centuries
    d. fifteenth and sixteenth centuries
    e. seventeenth and eighteenth centuries

2. Historians dislike the term "Scientific Revolution" for of all of the following reasons EXCEPT
    a. it did not take place in only one location
    b. it did not occur in one discrete timeframe
    c. the changes it brought were not as revolutionary as had once been believed
    d. it did not consist of one single event
    e. the Church was still embroiled in scientific discoveries

3. Which scientist proposed that the four elements were earth, air, water, and fire?
    a. Aristarchus
    b. Aristotle
    c. Galileo
    d. Archimedes
    e. Vesalius

4. In the early seventeenth century, the field of chemistry MOST overlapped with that of
    a. philosophy
    b. astrology
    c. mesmerism
    d. alchemy
    e. biology

5. Empiricism emphasized the use of
    a. religion
    b. evidence
    c. reason
    d. deduction
    e. justification

6. The creation of scientific societies prompted all of the following changes EXCEPT
    a. the exposition of mistakes
    b. the replication of experiments
    c. the privatization of discoveries
    d. the review of new theories
    e. the publication of research

7. The smallest distinct particles that make up matter are
    a. compounds
    b. molecules
    c. alloys
    d. atoms
    e. fermions

8. An ion forms when an atom loses or gains a(n)
    a. proton
    b. neutron
    c. electron
    d. neutrino
    e. positron

9. How did chemists develop a table of atomic masses?
    a. by measuring the wavelength emitted by each element
    b. by measuring the energies of electrons in each element
    c. by measuring the mass ratios of elements in a compound
    d. by measuring the number of protons and neutrons in each atom
    e. by measuring absorption spectra of elements in a compound

10. Mass spectrometers are sometimes used to
    a. analyze the coatings of non-stick pans
    b. create emission spectra
    c. sanitize surgical instruments
    d. search for explosives
    e. determine chemical levels of a pool

11. A mass spectrometer's results depend upon the interaction of
    a. velocity and temperature
    b. mass and charge
    c. mass and velocity
    d. charge and distance
    e. distance and intermolecular force

12. A mass analyzer functions to
    a. make substances gaseous
    b. create a vacuum
    c. accelerate ions into different pathways
    d. measure the number of ions present in each mass
    e. convert atoms to ions

13. Why do some mass spectrometers use a curved magnet?
    a. to repel negative ions
    b. to make the mass spectrometer more compact
    c. to change the shape of the ion path
    d. to create a cycling magnetic field
    e. to trap electrons

14. Samples for mass spectrometers with straight pathways are generally
    a. liquid
    b. superfluid
    c. plasma
    d. gaseous
    e. solid

15. How does a mass spectrometer scan several different masses?
    a. by altering the ion repeller
    b. the ion detector
    c. by altering the amplifier position
    d. by altering the magnetic field
    e. by altering the Faraday collector

1. Of which nationality was John Dalton?
   a. Scottish
   b. American
   c. German
   d. English
   e. French

2. Who FIRST proposed the concept of atomism?
   a. Leucippus
   b. John Dalton
   c. Amedeo Avogadro
   d. Democritus
   e. Dmitri Mendeleev

3. John Dalton proposed his atomic theory during the
   a. late nineteenth century
   b. early eighteenth century
   c. early nineteenth century
   d. mid-nineteenth century
   e. late eighteenth century

4. Which branch of science MOST interested John Dalton?
   a. hydrology
   b. geology
   c. astronomy
   d. meteorology
   e. ecology

5. For how many years did John Dalton keep daily scientific records?
   a. 57
   b. 39
   c. 61
   d. 25
   e. 43

6. John Dalton focused MOST on the behavior of
   a. metalloids
   b. metals
   c. liquids
   d. gases
   e. salts

7. John Dalton found that if two liquids at constant pressure were subjected to the same temperature change, their
   a. volumes would vary by the same amount
   b. vapor pressures would vary proportionally to the strength of the bonds in each substance
   c. bond energies would vary by the same amount
   d. molecular velocities would vary by the same amount
   e. densities would vary proportionally to the strength of the bonds in each substance

8. The law of partial pressures is BEST represented by the equation
   a. $PV = nRT$
   b. $P_1V_1 = P_2V_2$
   c. $P_{total} = P_A + P_B$
   d. $V_1/T_1 = V_2/T_2$
   e. $P_1V_1/T_1 = P_2V_2/T_2$

9. Who proposed the law of conservation of mass?
   a. Antoine Lavoisier
   b. Isaac Newton
   c. John Dalton
   d. Amedeo Avogadro
   e. Joseph Louis Proust

10. John Dalton proposed the law of multiple proportions based on the works of
    a. Joseph Louis Gay-Lussac
    b. Henry Louis Le Chatelier
    c. Louis Pasteur
    d. Jacques Charles
    e. Joseph Louis Proust

11. How many years separated the proposal of the law of definite proportions and the law of conservation of mass?
    a. 8
    b. 10
    c. 15
    d. 5
    e. 12

12. John Dalton's atomic theory was MOST closely related to the law of

    a. law of conservation of energy
    b. conservation of mass
    c. partial pressures
    d. multiple proportions
    e. definite proportions

13. Which of the following major conclusions did NOT appear in *A New System of Chemical Philosophy*?

    a. All matter is composed of atoms.
    b. Atoms of different elements combine in ratios to form compounds.
    c. Atoms of different elements have different properties.
    d. Atoms cannot be further divided.
    e. Atoms in the same family have similar properties.

14. Why was *A New System of Chemical Philosophy* particularly significant?

    a. It contained the first description of atomism.
    b. All its conclusions remain true.
    c. It was rooted in experimentation.
    d. It led to the development of the periodic table.
    e. It gave credit to a female lab assistant.

15. Which technology did scientists use to prove that different isotopes of the same atom can have different atomic masses?

    a. scanning electron microscopy
    b. photoelectron spectroscopy
    c. titration
    d. mass spectroscopy
    e. gas chromatography

1. Which of the following subatomic particles have a negative charge?
   a. gravitons
   b. electrons
   c. positrons
   d. protons
   e. neutrons

2. Which element does NOT contain any neutrons?
   a. oxygen
   b. nitrogen
   c. helium
   d. carbon
   e. hydrogen

3. The symbol "n^0" represents a
   a. neutrino
   b. nuclide
   c. nucleus
   d. neutron
   e. nucleotide

4. If Alpacatinium has elemental symbol Ac, atomic number 89, and mass number 227, which symbol represents it?
   a. ${}^{89}_{227}Ac$
   b. ${}^{138}_{89}Ac$
   c. ${}^{227}_{89}Ac$
   d. ${}^{227}_{138}Ac$
   e. ${}^{316}_{89}Ac$

5. The atomic number denotes the
   a. number of protons in an atom's nucleus
   b. sum of the number of neutrons and protons in an atom
   c. most common oxidation state of an atom
   d. average mass of all the isotopes of an element
   e. distance from the center of an atom's nucleus to its farthest electron

6. Which of the following terms represents an arrangement of protons and/or neutrons forming a nucleus?
   a. hypernucleus
   b. nuclide
   c. isotopomer
   d. isotope
   e. atom

7. How many naturally occurring isotopes does carbon have?
   a. 6
   b. 3
   c. 2
   d. 5
   e. 4

8. Radioactive atoms are BEST described as atoms that
   a. have lost or gained one or more electrons
   b. exhibit luminescence in the absence of light
   c. are unstable and decay over time
   d. form octahedral bond angles
   e. combine atomic orbitals to form hybrid orbitals

9. Why was carbon established as the standard for atomic mass?
   a. Carbon, a solid, is more convenient to use for comparison purposes than gaseous hydrogen.
   b. Carbon is the naturally occurring element with the least isotopes.
   c. Carbon was the first element whose atomic mass was measured.
   d. Carbon is found more abundantly in nature than hydrogen.
   e. Carbon has many inexpensive, accessible, and easy-to-transport allotropes.

10. An atom of tritium is represented by
    a. ${}^{1}_{1}H$
    b. ${}^{2}_{1}H$
    c. ${}^{1}_{3}H$
    d. ${}^{3}_{1}H$
    e. ${}^{2}_{3}H$

11. Which nuclide is MOST commonly used in fusion reactors?
    a. ${}^{1}_{3}H$
    b. ${}^{1}_{1}H$
    c. ${}^{14}_{6}C$
    d. ${}^{13}_{6}C$
    e. ${}^{2}_{1}H$

12. Which element interacts with cosmic rays to form ${}^{14}_{6}C$?
    a. helium
    b. oxygen
    c. nitrogen
    d. argon
    e. hydrogen

13. Which of the following isotopes is radioactive?
    a. ${}^{60}_{27}Co$
    b. ${}^{1}_{1}H$
    c. ${}^{2}_{1}H$
    d. ${}^{13}_{6}C$
    e. ${}^{59}_{27}Co$

14. ${}^{14}_{6}C$ is NOT used for
    a. tracing carbon atoms in reactions
    b. dating archaeological artifacts
    c. understanding the chemistry of photosynthesis
    d. sterilizing medical equipment
    e. studying climate change

15. For an element with 3 isotopes, of which 70% have a mass of 40, 20% have a mass of 35, and 10% have a mass of 38, the weighted average atomic mass is
    a. 36.3 amu
    b. 37.7 amu
    c. 39.1 amu
    d. 36.1 amu
    e. 38.8 amu

SCIENCE

FOCUSED QUIZ 04

ELECTRON ENERGY LEVELS –
THE WAVE AND PARTICLE NATURE, PP. 11-12

1. Why might a chemist strongly heat an atom?
   a. to slow down the electrons
   b. to excite the electrons
   c. to discover the band structure of the electrons
   d. to repulse the electrons
   e. to investigate the electrons' spins

2. A quantum of light energy is called a
   a. magnon
   b. boson
   c. plasmon
   d. photon
   e. phonon

3. An absorption spectrum is BEST defined as a spectrum produced when
   a. trace gases absorb infrared light through an air sample
   b. atoms absorb specific wavelengths of light
   c. optical fibers absorb electromagnetic waves due to trapped ions
   d. water absorbs electromagnetic radiation
   e. a solution absorbs a quantity of light

4. In the Bohr model, electrons occupy orbits that are
   a. fluctuating
   b. hybridized
   c. mobile
   d. undeterminable
   e. stationary

5. Which of the following entities is ALWAYS positively charged?
   a. nucleus
   b. neutrino
   c. electron
   d. photon
   e. ion

6. In the Bohr model, the nucleus-electron pair has potential energy due to the
   a. displacement between each shell of electrons
   b. spins of each electron
   c. distance between the electron orbit and the nucleus
   d. motion of subatomic particles within the atom
   e. differing charges of neutrons and electrons

7. When an electron absorbs light, it
   a. emits a photon
   b. approaches the nucleus
   c. transfers to another atom
   d. moves further from the nucleus
   e. attracts an electron from a nearby atom

8. At maximum, how many electrons occupy the exact same orbit?
   a. 3
   b. 8
   c. 4
   d. 2
   e. 6

9. Which observation provides the basis for the quantum mechanical model?
   a. The nucleus comprises most of an atom's mass.
   b. Atoms are divisible into subatomic particles.
   c. Subatomic particles have electrical charges.
   d. The positive charge of an atom is concentrated in one area.
   e. Matter has both mass and wave properties.

10. Which model of the atom places electrons at exact locations around the nucleus?
    a. the Bohr model of the atom
    b. the solid sphere model of the atom
    c. the nuclear model of the atom
    d. the quantum mechanical model of the atom
    e. the plum pudding model of the atom

11. Quantum physics discusses
    a. matter's wave-particle duality
    b. physical interaction between electrically charged particles
    c. behavior of electromagnetic phenomena
    d. interactions of atomic nuclei
    e. atoms as an isolated system of electrons and a nucleus

12. How does the quantum mechanical model of the atom differ from the Bohr model of the atom?
    a. The quantum mechanical model does not explain the shielding behavior of electrons in heavier atoms.
    b. The quantum mechanical model does not acknowledge that electrons have negligible mass.
    c. The quantum mechanical model does not illustrate the path of electrons around the nucleus.
    d. The quantum mechanical model does not limit an electron's energy to certain values.
    e. The quantum mechanical model does not recognize electrons' wave-like behavior.

13. Which effect MOST demonstrated that light acts as a particle?
    a. the photomagnetic effect
    b. the photovoltaic effect
    c. the photoelectrochemical effect
    d. the photoconductive effect
    e. the photoelectric effect

14. Which famous experiment displayed the wave properties of the electron?
    a. the Mach–Zehnder experiment
    b. the Davisson-Germer experiment
    c. the Elitzur–Vaidman experiment
    d. the Franck–Hertz experiment
    e. the Stern-Gerlach experiment

15. The crystal used in the classic experiment that demonstrated the wave properties of the electron was made of
    a. niobium
    b. nitrogen
    c. nobelium
    d. neodymium
    e. nickel

1. A location in an atom's orbital where electrons cannot exist is called a
   a. cloud
   b. isotope
   c. group
   d. period
   e. node

2. Atomic orbitals are differentiated by their shapes and given names such as s and
   a. p
   b. a
   c. g
   d. e
   e. b

3. In the periodic table, elements are organized into periods and
   a. groups
   b. nodes
   c. phases
   d. clusters
   e. orbitals

4. How many groups does the periodic table contain?
   a. 18
   b. 8
   c. 7
   d. 12
   e. 21

5. Which of the following characteristics, when similar, BEST indicates that two elements have similar properties?
   a. number of neutrons
   b. atomic mass
   c. group number
   d. number of protons
   e. period number

6. A higher number of nodes corresponds to
   a. more neutrons
   b. more electrons
   c. a larger atomic number
   d. a smaller orbital
   e. higher energy

7. The characteristics of a particular atom's electron orbitals indicates the
   a. atom's reactivity with itself
   b. distance of each electron from the nucleus
   c. types of chemical bonding that the atom can perform
   d. size of the atom
   e. position of the atom in relation to other atoms

8. Which of the following statements about main groups is NOT true?
   a. There are 8 main groups .
   b. Elements that are in the same main group share similar properties.
   c. Elements in main groups are more commonly seen in nature.
   d. Main groups are labelled with an "A".
   e. Elements in main groups have the same number of valence electrons as their group number.

9. Which of the following statements about orbitals is NOT true?
   a. Orbitals contain both electrons and neutrons.
   b. Orbitals exist because electrons exhibit wave-like properties.
   c. Orbitals indicate the locations of electrons within an atom.
   d. Orbitals are differentiated by their shape.
   e. Orbitals determines the outcomes of an atoms' chemical reactions with other atoms.

10. The modern periodic table is NOT ordered by
    a. the structures of elements' electronic orbitals
    b. elements' physical and chemical properties
    c. the number of neutrons in each element
    d. the number of protons in an element
    e. the elements' atomic number

1. Periodic trends come MOST directly from variations in
   a. kinetic energy
   b. mass
   c. attractive forces
   d. potential energy
   e. size

2. Which of the followings properties DECREASES from left to right across a period?
   a. electron affinity
   b. melting point
   c. electronegativity
   d. atomic radius
   e. ionization energy

3. Carbon's atomic number is
   a. 7
   b. 17
   c. 12
   d. 3
   e. 6

4. Ionization energy is BEST defined as the energy
   a. contained by in the crystal lattice of a compound
   b. needed to break apart a mole of molecules into the atoms that comprise it
   c. required to break a bond and form two molecular fragments
   d. change resulting from the addition of an electron to an atom
   e. required to remove a valence electron from an atom

5. Why does the force pulling on an atom's electrons increase from left to right across a period?
   a. Across a period, elements are more likely to attract electrons, which increases the force pulling on electrons.
   b. Across a period, increasing neutrons increases mass, causing the force on electrons to increase.
   c. Across a period, more protons are added, creating more attractive force pulling on electrons.
   d. Across a period, atomic radius decreases, so potential energy increases, increasing force.
   e. Across a period, repulsion between electrons increases, increasing the nuclear force pulling on electrons.

6. A negative ion is formed when a(n)
   a. neutron is removed from an atom
   b. proton is added to an atom
   c. electron is added to an atom
   d. proton is subtracted from an atom
   e. electron is subtracted from an atm

7. Which of the following elements provides an exception to the general trend of electron affinity?
   a. carbon
   b. hydrogen
   c. fluorine
   d. nitrogen
   e. nickel

8. An added orbital at a greater distance from the nucleus is MOST likely to be called a
   a. valence
   b. ring
   c. set
   d. shield
   e. shell

9. Atomic radius increases down a group on the periodic table because of
   a. added protons
   b. effective nuclear charge
   c. increasing distance
   d. more hybridized orbitals
   e. larger mass

10. As distance increases, attraction
    a. decreases, then increases
    b. increases, then decreases
    c. increases
    d. decreases
    e. decreases, then plateaus

11. Which of the following scientists is MOST connected to the concept of electronegativity?
    a. Linus Pauling
    b. John Dalton
    c. Henry Cavendish
    d. Joseph Louis Proust
    e. Antoine Lavoisier

12. Which of the following elements is the MOST electronegative?
    a. carbon
    b. oxygen
    c. fluorine
    d. sodium
    e. hydrogen

13. An ionic bond is MOST likely to form if the difference between two atoms' electronegativity is
    a. 2
    b. 0.8
    c. 1.6
    d. 0.5
    e. 0

14. Which of the following compounds MOST likely experiences dipole moments?
    a. H_2
    b. CCl_4
    c. F_2
    d. CO
    e. NaF

15. Which of the following compounds will have an electronegativity value of 0?
    a. CO
    b. CH_4
    c. H_2
    d. NH_3
    e. NaF

1. Which of the following definitions BEST describes a molecule?
   a. an attractive force between two atoms that requires force to break
   b. a group of atoms held in a constant ratio by strong covalent bonds
   c. a group of two or more substances that are combined but not chemically bonded
   d. a group of atoms of at least 2 different elements held together in a fixed ratio
   e. an extended lattice of covalent bonds

2. Which of the following compounds exists in the form of a network?
   a. magnesium oxide
   b. silica
   c. sodium bromide
   d. lithium fluoride
   e. table salt

3. Which of the following pairs INCORRECTLY matches the compound with its formula?
   a. Ammonia | NH_3
   b. Methane | CH_4
   c. Hydrogen | H_2
   d. Table salt | NaCl
   e. Silica | SiO_4

4. The structure of salt is BEST described as a
   a. network
   b. hexagonal crystal
   c. matrix
   d. lattice
   e. tetragonal crystal

5. Which type of attraction creates the physical properties of a substance?
   a. ionic bonds
   b. intermolecular forces
   c. nonpolar covalent bonds
   d. polar covalent bonds
   e. metallic bonds

6. The boiling point of water is
   a. 0 degrees Celsius
   b. 100 degrees Celsius
   c. 373 degrees Celsius
   d. 212 degrees Celsius
   e. 32 degrees Celsius

7. All attractive forces regarding atoms and molecules stem from the
   a. electromechanical force
   b. electromagnetic force
   c. electromotive force
   d. electrostatic force
   e. electrochemical force

8. How do ionic and covalent bonds differ?
   a. Ionic bonds occur when exactly one electron is transferred from one atom to another, whereas covalent bonds occur when more than one electron is transferred from one atom to another atom.
   b. Ionic bonds occur when two atoms share an electron pair, whereas covalent bonds occur when one or more electron is transferred from one of the atoms to another.
   c. Ionic bonds occur when one or more electron is transferred from one atom to another, whereas covalent bonds occur when two atoms share a pair of electrons.
   d. Ionic bonds occur between atoms within a molecule, whereas covalent bonds occur between discrete molecules.
   e. Ionic bonds occur when one or more electron is transferred from one atom to another, whereas covalent bonds occur when electrons can move freely among the atoms.

9. The term “electron sea” is MOST associated with
   a. ionic bonds
   b. polar covalent bonds
   c. hydrogen bonds
   d. metallic bonds
   e. nonpolar covalent bonds

10. What causes electricity?
    a. negative ions
    b. moving electrons
    c. magnetic dipoles
    d. electronegativity
    e. transferred electrons

11. When multiple elements form a metal, the result is called an
    a. allotrope
    b. alloy
    c. electrum
    d. admixture
    e. isomer

12. Brass is composed of copper and
    a. carbon
    b. iron
    c. silicon
    d. zinc
    e. tin

13. In an ionic bond, the basis of attraction is
    a. ion charge-polarizable electron cloud
    b. cations-decentralized electrons
    c. ion charge-dipole charge
    d. nuclei-shared electron pair
    e. cation-anion

14. The energy of a covalent bond MOST exactly occurs in a range between
    a. 40 kJ/mol and 600 kJ/mol
    b. 150 kJ/mol and 1100 kJ/mol
    c. 75 kJ/mol and 1000 kJ/mol
    d. 2 kJ/mol and 10 kJ/mol
    e. 400 kJ/mol and 4000 kJ/mol

15. Which of the following elements is NOT a metal?
    a. Br
    b. Zn
    c. Na
    d. Au
    e. Fe

1. Which property causes bonds between identical atoms to be nonpolar?
   a. radioactivity
   b. velocity
   c. symmetry
   d. distance
   e. energy

2. Which of the following symbols denotes a partial charge?
   a. δ
   b. μ
   c. α
   d. σ
   e. π

3. Which of the following forces is STRONGEST?
   a. covalent bonds
   b. London dispersion forces
   c. hydrogen bonds
   d. ion-dipole forces
   e. dipole-dipole forces

4. Which type of force is NOT do physical changes NOT affect?
   a. dipole-dipole bonds
   b. ionic bonds
   c. London dispersion forces
   d. van der Waals forces
   e. ion-dipole forces

5. Which of the following forces are based on multiple polarizable electron clouds?
   a. dipole-dipole forces
   b. dipole-induced dipole forces
   c. London dispersion forces
   d. ion-dipole forces
   e. ion-induced dipole forces

6. Dipole-induced dipole force occur between
   a. Fe^{2+} and O_2
   b. Na^+ and H_2O
   c. ICl and ICl
   d. CH_2O and CO
   e. HCl and Cl_2

7. Which of the following forces are the only types of forces that occur between two nonpolar molecules?
   a. hydrogen bonds
   b. ion-dipole forces
   c. ion-induced dipole forces
   d. dipole-dipole forces
   e. London dispersion forces

8. Which of the following atoms does NOT participate in hydrogen bonding?
   a. F
   b. N
   c. H
   d. C
   e. O

9. Hydrogen bonds are a type of strong
   a. London dispersion force
   b. ion-dipole interaction
   c. dipole-dipole interaction
   d. dipole-induced dipole force
   e. ion-induced dipole force

10. In liquid and solid water, the hydrogen bond is comprised of a weak bond between a hydrogen atom and the
    a. weakly electronegative hydrogen atom on the same water molecule
    b. weakly electronegative oxygen atom on a neighboring water molecule
    c. strongly electronegative hydrogen atom on a neighboring water molecule
    d. strongly electronegative oxygen atom on a neighboring water molecule
    e. strongly electronegative oxygen atom on the same water molecule

11. Ice's structure is BEST described as
    a. flexible
    b. tetrahedral
    c. open
    d. mobile
    e. narrow

12. Which of the following bases does NOT make up DNA?
    a. uracil
    b. adenine
    c. guanine
    d. thymine
    e. cytosine

13. DNA's backbone comprises a sugar and
    a. sulfate groups
    b. chlorate groups
    c. carbonate groups
    d. phosphate groups
    e. nitrate groups

14. In the 1950s, scientists established a model of DNA based on the X-ray studies of
    a. Johannes van der Waals
    b. James Watson
    c. Fritz London
    d. Rosalind Franklin
    e. Francis Crick

15. Why does DNA form a double helix structure?
    a. anticodon translation
    b. base pairing
    c. methylation
    d. hydrogen bonding
    e. isomerism

1. Which of the following elements is MOST abundant in our universe?
   a. sodium
   b. carbon
   c. oxygen
   d. hydrogen
   e. nitrogen

2. Who was the FIRST chemist to identify hydrogen as an element distinct from other gases and to describe its properties?
   a. Antoine Lavoisier
   b. Henry Cavendish
   c. Robert Boyle
   d. Paracelsus
   e. Gilbert Lewis

3. Many observations of hydrogen's properties came from reactions between metal and a(n)
   a. actinide
   b. salt
   c. base
   d. hydrocarbon
   e. acid

4. Who gave hydrogen its current name?
   a. Paracelsus
   b. Robert Boyle
   c. Henry Cavendish
   d. Dmitri Mendeleev
   e. Antoine Lavoisier

5. Why do ionic lattice networks form?
   a. Most ionic substances exist as solids at room temperature, decreasing distance between ions, allowing them to form a lattice.
   b. Larger anions alternate with smaller cations, reducing distance and increasing attraction strength.
   c. Ions have a sea of electrons which attract more cations, creating a lattice network.
   d. Ions create charges that extend in all directions, attracting multiple neighboring ions.
   e. Atoms can share electrons with multiple other atoms.

6. Which of the following substances exists as a covalent network solid?
   a. caffeine
   b. brass
   c. methane
   d. table salt
   e. diamond

7. To how many other atoms is each carbon atom in diamond joined?
   a. 1
   b. 2
   c. 4
   d. 3
   e. 5

8. Why can metals be bent, hammered into sheets, or pulled into wires?
   a. Metals exhibit translational motion, so constituent atoms can move rather freely.
   b. Metallic bonds have layers of atoms that can slide over each other easily.
   c. Metals have weak intermolecular forces, so the molecules are easier to separate.
   d. Metals experience a clustering effect, so they can flex without breaking.
   e. Metallic bonds require less energy than other intramolecular forces to change physically.

9. Which type of motion MOST determines a substance's state of matter?
   a. translational motion
   b. vibrational motion
   c. rotational motion
   d. oscillational motion
   e. transformational motion

10. Which change MUST take place for a solid to melt?
    a. broken intramolecular bonds
    b. increased electronegativity
    c. added energy
    d. decreased velocity
    e. reduced dipoles

11. A substance's temperature is a measure of its
    a. molecular acceleration
    b. molecular motion
    c. current
    d. potential energy
    e. wavelength

12. Which of the following substances is MOST likely to be a liquid at room temperature?
    a. NaCl
    b. SiO_2
    c. CH_4
    d. H_2O
    e. CO_2

13. If a compound has only dispersion forces,
    a. its electron clouds are not polarizable
    b. it exists in a network structure
    c. it never has induced dipoles
    d. it has no permanent dipoles
    e. it has a high bond energy

14. When a substances changes from a solid directly to a gas, it has
    a. recombined
    b. deposited
    c. evaporated
    d. condensed
    e. sublimated

15. Which of the following statements about water's bonding is FALSE?
    a. Its O-H bonds are strong covalent bonds.
    b. Each water molecule can form up to four hydrogen bonds.
    c. It has 2 O-H bonds per molecule.
    d. Hydrogen bonds cause adjacent water molecules to adjoin.
    e. Hydrogen bonds constitute its intramolecular bonds.

1. Why are some incorrect scientific models retained?
    a. There is not enough evidence to offer strong alternative hypotheses.
    b. They are too prominent to cease using them.
    c. Not all scientists agree on the models' inaccuracy.
    d. They may still be useful in making predictions.
    e. They are less time consuming to create.

2. In which institution did Gilbert Lewis work?
    a. the University of Chicago
    b. the University of Pennsylvania
    c. the University of California
    d. the University of Michigan
    e. the University of Wisconsin

3. How do modern understandings of chemistry conflict with Gilbert Lewis's model of compounds?
    a. Scientists today know that electrons have negative charges.
    b. Scientists today do not visualize electrons as dots.
    c. Scientists today reject the idea that electrons exist in discrete locations.
    d. Scientist today disagree that electrons are uniformly spread out.
    e. Scientists today dispute the presumption that electrons are stationary.

4. Valence electrons are the
    a. additional transferred electrons in an anion
    b. electrons that always form bonds
    c. electrons that have the potential to form bonds
    d. innermost electrons in an atom
    e. electrons that never form bonds

5. How many chlorine atoms bond with a carbon atom in one molecule of carbon tetrachloride?
    a. 3
    b. 1
    c. 4
    d. 5
    e. 2

6. In Lewis structures, a dash represents a(n)
    a. neutron
    b. nucleus
    c. orbital
    d. proton
    e. bond

7. Which of the following statements regarding bonding electrons is TRUE?
    a. There are usually fewer electrons involved in bonding than there are in the whole atom.
    b. The number of bonding electrons in an atom is equivalent to the number of neutrons.
    c. The ratio of bonding electrons to total electrons varies with ionization energy.
    d. There are usually equally many electrons involved in bonding than there are in the whole atom.
    e. There are usually more electrons involved in bonding than there are in the whole atom.

8. Electron orbitals represent maps of
    a. electron quantity
    b. electron quality
    c. electron conductivity
    d. electron density
    e. electron velocity

9. Which of the following forces occurs when two orbitals overlap?
    a. a covalent bond
    b. an ionic bond
    c. a double covalent bond
    d. a dispersion force
    e. an on-dipole force

10. Electron pairs that are NOT involved in bonding are known as
    a. lone pairs
    b. shared pairs
    c. valance pairs
    d. antibonding pairs
    e. inert pairs

11. How many electrons are transferred during the bonding of NaCl?
    a. 1
    b. 7
    c. 5
    d. 2
    e. 6

12. How many outer shell electrons does one chlorine atom have?
    a. 7
    b. 8
    c. 2
    d. 4
    e. 1

13. The original Lewis structure model did NOT consider the existence of
    a. multiple bonds
    b. orbitals
    c. nonbonding pairs
    d. octets
    e. oxidation states

14. How many electrons compose the bonds in CO_2?
    a. 4
    b. 16
    c. 6
    d. 2
    e. 8

15. Each molecule of nitrogen contains
    a. one double bond
    b. two single bonds
    c. one triple bond
    d. one single bond
    e. three single bonds

1. Chemists developed the concept of hybridization when the interaction of atomic orbitals did not always predict the correct molecular
   a. resonance
   b. shape
   c. energy
   d. spectra
   e. radius

2. In the context of atomic orbitals, hybridization means that two or more
   a. metals combine to form a unique metal with desirable qualities
   b. electron orbitals combine to form new orbitals with a different shape
   c. molecules form a different and unique compound
   d. subatomic particles combine to form a new subatomic particle with different properties
   e. electron pairs combine to form double or triple bonds

3. One s and three p orbitals combine to form
   a. four s^3p orbitals
   b. three sp^4 orbitals
   c. one sp^3 orbital
   d. four s^3p^3 orbitals
   e. four sp^3 orbitals

4. The positions of bonding electrons in bonds are modeled as
   a. seas
   b. ellipses
   c. waves
   d. lines
   e. ripples

5. The formation of two sp orbitals results from the combination of
   a. 2 s orbitals and 1 p orbital
   b. 1 s orbital and 1 2p orbital
   c. 2 s orbitals and 2 p orbitals
   d. 1 s orbital and 1 p orbital
   e. 1 s orbital and 2 p orbitals

6. In chemistry, MO stands for
   a. multiple orbitals
   b. model orbitals
   c. mapped orbitals
   d. molecular orbitals
   e. mixed orbitals

7. How do sigma and pi bonds differ?
   a. In a sigma bond, electrons are concentrated along an imaginary axis in the x-direction, whereas in pi bonds, electrons are concentrated along an imaginary axis in the y-direction.
   b. In a sigma bond, electrons are concentrated along an imaginary axis connecting the atoms, whereas in pi bonds, electrons are concentrated between the atoms but away from the center line.
   c. In a sigma bond, electrons are concentrated along an imaginary axis in the x-direction, whereas in pi bonds, electrons are concentrated along an imaginary axis in both the x and y-directions.
   d. In a sigma bond, electrons are concentrated along an imaginary axis in the y-direction, whereas in pi bonds, electrons are concentrated along an imaginary axis in the x-direction.
   e. In a sigma bond, electrons are concentrated between the atoms but away from the center line, whereas in pi bonds, electrons are concentrated along an imaginary axis connecting the atoms.

8. A sigma bond forms from
   a. 1 s orbital and 1 p orbital
   b. 2 2p orbitals
   c. 2 s orbitals and 1 p orbital
   d. 1 s orbital and 2 p orbitals
   e. 2 s orbitals

9. In a pi bond, how must the atomic orbitals align?
   a. in the y-coordinate direction
   b. in the direction of the line y = -x
   c. in the x-coordinate direction
   d. in the direction of the line y = x
   e. in the z-coordinate direction

10. How many valence electrons does one molecule of F_2 have?
   a. 10
   b. 16
   c. 14
   d. 8
   e. 12

11. O_2 forms
   a. one sigma bond
   b. one pi bond and two sigma bonds
   c. one pi bond
   d. one sigma bond and two pi bonds
   e. one sigma bond and one pi bond

12. How many electrons are shared in 1 molecule of N_2?
   a. 4
   b. 6
   c. 2
   d. 1
   e. 3

13. How many electron pairs exist between the carbon atom and 1 oxygen atom in CO_2?
   a. 8
   b. 1
   c. 2
   d. 3
   e. 4

14. How many total valence electrons does 1 molecule of CO_2 have?
   a. 14
   b. 16
   c. 10
   d. 12
   e. 8

15. In addition to non-bonding and bonding, orbitals can be
   a. trans-bonding
   b. un-bonding
   c. anti-bonding
   d. semi-bonding
   e. reverse-bonding

1. Oxidation states are values that signify a change in
   a. electron affinities
   b. ionization energies
   c. electronegativities
   d. electrons
   e. ionic potentials

2. If an atom has 3 electrons and loses 1, its oxidation state is
   a. +1
   b. -1
   c. +4
   d. +2
   e. -2

3. The oxidation number of an oxygen atom in H_2O is
   a. -2
   b. -1
   c. +1
   d. +2
   e. -3

4. If the total oxidation state for all atoms in a molecule sums to zero, the molecule MUST be
   a. charged
   b. hybridized
   c. resonant
   d. neutral
   e. linear

5. The oxidation number of N in NH_4^+ is
   a. +4
   b. -4
   c. +3
   d. -3
   e. +1

6. Each H in NH_4^+ has an oxidation number of
   a. -4
   b. -1
   c. -3
   d. +1
   e. +4

7. If in PO_4^{3-}, each O atom has an oxidation number of -2, the oxidation number of P must be
   a. +4
   b. -3
   c. +8
   d. +5
   e. -5

8. What is the oxidation number of S in H_2S if hydrogen's oxidation number is +1?
   a. +1
   b. -1
   c. -3
   d. +2
   e. -2

9. The VSEPR model aims to predict
   a. orbital hybridization
   b. electron geometry
   c. molecular geometry
   d. oxidation number
   e. atomic radius

10. The VSEPR model depends upon
    a. inert-pair effects
    b. electron pair repulsion
    c. relative atom size
    d. resonance structure
    e. lone pair location

11. The MAIN difference between the VSEPR model and Lewis structures is that the VSEPR model
    a. represents molecules in two dimensions rather than three
    b. considers only lone pairs and not bonding electrons
    c. considers lone pairs in addition to bonding electrons
    d. represents molecules in three dimensions rather than two
    e. considers only bonding electrons and not lone pairs

12. Which of the following molecules has tetrahedral geometry?
    a. BF_3
    b. PF_5
    c. ClF_4^-
    d. SF_6
    e. CH_4

13. Which of the following molecules has linear geometry?
    a. H_2O
    b. SF_6
    c. BF_3
    d. BeF_2
    e. PF_5

14. How many nonbonding electrons does BF_3 have?
    a. 12
    b. 24
    c. 16
    d. 6
    e. 18

15. How many total electrons comprise the bonds of SF_6?
    a. 18
    b. 10
    c. 36
    d. 30
    e. 12

1. The resonance model was introduced due to experimental findings that bond properties were equal even in the presence of
   a. delocalized electrons
   b. multiple bond types
   c. irregular molecular geometry
   d. insufficient valence electrons
   e. London dispersion forces

2. How many oxygen atoms make up one ozone molecule?
   a. 5
   b. 3
   c. 6
   d. 4
   e. 2

3. How many valence electrons does one molecule of ozone have?
   a. 14
   b. 18
   c. 16
   d. 10
   e. 12

4. How many resonance structures does ozone have?
   a. 3
   b. 1
   c. 5
   d. 2
   e. 4

5. The term "resonance" came from the idea that
   a. ionic compounds can form lattices
   b. orbitals combine to form new hybrid orbitals
   c. electrons can delocalize
   d. compounds can have multiple isomers
   e. a molecule oscillates between two models

6. The bond total between an atom on sulfur and an atom of oxygen in SO_3 is
   a. 1 1/2
   b. 2
   c. 1 1/3
   d. 1 1/4
   e. 3

7. Why was the resonance model replaced?
   a. It does not explain the presence of nonbonding electrons.
   b. It does not explain all the known types of molecular geometries.
   c. It does not explain molecules with non-integer bond values.
   d. It does not explain sigma and pi bond formation and "delocalized" orbitals.
   e. It does not explain equal bond properties in cases of incomplete octets.

8. How many bonds does each side of an ozone molecule contain?
   a. 2
   b. 1 1/2
   c. 1 1/4
   d. 1 1/3
   e. 1

9. How many nonbonding valence electrons are in one molecule of SO_3?
   a. 12
   b. 18
   c. 16
   d. 14
   e. 10

10. Which theory has replaced the resonance model?
   a. Lewis theory
   b. molecular orbital theory
   c. spin-coupled theory
   d. VSEPR theory
   e. valence bond theory

11. A molecule whose bond dipoles do not cancel will MOST likely have
   a. hybridized orbitals
   b. multiple resonance structures
   c. a significant dipole moment
   d. irregular geometry
   e. symmetry

12. Which type of geometry do methane molecules have?
    a. trigonal planer
    b. linear
    c. square planar
    d. tetrahedral
    e. trigonal pyramidal

13. The bond angle in a molecule of methane is MOST nearly
    a. 90°
    b. 107°
    c. 105°
    d. 109.5°
    e. 120°

14. The structure of a CO_2 molecule is
    a. tetrahedral
    b. bent
    c. linear
    d. t-shaped
    e. trigonal planar

15. Which molecule will have the LOWEST melting and boiling points?
    a. HF, a polar covalent molecule
    b. CCl_4, a nonpolar covalent molecule
    c. N_2, a diatomic molecule
    d. Cu, a metal
    e. NaCl, an ionic compound

1. Which of the following equations BEST represents an atom's mass number?
   a. $A = N/Z$
   b. $A = A_Z + A_N$
   c. $A = f_1M_1 + f_2M_2 + \ldots + f_nM_n$
   d. $A = Z \times N$
   e. $A = Z + N$

2. Which of the following changes occurs during alpha decay?
   a. A beta minus particle is emitted.
   b. A hydrogen nucleus is lost.
   c. A proton transforms into a neutron.
   d. A helium nucleus is lost.
   e. A neutron transforms into a proton.

3. An electron may also be known as a(n)
   a. beta minus particle
   b. positron
   c. beta plus particle
   d. antineutrino
   e. alpha particle

4. Radioactive decay occurs due to unstable
   a. nuclides
   b. neutrons
   c. antineutron
   d. neutrinos
   e. nuclei

5. Alpha decay is MOST likely to occur if a nucleus is larger than
   a. radium
   b. mercury
   c. xenon
   d. bismuth
   e. ruthenium

6. How does a neutrino differ from an electron?
   a. Neutrinos have positive charge and larger mass.
   b. Neutrinos have negative charge and larger mass.
   c. Neutrinos have no charge and very small mass.
   d. Neutrinos have positive charge and very small mass.
   e. Neutrinos have no charge and larger mass.

7. Plutonium-240 undergoes alpha decay to form
   a. U-235
   b. U-236
   c. Ba-141
   d. Th-234
   e. Pu-239

8. In beta decay, if there are too few neutrons,
   a. electrons are emitted
   b. neutrons are attracted
   c. neutrinos are emitted
   d. positrons are emitted
   e. protons are attracted

9. ^{22}Na undergoes beta decay to form
   a. ^{22}Ne
   b. ^{14}C
   c. ^{21}F
   d. ^{14}N
   e. ^{10}C

10. The symbol v represents a(n)
    a. antineutrino
    b. beta particle
    c. neutrino
    d. positron
    e. alpha particle

11. During beta-plus decay, carbon-10 forms
    a. fluorine-10
    b. oxygen-10
    c. beryllium-10
    d. nitrogen-10
    e. boron-10

12. Who showed that the bombardment of nitrogen atoms with alpha particles created oxygen atoms?

    a. Henry Cavendish
    b. Antoine Lavoisier
    c. Ernest Rutherford
    d. Linus Pauling
    e. John Dalton

13. During the fission of uranium, it is hit by a free

    a. positron
    b. proton
    c. neutron
    d. electron
    e. neutrino

14. In nuclear fission reactions, energy production occurs as a result of the

    a. release of neutrons
    b. heat used to drive the reaction
    c. release of electrons
    d. production of a reactive nuclide
    e. products flying apart

15. Why is large scale fusion for electric power NOT feasible?

    a. The equipment necessary for fusion is prohibitively expensive.
    b. Large scale fusion requires exceedingly large facilities.
    c. Too much energy is needed to overcome the repulsion of two positive forces.
    d. The necessary isotopes rarely occur naturally.
    e. Large scale fusion produces hazardous byproducts.

1. Elements are distinguished by their number of
   a. protons
   b. neutrinos
   c. neutrons
   d. electrons
   e. positrons

2. Atomic masses are defined relative to the mass of an isotope of
   a. helium
   b. hydrogen
   c. calcium
   d. oxygen
   e. carbon

3. Who formulated the law of partial pressures?
   a. Antoine Lavoisier
   b. Henry Cavendish
   c. John Dalton
   d. Joseph Louis Proust
   e. Linus Pauling

4. Which equation represents the law of partial pressures?
   a. $P_{Total} = P_A + P_B$
   b. $PV = nRT$
   c. $P_a = \frac{N}{m^2}$
   d. $V_1P_1 = V_2P_2$
   e. $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

5. Which value orders elements in the periodic table?
   a. electronegativity
   b. atomic mass
   c. atomic number
   d. electron affinity
   e. atomic radius

6. Which of the following forces are NOT intermolecular?
   a. dipole-induced dipole forces
   b. ion-dipole forces
   c. metallic bonds
   d. London dispersion forces
   e. hydrogen bonds

7. Polar bonds occur due to differences in
   a. electronegativity
   b. atomic radius
   c. ionization potential
   d. ionization energy
   e. electron affinity

8. Which of the following properties are intermolecular forces MOST likely to influence?
   a. atomic radius
   b. electron affinity
   c. atomic mass
   d. physical state
   e. electronegativity

9. Hydrogen bonds are a specific type of
   a. ion-induced dipole forces
   b. London dispersion forces
   c. dipole-induced dipole forces
   d. dipole-dipole forces
   e. ion-dipole forces

10. Who FIRST described hydrogen's properties and identified it as an element distinct from other gases?
    a. Paracelsus
    b. Henry Cavendish
    c. Robert Boyle
    d. John Dalton
    e. Antoine Lavoisier

11. How many electrons comprise a covalent bond, according to Gilbert Lewis?
    a. 8
    b. 4
    c. 2
    d. 6
    e. 3

12. Molecular orbitals can be sigma bonds or
    a. gamma bonds
    b. psi bonds
    c. phi bonds
    d. delta bonds
    e. pi bonds

13. VSEPR theory is used to predict molecules'
    a. molar masses
    b. shapes
    c. melting points
    d. conductivities
    e. solubilities

14. A molecule's shape can be used to predict its
    a. stability
    b. solubility
    c. polarity
    d. conductivity
    e. electronegativity

15. As radioactive decay occurs, a nucleus
    a. excites
    b. expands
    c. dissolves
    d. stabilizes
    e. combusts

1. A combination of substances in which each substance remains distinct is a
   a. reagent
   b. mixture
   c. compound
   d. catalyst
   e. solution

2. A combination of substances that looks like one substance is a
   a. solution
   b. catalyst
   c. mixture
   d. reagent
   e. compound

3. Boyle's law relates
   a. solubility to temperature
   b. mass to diffusion
   c. volume to pressure
   d. electricity to temperature
   e. mass to temperature

4. Early scientists FIRST observed that gases
   a. groups of rapidly moving particles
   b. have a kinetic energy proportional to temperature
   c. undergo perfectly elastic collisions
   d. lost mass as they transitioned between states
   e. were highly regular in behavior

5. What is Boyle's law?
   a. $V/T = D$
   b. $V = kN$
   c. $PV = nRT$
   d. $PV = C$
   e. $d = M/V$

6. What is Charles's law?
   a. $PV = C$
   b. $V/T = D$
   c. $d = M/V$
   d. $PV = nRT$
   e. $V = kN$

7. Which Celsius temperature is 0 K?
   a. 100
   b. 0
   c. 43
   d. -273
   e. -540

8. If the volume of a gas in a cylinder with a movable piston has decreased, what has MOST likely happened?
   a. The pressure has decreased.
   b. The mass has increased.
   c. The pressure has increased.
   d. The mass has decreased.
   e. The temperature has decreased.

9. A sample of gas in a rigid container of volume 2.0 liters has a pressure of 4.0 atmospheres at a temperature of 300 K. What will be the pressure at 600 K?
   a. 4.0 atmospheres
   b. 2.0 atmospheres
   c. 16 atmospheres
   d. 8.0 atmospheres
   e. 6.0 atmospheres

10. French scientist Jacques Charles was interested in
    a. bicycles
    b. helicopters
    c. steam engines
    d. hot air balloons
    e. dirigibles

11. Which gas did Jacques Charles use to fuel his invention?
    a. nitrogen
    b. helium
    c. oxygen
    d. hydrogen
    e. neon

12. Charles's law relates
    a. electricity to temperature
    b. solubility to temperature
    c. volume to temperature
    d. volume to pressure
    e. mass to temperature

13. Which scientist FIRST published on Jacques Charles's findings?
    a. Joseph-Louis Gay-Lussac
    b. Antoine Lavoisier
    c. Robert Boyle
    d. John Dalton
    e. Amedeo Avogadro

14. 600.0 mL of air is at 273 K. What is the volume at 333 K, rounded to the unit place?
    a. 492 mL
    b. 873 mL
    c. 706 mL
    d. 732 mL
    e. 423 mL

15. What happened to the first draft of Jacques Charles's invention?
    a. Peasants destroyed it.
    b. It exploded.
    c. He preserved it in a museum.
    d. His servant threw it away.
    e. His children crashed it.

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. 35-6

1. Which law allows for the consideration of gases in a mixture separately?
   a. Gay-Lussac's Law
   b. Avogadro's Law
   c. Dalton's Law of Partial Pressure
   d. Boyle's Law
   e. Charles's Law

2. Particles that behave ideally
   a. exhibit repulsive forces
   b. are compressible
   c. have significant volume
   d. are all the same speed
   e. do not react with each other

3. "Mole fraction" is BEST represented by the expression
   a. $\frac{P_A}{P_A+P_B+P_C+\cdots}$
   b. $\frac{[C]^C[D]^D}{[A]^A[B]^B}$
   c. $\frac{n_A}{n_A+n_B+n_C+\cdots}$
   d. $\frac{(P_C)^C(P_D)^D}{(P_A)^A(P_B)^B}$
   e. $\frac{P_1V_1}{T_1}$

4. Which law originates the direct relationship of pressure to number of moles?
   a. the ideal gas law
   b. the real gas law
   c. the law of partial pressures
   d. the law of partial volumes
   e. the combined gas law

5. The value of water vapor pressure depends only upon
   a. enthalpy
   b. viscosity
   c. volume
   d. mass
   e. temperature

6. MmHg is a measure of
   a. force
   b. speed
   c. length
   d. density
   e. pressure

7. In chemistry, T has units of
   a. Kelvin
   b. Celsius
   c. minutes
   d. seconds
   e. Fahrenheit

8. Which of the following assumptions does KMT NOT include?
   a. Gases lose energy during collisions.
   b. The molecules are moving quickly, but not all at the same speed.
   c. Pressure results from collisions of gas molecules and the container's walls.
   d. The volume of the gas molecules is negligible.
   e. There are no attractive or repulsive forces between molecules.

9. The symbol "u" is defined as the average
   a. frequency of a molecule
   b. density of a molecule
   c. acceleration of a molecule
   d. speed of a molecule
   e. force of a molecule

10. Force on a container wall is a direct result of a molecule's change in
    a. velocity
    b. momentum
    c. inertia
    d. acceleration
    e. position

11. Which equation BEST represents the force of a molecule's collisions?
    a. Force = μN
    b. Force = ma
    c. Force = $\frac{\tau}{r_\perp}$
    d. Force = $\frac{\Delta x}{\Delta t}$
    e. Force = $\frac{m\mathrm{u}^2}{\ell}$

12. Out of N molecules, how many travel in each direction?
    a. 1/3 N
    b. 1/6 N
    c. 1/2 N
    d. 1 N
    e. 1/4 N

13. The expression $A\ell$ represents
    a. volume
    b. perimeter
    c. surface area
    d. weight
    e. area

14. One way of calculating pressure involves force and
    a. unit area
    b. radius
    c. mass
    d. acceleration
    e. velocity

15. Which equation explains Boyle's law?
    a. A = $\frac{n_A}{n_A+n_B+n_C+\cdots}$
    b. $P_1V_1 = P_2V_2$
    c. PV = nRT
    d. F = $\frac{m\mathrm{u}^2}{\ell}$
    e. PV = 1/3 $N m u^2$

FOCUSED QUIZ 18

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

1. Which number contributes MOST to a gas's change in pressure and volume when the amount of the gas is varied?
   a. the number of moles of the gas
   b. the ratio of molar mass of the gas and the atomic mass of the gas
   c. the ratio of the total moles of a gas to the mass of one gas molecule
   d. the mass of the gas
   e. the ratio of total mass of gas to mass of an individual gas molecule

2. Which of the following groups, at the same temperature and pressure, have equal volumes?
   a. 4g helium, 2g hydrogen, and 16g methane
   b. 16g helium, 2g hydrogen, and 4g methane
   c. 2g helium, 16g hydrogen, and 4g methane
   d. 4g helium, 16g hydrogen, and 2g methane
   e. 2g helium, 4g hydrogen, and 16g methane

3. Which quantity MOST determines the properties of a substance?
   a. entropy of molecules
   b. number of molecules
   c. mass of molecules
   d. heat capacity of molecules
   e. charge of molecules

4. A mole is a unit that measures a substance's
   a. weight
   b. absorbance
   c. amount
   d. concentration
   e. mass

5. Who FIRST determined Avogadro's number?
   a. Antoine Lavoisier
   b. Joseph-Louis Gay-Lussac
   c. John Dalton
   d. Jean Baptiste Perrin
   e. Amedeo Avogadro

6. When did Joseph-Louis Gay-Lussac make his famous discovery?
   a. 1856
   b. 1870
   c. 1909
   d. 1776
   e. 1805

7. Who explained that many gases are composed of two identical molecules bound together?
   a. John Dalton
   b. Amedeo Avogadro
   c. Jean Baptiste Perrin
   d. Joseph-Louis Gay-Lussac
   e. Jacques Charles

8. In chemistry, the letter "n" represents the
   a. number of molecules of a substance
   b. number of atoms of a substance
   c. number of isotopes of a substance
   d. number of moles of a substance
   e. number of ions of a substance

9. Which equation BEST represents Avogadro's law?
   a. $P_1V_1 = P_2V_2$
   b. $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
   c. $\frac{P_1}{T_1} = \frac{P_2}{T_2}$
   d. $\frac{V_1}{T_1} = \frac{V_2}{T_2}$
   e. $\frac{V_1}{n_1} = \frac{V_2}{n_2}$

10. In chemistry, standard temperature is considered to be
   a. 273 K
   b. 212 K
   c. 305 K
   d. 100 K
   e. 32 K

11. The MOST recent definition of a mole is based on an isotope of
    a. helium
    b. oxygen
    c. hydrogen
    d. nitrogen
    e. carbon

12. Avogadro's number is MOST nearly
    a. 1.602×10^{19}
    b. 2.998×10^{18}
    c. 6.022×10^{23}
    d. 6.626×10^{34}
    e. 2.244×10^{27}

13. What is the molar mass of helium?
    a. 12 g/mol
    b. 4 g/mol
    c. 14 g/mol
    d. 1 g/mol
    e. 16 g/mol

14. Which compound has a molar mass of 58 g/mol, given that the atomic mass of H is 1, the atomic mass of C is 12, and the atomic mass of O is 16?
    a. C_5H_{12}
    b. $C_2H_4O_2$
    c. C_3H_6O
    d. $C_4H_{10}O$
    e. C_2H_6O

15. What is the molar mass of table salt?
    a. 26.0 g/mol
    b. 49.5 g/mol
    c. 35.0 g/mol
    d. 58.5 g/mol
    e. 94.2 g/mol

FOCUSED QUIZ 19

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

1. "N" represents
    a. the universal molar gas constant
    b. Avogadro's number
    c. the number of moles
    d. molar mass
    e. the Boltzmann constant

2. Which expression represents the speed of molecules in one mole of gas?
    a. $\sqrt{\frac{3RT}{M}}$
    b. $\sqrt{\frac{3kT}{m}}$
    c. 1/3 Mu^2
    d. ½ mv^2
    e. PV/RT

3. The ideal gas law does NOT contain the
    a. universal molar gas constant
    b. Boltzmann constant
    c. pressure
    d. volume
    e. temperature

4. The Boyle's and Charles's law equations combine to form the equation
    a. $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
    b. PV = 1/3 Mu^2
    c. PV = nRT
    d. $u = \sqrt{\frac{3kT}{m}}$
    e. $P = \frac{nRT}{V-nb} - \frac{n^2a}{V^2}$

5. When did Carl Wilhelm Scheele publish his work on oxygen?
    a. 1772
    b. 1774
    c. 1770
    d. 1777
    e. 1775

6. Which of the following properties did Carl Wilhelm Scheele NOT note about oxygen gas?
    a. its combustion
    b. its high density
    c. its lack of odor
    d. its tastelessness
    e. its accommodation of respiration

7. How did Carl Wilhelm Scheele and Joseph Priestley experiment with oxygen?
    a. by burning oxides
    b. by maintaining weather records
    c. by dissolving metals in acid
    d. by observing ice and water at different temperatures
    e. by heating calcium carbonate

8. What did Joseph Priestley call the gas now known as oxygen?
    a. fixed air
    b. dephlogisticated air
    c. calx
    d. fire-air
    e. oxygéne

9. How many gases is Joseph Priestley credited with isolating?
    a. 5
    b. 4
    c. 8
    d. 6
    e. 7

10. Which of the following statements does NOT describe Antoine Lavoisier?
    a. He gave oxygen its current name.
    b. He claimed to have discovered oxygen independently.
    c. He debunked phlogiston theory.
    d. He published his work on oxygen in 1774.
    e. He understood combustion reactions.

11. Which principle did Antoine Lavoisier establish?
    a. the law of conservation of energy
    b. the law of partial pressures
    c. the law of conservation of mass
    d. the law of definite proportions
    e. the law of multiple proportions

12. Which term in the van der Waals equation corrects for volume of gas molecules?
    a. $\frac{n^2a}{V^2}$
    b. V
    c. $\sqrt{\frac{3kT}{m}}$
    d. nRT
    e. nb

13. Which of the following gases behaves MOST ideally?
    a. CH_4
    b. NH_3
    c. O_2
    d. CO_2
    e. He

14. What is the ideal value for PV/RT for one mole of gas?
    a. 1.5
    b. 2
    c. 0.5
    d. 1
    e. 0

15. Why are chlorofluorocarbons no longer ideal for use in refrigerators?
    a. They cause acidic precipitation.
    b. They do not liquify easily enough.
    c. They damage the ozone layer.
    d. They enter oceans and cause eutrophication.
    e. They do not cool rapidly enough.

SCIENCE

FOCUSED QUIZ 20

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

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1. The main difference between diffusion and effusion is the
    a. number of particles of the moving gas
    b. surroundings of the container
    c. type of gas involved
    d. size of the opening of the container
    e. temperature of the container

2. The dispersion of perfume is an example of
    a. adsorption
    b. affusion
    c. osmosis
    d. diffusion
    e. effusion

3. Why are elaborate computer models necessary to predict exact diffusion and effusion rates?
    a. Real gas molecules experience vibration.
    b. Real gas molecules experience friction.
    c. Real gas molecules experience repulsion.
    d. Real gas molecules experience attraction.
    e. Real gas molecules experience collisions.

4. The equation describes the relationship between which two values?
    a. the moles and mass of gas molecules
    b. the speed and temperature of gas molecules
    c. the speed and concentration of gas molecules
    d. the speed and mass of gas molecules
    e. the temperature and mass of gas molecules

5. By which factor does the mass of a gaseous helium atom differ from the mass of a gaseous hydrogen molecule?
    a. ¼
    b. 2
    c. 1
    d. ½
    e. 4

6. Which molecule will have the fastest diffusion and effusion rates?
    a. Cl_2, with a molar mass of about 70.9 g/mol.
    b. O_2, with a molar mass of about 32.0 g/mol.
    c. Ne, with a molar mass of about 20.2 g/mol.
    d. Ar, with a molar mass of about 39.9 g/mol.
    e. N_2, with a molar mass of about 28.0 g/mol.

7. Which chemical formula represents acetone?
    a. $C_2H_4O_2$
    b. CH_4
    c. NH_3
    d. C_8H_{18}
    e. C_3H_6O

8. What is the molar mass of ammonia?
    a. 14 g/mol
    b. 58 g/mol
    c. 36 g/mol
    d. 3 g/mol
    e. 17 g/mol

9. Which family on the periodic table has gas atoms rather than molecules?
    a. alkaline gases
    b. chalcogens
    c. halogens
    d. alkali gases
    e. noble gases

10. Which form of motion do liquid and solid substances NOT share?
    a. vibrational motion
    b. translational motion
    c. revolutional motion
    d. rotational motion
    e. dilational motion

11. Which change occurs during sublimation?
    a. A gas transforms directly into a solid.
    b. A solid transforms to a liquid.
    c. A gas transforms to a plasma.
    d. A solid transforms directly into a gas.
    e. A gas transforms to a liquid.

12. Which type of order do liquids have?
    a. long-range order
    b. full order
    c. short-range order
    d. no order
    e. medium-range order

13. A liquid's degree of ordering depends MOST upon
    a. molar mass
    b. volume
    c. temperature
    d. solubility
    e. pressure

14. Which liquid is a component of gasoline?
    a. acetylene
    b. methane
    c. acetone
    d. phenol
    e. octane

15. Which series correctly lists the state of matters in order of least to most dense?
    a. gas, solid, liquid
    b. solid, gas, liquid
    c. gas, liquid, solid
    d. solid, liquid, gas
    e. liquid, solid, gas

1. Which of the following structures are NOT types of solid structures?
   a. ionic lattice solids
   b. molecular solids
   c. simple cubic solids
   d. covalent network solids
   e. metallic solids

2. An allotrope is
   a. an atom of the same element with different number of electrons
   b. one of many possible forms of an element
   c. a molecule with the same chemical formula but different arrangements
   d. a mixture of two or more metals
   e. an alternative form of an element with a differing number of neutrons

3. What is the upper number of allotropes of carbon that are hypothesized to exist?
   a. 400
   b. 500
   c. 200
   d. 100
   e. 300

4. Which form of carbon is the MOST stable?
   a. diamond
   b. buckminsterfullerene
   c. graphite
   d. lonsdaleite
   e. fullerene

5. Which form of carbon is also known as C_{60}?
   a. graphite
   b. buckminsterfullerene
   c. fullerene
   d. lonsdaleite
   e. diamond

6. Which of the following forms of carbon is the WORST conductor?
   a. diamond
   b. single-walled carbon nanotube
   c. graphite
   d. C_{70}
   e. buckminsterfullerene

7. Which factor MOST affects a solid's arrangement?
   a. the energy of the particles
   b. the oxidation state of the particles
   c. the absorbance of the particles
   d. the mass of the particles
   e. the size of the particles

8. Which substance has the MOST closely packed structure?
   a. pure metals
   b. covalent solids
   c. ionic solids
   d. alloys
   e. molecular solids

9. In which structure are there eight nearest neighbors?
   a. cubic close-packed
   b. simple cubic
   c. body-centered cubic
   d. face-centered cubic
   e. close-packed hexagonal

10. Which metal has a body-centered cubic structure?
    a. tungsten
    b. cadmium
    c. aluminum
    d. zinc
    e. silver

11. Which factor MOST affects the softness of a metal?
    a. the enthalpy of the particles
    b. the melting point of the metal
    c. the reactivity of the atoms
    d. the size of the atoms
    e. the closeness of the packing

12. Metals are conductive because
    a. the outer electrons are further from the nuclei
    b. the outer electrons can move freely
    c. metals are extremely volatile and reactive
    d. metals are relatively insoluble
    e. metals have negative charges

13. Which property of metals does NOT result from metal's structure?
    a. their sonorous quality
    b. their ductile quality
    c. their lustrous quality
    d. their malleable quality
    e. their low melting point

14. Ductility is BEST defined as the ability to
    a. reflect light
    b. conduct electricity
    c. ring when struck
    d. be pulled into a wire
    e. be flattened into sheets

15. How do alloys increase metal strength?
    a. The impurities disrupt smooth layers and decrease their sliding.
    b. The addition of metals with more electrons increases the alloy strength.
    c. The alternation of differently sized metals creates smoother layers.
    d. The different charges increase the strength of the metallic bonds.
    e. The combination of metals causes the layers to pack more tightly.

1. Which quantity is on the vertical axis of a phase change diagram?
    a. volume
    b. mass
    c. pressure
    d. temperature
    e. heat capacity

2. The triple point represents the temperature and pressure at which
    a. the solid and liquid states exist in equilibrium
    b. the gas and liquid phases are no longer distinct
    c. a substance loses its magnetic properties
    d. a substance has the density of a liquid but the mobility of a gas
    e. solid, liquid, and gaseous states exist in equilibrium

3. Horizontal lines on a phase diagram represent the effects of changes in
    a. heat of vaporization
    b. temperature
    c. volume
    d. heat of formation
    e. pressure

4. At which temperature does water reach its critical point?
    a. 31.1°C
    b. 0°C
    c. 100°C
    d. 374°C
    e. 218°C

5. Why is the phase diagram for water unusual?
    a. The solid-liquid equilibrium line is not linear.
    b. the liquid-gas line is not linear.
    c. The liquid-gas line slopes backward.
    d. The solid-liquid equilibrium line slopes backward.
    e. The solid-liquid line slopes forward.

6. Which of the following adjectives BEST describes ice's structure?
    a. packed
    b. mobile
    c. flexible
    d. weak
    e. open

7. Which effect does an increase in pressure have on water's solid phase?
    a. lower melting point
    b. lower subliming point
    c. higher condensation point
    d. higher melting point
    e. lower boiling point

8. What is the LOWEST pressure at which CO_2 becomes liquid?
    a. 73 atm
    b. 56.4 atm
    c. 5.11 atm
    d. 218 atm
    e. 31.1 atm

9. A substance that is supercritical is BEST defined as behaving as
    a. both a solid and a liquid
    b. both a gas and a solid
    c. both a gas and a plasma
    d. a gas, a liquid, and a solid
    e. both a gas and a liquid

10. Which combination of properties is MOST ideal for a solvent?
    a. the density of a solid and the mobility of a liquid
    b. the density of a liquid and the mobility of a gas
    c. the density of a liquid and the mobility of a solid
    d. the mobility of a liquid and the density of a gas
    e. the density of a solid and the mobility of a gas

11. Why is supercritical CO_2 NOT feasible to use as a solvent?
    a. It must be transported from location to location prohibitively quickly.
    b. It is highly toxic and has negative health repercussions for humans.
    c. It is not recyclable and therefore adds to the carbon dioxide burden in the atmosphere.
    d. It must maintain a specific pressure and temperature to remain supercritical.
    e. It does not dissolve substances as well as non-supercritical CO_2.

12. What is one of the MAIN advantages of the use of supercritical CO_2 as a solvent?
    a. It is very reactive and acts as a solvent for many substances.
    b. It is easily removed by reducing the pressure.
    c. It is very light and inexpensive to transport.
    d. It is inexpensive and widely available.
    e. It is stable and does not react with its surroundings.

13. What did Joseph Black name carbon dioxide?
    a. fire air
    b. wild spirit
    c. dephlogisticated air
    d. fixed air
    e. phlogiston

14. Joseph Black discovered carbon dioxide by heating
    a. sodium bicarbonate
    b. barium carbonate
    c. cesium carbonate
    d. lithium carbonate
    e. calcium carbonate

15. When did Joseph Black present his findings?
    a. 1755
    b. 1745
    c. 1754
    d. 1799
    e. 1728

1. A solution must be
    a. diluted
    b. homogenous
    c. composite
    d. balanced
    e. activated

2. What distinguishes a solvent from a solute?
    a. It is present in a larger quantity.
    b. It is denser.
    c. It is at a higher temperature.
    d. It is an ionic salt.
    e. It is under more pressure.

3. The solvent in an aqueous solution is
    a. salt
    b. organic
    c. water
    d. nonpolar
    e. oil

4. What is the name of a solution in which the solvent is a nonpolar substance?
    a. homogenous solution
    b. molar solution
    c. colligative solution
    d. organic solution
    e. concentrated solution

5. The relative amount of solute and solvent present in a solution is the
    a. solubility
    b. concentration
    c. percentage
    d. pressure
    e. oxidation

6. A general rule of solubility is that
    a. cold dissolves better
    b. polar dissolves nonpolar
    c. like dissolves like
    d. pressure does not dissolve
    e. less dissolves more

7. Which of the following substances will dissolve BEST in water?
    a. benzene
    b. gasoline
    c. octane
    d. vegetable oil
    e. salt

8. An ionic salt will be insoluble if
    a. the positive and negative ions are separated
    b. the water molecules completely surround the ions
    c. the temperature of the solvent is raised
    d. the solvent is polar
    e. the lattice forces are especially strong

9. Which of the following sulfates is NOT soluble?
    a. barium sulfate
    b. sodium sulfate
    c. copper sulphate
    d. magnesium sulfate
    e. calcium sulfate

10. In which of the following groups are all common compounds soluble?
    a. Group V
    b. Group III
    c. Group I
    d. Group II
    e. Group IV

11. A solution containing exactly the maximum amount of a solute is
    a. heterogenous
    b. concentrated
    c. precipitated
    d. saturated
    e. colligated

12. An insoluble substance that forms in a solution is a(n)
    a. sediment
    b. silicone
    c. precipitate
    d. substance
    e. mass

13. A saturated solution is MOST likely to form solid crystals if it is

    a. agitated
    b. pressurized
    c. titrated
    d. cooled
    e. homogenous

14. Which insoluble compound forms stalactites and stalagmites?

    a. carbon dioxide
    b. silver chloride
    c. potassium iodide
    d. calcium carbonate
    e. hydroxide

15. A substance is mostly likely to dissolve easily in water if it contains

    a. O-H bonds
    b. F-N bonds
    c. N-O bonds
    d. F-O bonds
    e. H-N bonds

1. Intermolecular forces in nonpolar solvents are
   a. weak
   b. ideal
   c. metallic
   d. unidirectional
   e. diffuse

2. Carbon and hydrogen bonds have little polarity because they have similar
   a. electronegativity
   b. activation energy
   c. adhesion
   d. boiling points
   e. lattice structures

3. Nonpolar solvents are MOST particularly useful in
   a. hydrating
   b. emulsifying
   c. pressurizing
   d. oxidizing
   e. cleaning

4. What is a primary downside of using nonpolar substances?
   a. expense
   b. difficulty
   c. rarity
   d. explosivity
   e. toxicity

5. Which two substances form when nonpolar substances are burned?
   a. methane and nitrogen
   b. sodium chloride and oxygen
   c. carbon dioxide and water
   d. potassium iodide and iron
   e. hydrogen and helium

6. A substance with one relatively polar end attached to one relatively nonpolar end is a(n)
   a. catalyst
   b. soap
   c. acid
   d. ion
   e. alcohol

7. A solution's composition is MOST often expressed as a(n)
   a. atmosphere
   b. mole
   c. ratio
   d. proportion
   e. percentage

8. A substance is labeled "0.5% sodium fluoride %w/w". Here, the "w" stands for
   a. weak
   b. weight
   c. water
   d. work
   e. wave

9. Which measurement is expressed as moles of a solute in 1 liter of solution?
   a. concentration
   b. mass fraction
   c. molality
   d. molarity
   e. mole fraction

10. A solution containing 18.02 grams of water (molar mass = 18.02 g/mol) in 1 liter (exactly) of a salt solution has a concentration of
    a. 18.02 M
    b. 0.000 M
    c. 0.100 *M*
    d. 1.000 M
    e. 36.04 M

11. The relative numbers of moles of solute and solvent give rise to a substance's
    a. physical properties
    b. electrical properties
    c. colligative properties
    d. characteristic properties
    e. intensive properties

12. The relative numbers of moles of solute and solvent can be expressed in the unit of
    a. mole fractions
    b. mass fraction
    c. molality
    d. concentration
    e. molarity

13. The number of moles of solute dissolved in 1 kilogram (exactly) of solvent is

    a. mole fraction
    b. molarity
    c. molality
    d. mass fraction
    e. concentration

14. 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

    a. -0.486 °C
    b. -0.062 °C
    c. -0.114 °C
    d. -0.558 °C
    e. -0.372 °C

15. What does molality express?

    a. the relative amounts of different components of a mixture
    b. change in properties as concentration changes
    c. the abundance of a constituent of a mixture
    d. the strength of a solute in a solution
    e. the total internal energy of a system

SCIENCE

FOCUSED QUIZ 25

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

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1. Which of the following properties is a colligative property?
   a. concentration
   b. boiling point
   c. mass
   d. density
   e. volume

2. Which of the following properties is a colligative property?
   a. surface tension
   b. color
   c. chemical potential
   d. osmotic pressure
   e. viscosity

3. Which chemist formulated rules to explain colligative properties?
   a. François Raoult
   b. Linus Pauling
   c. John Dalton
   d. Antoine Lavoisier
   e. Amedeo Avogadro

4. Colligative properties depend on
   a. the amounts of substances present
   b. a substance's temperature
   c. the specific solute and solvent
   d. the size of a solution sample
   e. the pressure of the substance

5. 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 change to
   a. 0.186 °C
   b. 1.860 °C
   c. 0.002 °C
   d. 0.017 °C
   e. 18.60 °C

6. Why does Raoult's law work?
   a. Solute particles are heavier than solvent molecules.
   b. Solvent molecules become more resistant to change at lower temperatures.
   c. Solute particles prevent solvent molecules from entering the vapor phase.
   d. Excess solute settles to the bottom in saturated solutions.
   e. Solvents move solute when they enter the vapor phase.

7. Which type of substance causes larger than expected changes in the point at which a substance's phase shifts?
   a. lipids
   b. intermetallics
   c. salts
   d. carbides
   e. alcohols

8. What are the x and y axes on a phase diagram?
   a. pressure and volume
   b. temperature and pressure
   c. temperature and volume
   d. volume and amount
   e. volume and amount

9. Which of the following applications can be attributed to colligative properties?
   a. sterilizing equipment with UV light
   b. pretreating a stain
   c. spraying off a car window
   d. dissolving sugar in warm water
   e. adding salts to freezing roadways

10. Distillation depends on the fact that
    a. low temperatures dissolve solutions better
    b. adding solute shifts a substance's triple point
    c. vapor does not contain solute
    d. salts are heavier than water
    e. pressure increases when volume decreases

11. If one mole of $CaCl_2$ is added to a solvent, how much greater than expected will its effect be on the freezing and boiling points?
    a. three times
    b. five times
    c. two times
    d. four times
    e. six times

12. Adding more solute to a solvent will shift the triple point
    a. up and right
    b. down and left
    c. up and left
    d. down and right
    e. straight up

13. Which of the following pieces of equipment would a laboratory water distillation setup use?
    a. Erlenmeyer flask
    b. graduated cylinder
    c. friability tester
    d. buret
    e. Liebig condenser

14. What is NOT an advantage of distillation over reserve osmosis?
    a. less pretreatment needed
    b. less need for shut-downs
    c. greater water reclamation
    d. less waste to dispose of
    e. easier to scale up

15. What is NOT an advantage of reverse osmosis over distillation?
    a. lower energy needs
    b. smaller plant requirements
    c. lower discharge water temperature
    d. purer water output
    e. fewer barriers to entry

1. Whose law states that the volume of a gas is inversely proportional to its pressure?
    a. Charles's
    b. Boyle's
    c. Dalton's
    d. Avogadro's
    e. Raoult's

2. Whose law states that the volume of a gas is directly proportional to its temperature?
    a. Raoult's
    b. Dalton's
    c. Avogadro's
    d. Boyle's
    e. Charles's

3. Which measurement is based on absolute zero?
    a. M
    b. atm
    c. K
    d. F
    e. C

4. Whose law sets rules for dealing with mixtures of gases?
    a. Charles's
    b. Raoult's
    c. Dalton's
    d. Avogadro's
    e. Boyle's

5. How many particles are in any 1 mole of a substance?
    a. 6.022×10^{23}
    b. 6.26×10^{1}
    c. 4.1×10^{3}
    d. 3.303×10^{-6}
    e. 1.71771×10^{3}

6. Who proposed that every gas contains the same specific number of small particles under given conditions of temperature and pressure?
    a. Charles
    b. Boyle
    c. Avogadro
    d. Raoult
    e. Dalton

7. Which gas was discovered in the 1770s?
    a. helium
    b. oxygen
    c. neon
    d. hydrogen
    e. argon

8. What does the Van der Waals equation do?
    a. calculates temperature or volume changes
    b. converts Celsius to Kelvin
    c. adjusts the ideal gas law to account for the behavior of non-ideal gases
    d. relates the vapor pressure of a liquid to its vapor pressure in a solution
    e. determines the molality of a solution

9. What PRIMARILY differentiates liquids and solids from gases?
    a. number of electrons
    b. constituent elements
    c. inability to change shape
    d. particle weight
    e. particle freedom of motion

10. What is a foundational idea specific to kinetic-molecular theory?
    a. Entropy increases over time.
    b. Gas consists of very small rapidly moving particles.
    c. Matter can neither be created nor destroyed.
    d. All matter is composed of small, indivisible particles.
    e. Equal volumes of gases under identical conditions will contain equal numbers of particles

11. Absolute zero is defined as the temperature at which
    a. heat ceases to move from areas of high temperature to low temperature
    b. the three phases of water coexist
    c. the minimum temperature at which liquid can exist
    d. water phase shifts
    e. a thermodynamic system has the lowest energy

12. Which characteristic do metals NOT have?
    a. luster
    b. slimy feel
    c. conductivity
    d. ductility
    e. malleability

13. Compared to water, carbon dioxide
    a. dissolves polar compounds more readily
    b. has stronger intermolecular forces
    c. is more soluble in nonpolar solvents
    d. has a more typical phase diagram
    e. drives global warming less

14. Joseph Black is credited with the discovery of
    a. carbon dioxide
    b. Avogadro's number
    c. the second law of thermodynamics
    d. absolute zero
    e. molecular-orbital theory

15. Percent composition, molarity, and molality all express
    a. changes in partial pressures
    b. how much solute dissolves in a solvent
    c. changes in pressure and temperature
    d. electronegatively relationships
    e. how many products reagents form

1. Which reaction type is NOT one of the 5 basic types of chemical reactions?
   a. oxidation-reduction
   b. double replacement
   c. decomposition
   d. combustion
   e. single replacement

2. The substances on the right side of a chemical equation are called
   a. reagents
   b. reactants
   c. solvents
   d. products
   e. catalysts

3. How many products form in a synthesis reaction?
   a. 1
   b. 3
   c. 2
   d. 4
   e. 5

4. How many potassium atoms react with one chlorine molecule to form potassium chloride?
   a. 1
   b. 5
   c. 2
   d. 3
   e. 4

5. Which equation represents the decomposition of H_2O_2?
   a. $2H_2O_2 \rightarrow H_2O + O_2$
   b. $2H_2O_2 \rightarrow H_2O + 2O_2$
   c. $H_2O_2 \rightarrow H_2O + O_2$
   d. $2H_2O_2 \rightarrow 2H_2O + 2O_2$
   e. $2H_2O_2 \rightarrow 2H_2O + O_2$

6. When was oxygen gas FIRST identified?
   a. 1754
   b. 1776
   c. 1755
   d. 1774
   e. 1770

7. Which general formula represents a double replacement reaction?
   a. $N + BC \rightarrow BN + C$
   b. $AB + CD \rightarrow AD + CB$
   c. $M + BC \rightarrow MC + B$
   d. $A + B \rightarrow C$
   e. $A \rightarrow B + C$

8. The format of a single replacement reaction depends on whether the atom is
   a. alkaline earth metal or alkali metal
   b. halogen or noble gas
   c. synthetic or natural
   d. metal or nonmetal
   e. lanthanide or actinide

9. Whether a single replacement reaction will take place depends on the
   a. oxidation series
   b. enthalpy series
   c. electrochemical series
   d. activity series
   e. entropy series

10. Which metal is the MOST reactive?
    a. Cr
    b. Fe
    c. Ba
    d. Co
    e. Mg

11. Which metal will NOT displace H_2 from any source?
    a. Pb
    b. Hg
    c. Al
    d. Cd
    e. Ca

12. A combustion reaction including an organic hydrocarbon will produce water and
    a. methane
    b. carbon dioxide
    c. carbon monoxide
    d. hydrogen
    e. oxygen

13. Which metal can displace H_2 from water?

    a. Ni
    b. Zn
    c. K
    d. Cu
    e. Mn

14. Which of the following reactions will NOT occur?

    a. $CH_4(g) + 2O_2(g)$
    b. $Pb(NO_3)_2(aq) + 2KI(aq)$
    c. $Ag(s) + 2H^+(aq)$
    d. $2Li(s) + 2H_2O(l)$
    e. $Sn(s) + 2H^+(aq)$

15. Which of the following formulas represents silver nitrate?

    a. $AgNO_2$
    b. Ag_2N
    c. Ag_3N
    d. $AgNO$
    e. $AgNO_3$

SCIENCE

FOCUSED QUIZ 28

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

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1. On which process did alchemists NOT focus?
   a. observation
   b. hydrolysis
   c. distillation
   d. percolation
   e. extraction

2. Who refuted phlogiston theory with a scientific explanation for combustion?
   a. Carl Wilhelm Scheele
   b. Joseph Black
   c. Robert Boyle
   d. Antoine Lavoisier
   e. Joseph Priestley

3. Which observation MOST inspired the principle of conservation of mass?
   a. Upon the dissolving of mercury in acid, a flammable gas that was less dense than air was produced.
   b. The heating of a mercury calx produced a gas that was heavier than air and could pour like a liquid.
   c. Upon the heating of mercury, the flame of a candle in the same container was extinguished
   d. The combined masses of mercury and gas were equal to the original mass of the calx that was heated.
   e. The heating of mercury oxide produced an odorless and tasteless gas that supported respiration and combustion better than air.

4. Which of the following statements regarding Antoine Lavoisier is FALSE?
   a. His system of chemical nomenclature is still used today.
   b. He founded the law of conservation of mass.
   c. He determined that air is composed of mostly nitrogen and oxygen.
   d. He determined the role of oxygen in combustion.
   e. He proposed a theory of acid-base reactions that is still used today.

5. Who proposed the FIRST theory of acid-base reactions?
   a. Thomas Lowry
   b. Johannes Brønsted
   c. Gilbert Lewis
   d. Svante Arrhenius
   e. Antoine Lavoisier

6. In the current theory of acid-base reactions, a basic solution has an excess of
   a. H^- ions
   b. OH^- ions
   c. H_3O^+ ions
   d. H^+ ions
   e. OH^+ ions

7. Which compound is amphoteric?
   a. H_2O
   b. HCl
   c. NaCl
   d. HBr
   e. NaOH

8. The pH scale is
   a. logarithmic
   b. quadratic
   c. trigonometric
   d. rational
   e. cubic

9. A solution with a low pH contains
   a. only base and no acid
   b. mostly base and some acid
   c. equal amounts of acid and base
   d. only acid and no base
   e. mostly acid and some base

10. Why are titrations performed?
    a. to prepare a solution of a certain concentration
    b. to extract a desired precipitate
    c. to identify a chemical substance based on its concentration
    d. to measure the concentration of an unknown
    e. to measure the amount of heat transferred to or from a substance

11. Which of the following compounds are basic?
    a. $NaCl$
    b. HNO_3
    c. $NaOH$
    d. HBr
    e. KBr

12. Neutral salts form from reactions between
    a. a weak acid and a weak base
    b. a strong acid and a strong base
    c. a strong acid and a weak base
    d. a strong acid or strong base and a buffer
    e. a weak acid and a strong base

13. Which piece of lab equipment is LEAST likely to be used in a titration?
    a. Erlenmeyer flask
    b. pH indicator
    c. pH meter
    d. crucible
    e. buret

14. The equation $M_aV_a = M_bV_b$ must be written for the total acid content due to the existence of
    a. conjugate acids
    b. polyprotic acids
    c. amphoteric substances
    d. buffer solutions
    e. amphiprotic substances

15. In the equation $M_aV_a = M_bV_b$, volume MUST be measured in
    a. cubic centimeters
    b. milliliters
    c. kiloliters
    d. cubic meters
    e. liters

1. Which of the following definitions BEST represents hydrolysis?
   a. the process of increasing the rate of reaction using a chemical that is not used up
   b. the nonspontaneous splitting of a substance by adding electrical energy
   c. a reaction in which a hydroxyl group is added to an organic compound
   d. the process of thermally decomposing a substance in a nonreacting atmosphere
   e. the process in which a salt dissolves in water and its ions interact with the water molecules

2. In a basic salt reaction, an anion of a salt reacts with water to produce its conjugate acid and
   a. hydroxide ions
   b. oxide ions
   c. hydrogen ion
   d. hydronium ions
   e. hydride ions

3. Which of the following acids is the conjugate acid of the acetate ion?
   a. benzoic acid
   b. stearic acid
   c. formic acid
   d. carbonic acid
   e. acetic acid

4. Which of the following pairs INCORRECTLY matches the type of ion to the salt it forms?
   a. anions of weak acids | basic salts
   b. cations of weak bases | acidic salts
   c. cations of strong bases | neutral salts
   d. cations of weak acids | acidic salts
   e. anions of strong acids | neutral salts

5. Which anion will form a basic salt?
   a. I^-
   b. NO_3^-
   c. HCO_3^-
   d. BrO_4^-
   e. ClO_4^-

6. Which cation will form an acidic salt?
   a. K^+
   b. Mg^{2+}
   c. Ca^{2+}
   d. Cu^{2+}
   e. Cs^+

7. Precipitation reactions are a type of
   a. combustion reaction
   b. single replacement reaction
   c. double replacement reaction
   d. acid-base reaction
   e. synthesis reaction

8. Which of the following changes is an indication that a precipitate has formed?
   a. A solution appears cloudy.
   b. A solution changes color.
   c. A solution produces bubbles.
   d. A solution changes temperature.
   e. A solution produces an odor.

9. Which of the following reactions represents the net ionic reaction between silver nitrate and sodium chloride?
   a. $Na^+ + Cl^- \rightarrow NaCl$
   b. $AgNO_3 + NaCl \rightarrow AgCl\ NaNO_3$
   c. $Ag^+ + Cl^- \rightarrow AgCl$
   d. $Ag^+ + NO_3^- + Na^+ + Cl^- \rightarrow AgCl + NO_3^- + Na^+$
   e. $NO_3^- + Na^+ \rightarrow NaNO_3$

10. Which of the following ions is a spectator ion in the reaction between silver nitrate and sodium chromate?
    a. Ag^+
    b. CrO_4^+
    c. Pb^{2+}
    d. Na^+
    e. Cl^-

11. Which of the following compounds is insoluble?
    a. $Zn_3(PO_4)_2$
    b. $MgSO_4$
    c. H_2S
    d. $Ca(ClO_3)_2$
    e. $BaBr_2$

12. Which sulfate is insoluble?
    a. calcium sulfate
    b. potassium sulfate
    c. sodium sulfate
    d. aluminum sulfate
    e. lithium sulfate

13. In which condition do metal hydroxides form soluble oxide ions?
    a. in slightly acidic pH
    b. in neutral pH
    c. in acidic pH
    d. in slightly basic pH
    e. in basic pH

14. Hg (II) turns into CH_3Hg through a process called
    a. glycosylation
    b. ethylation
    c. esterification
    d. fucosylation
    e. methylation

15. Solubility principles and precipitation reactions are NOT commonly applied to
    a. silver recovery
    b. HVAC
    c. soil pH
    d. mine run-off
    e. mercury release

1. If an atom has oxidized, it has
   a. lost an electron
   b. lost an oxygen
   c. gained an oxygen
   d. gained multiple oxygens
   e. gained an electron

2. Batteries are powered by
   a. acid-base reactions
   b. oxidation-reduction reactions
   c. combustion reactions
   d. precipitation reactions
   e. decomposition reactions

3. An oxidation number is equal to
   a. a measure of the heat content of an atom or molecule experiencing constant pressure and temperature
   b. the potential electrical difference between two electrodes in an electrolytic cell
   c. the electrical potential energy difference between two electrodes when no current is present
   d. 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 atom
   e. the amount of energy required to remove one electron from an atom or molecule

4. Hydrogen's oxidation number is MOST often
   a. +2
   b. +1
   c. -1
   d. -2
   e. 0

5. In which compounds does oxygen have an oxidation number of -1?
   a. peroxides
   b. oxyhalides
   c. suboxides
   d. oxyanions
   e. metal oxides

6. Which rule does NOT apply to oxidation numbers?
   a. The oxidation numbers in polyatomic ions add up to the charge of the ion.
   b. Most atoms have only one possible oxidation number.
   c. All neutral atoms have an oxidation number of 0.
   d. The oxidation numbers in a neutral compound add up to zero.
   e. The oxidation number of a monatomic ion is equal to the ion's charge.

7. What is Cr's oxidation number in the compound K_2CrO_4?
   a. -2
   b. +6
   c. +3
   d. +1
   e. +4

8. Which of the following terms represents a positive ion?
   a. anion
   b. cathode
   c. cation
   d. electrode
   e. anode

9. The half-reaction $Al \rightarrow Al^{3+} + 3e^-$ and the half-reaction $O_2 + 4e^- \rightarrow 2O^{2-}$ form the complete equation
   a. $Al + 3O_2 \rightarrow Al^{3+} + 6O_2^-$
   b. $Al + O_2 + 4e^- \rightarrow Al^{3+} + 2O^{2-} + 3e^-$
   c. $Al + O_2 \rightarrow Al^{3+} + 2O^{2-}$
   d. $4Al + 3O_2 \rightarrow 4Al^{3+} + 6O_2^-$
   e. $4Al + O_2 \rightarrow 4Al^{3+} + 2O_2^-$

10. The MOST common example of electroplating is
    a. antimony plating
    b. tungsten plating
    c. aluminum plating
    d. manganese plating
    e. chrome plating

11. The tendency of an electron to leave or join an atom is measured as electrical energy in
    a. amperes
    b. coulombs
    c. watts
    d. ohms
    e. volts

12. Standard half-cell potentials are measured relative to the reaction
    a. $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$
    b. $F_2 + 2e^- \rightarrow 2F^-$
    c. $2H^+ + 2e^- \rightarrow H_2$
    d. $Na^+ + e^- \rightarrow Na$
    e. $Cl_2 + 2e^- \rightarrow 2Cl^-$

13. The possible voltage of a reaction is represented by the expression
    a. $E°_{red} - E°_{ox}$
    b. $-nFE° + RT\ln Q$
    c. $\ln K$
    d. $E° - \ln Q$
    e. $-nFE°$

14. Which of the following reactions is nonspontaneous?
    a. $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$
    b. $3Mn + 2Cr^{3+} \rightarrow 3Mn^{2+} + 2Cr$
    c. $Cu + 2Ag^+ \rightarrow Cu^{2+} + 2Ag$
    d. $2Ag^+ + Ni \rightarrow 2Ag + Ni^{2+}$
    e. $Li^+ + Fe^{3+} \rightarrow Li + Fe^{2+}$

15. Who invented the galvanic cell powered by the reaction $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$?
    a. Luigi Brugnatelli
    b. Michael Faraday
    c. J. F. Daniell
    d. Alessandro Volta
    e. Luigi Galvani

LUIGI BRUGNATELLI AND THE DEVELOPMENT OF ELECTROPLATING – THE NERNST EQUATION, PP. 66-69

1. Who is credited with developing the modern process of electroplating?
   a. Walther Nernst
   b. Luigi Galvani
   c. Luigi Brugnatelli
   d. Michael Faraday
   e. Alessandro Volta

2. Whose experiments disproved "animal electricity"?
   a. Walther Nernst's
   b. Luigi Galvani's
   c. Luigi Brugnatelli's
   d. Michael Faraday's
   e. Alessandro Volta's

3. How do electrochemical reduction reactions work?
   a. Negatively charged metal ions are attracted to the positive electrode, which accepts electrons, turning the metal into its neutral species, causing it to precipitate out of the solution.
   b. Positively charged metal ions are attracted to the negative electrode, which donates electrons, turning the metal into its neutral species, causing it to precipitate out of the solution.
   c. Neutral metal species oxidize upon contact with the solution, precipitate out of the solution, and collect on the electrodes.
   d. Positively charged metal ions are repulsed by the positive electrode, causing it to precipitate out of the solution.
   e. Negatively charged metal ions are repulsed by the negative electrode, causing it to precipitate out of the solution.

4. Which of the following benefits does electroplating NOT provide?
   a. wear resistance
   b. added hardness
   c. increased magnetism
   d. enhanced appearance
   e. corrosion protection

5. Which space shuttle featured 41 kilograms of gold?
   a. *Columbia*
   b. *Atlantis*
   c. *Endeavor*
   d. *Challenger*
   e. *Discovery*

6. Which of the following metals is used LEAST in electroplating?
   a. sodium
   b. rhodium
   c. nickel
   d. copper
   e. silver

7. Which telescope's mirrors are coated with gold to optimize the reflection of IR light?
   a. the James Webb Space Telescope
   b. the Spitzer Space Telescope
   c. the Hubble Space Telescope
   d. the Stratospheric Observatory for Infrared Astronomy
   e. the Kepler Space Telescope

8. Which metal can be used to make finishes that are mistakenly called white gold?
   a. platinum
   b. nickel
   c. rhodium
   d. palladium
   e. silver

9. Which metal is MOST often used to plate sterling silver and surgical instruments?
   a. platinum
   b. rhodium
   c. silver
   d. copper
   e. palladium

10. How many Faradays are necessary to convert one mole of Cr^{3+} to chromium metal?
    a. 4
    b. 2
    c. 1
    d. 6
    e. 3

11. The Nernst equation was developed to make predictions about
    a. precipitation reactions
    b. acid-base reactions
    c. combustion reactions
    d. oxidation-reduction reactions
    e. decomposition reactions

12. Which of the following equations is the MOST useful version of the Nernst equation?
    a. $E = E° - RT \ln Q / nF$
    b. $\Delta G = \Delta G° + RT \ln Q$
    c. $\Delta G = -nFE$
    d. $E = E° - (0.0592 \log Q) / n$
    e. $-nFE = -nFE° + RT \ln Q$

13. Which symbol denotes that a value is under standard state conditions?
    a. δ
    b. [ ]
    c. σ
    d. Δ
    e. °

14. When the concentrations of reactants and products are equal,
    a. $Q = 0$
    b. $n = 1$
    c. $E = E° - 1$
    d. $T = 273$
    e. $E = E°$

15. In a concentration cell, E° =
    a. R
    b. Q
    c. 0
    d. F
    e. 1

1. Stoichiometry is MOST analogous to
   a. altering baking length
   b. altering oven temperature
   c. substituting an ingredient
   d. adding extra ingredients
   e. doubling a recipe

2. What product forms when hydrogen burns in air?
   a. hydronium
   b. carbon monoxide
   c. hydrogen peroxide
   d. water
   e. carbon dioxide

3. Net ionic equations eliminate
   a. precipitates
   b. analytes
   c. supernates
   d. spectator ions
   e. titrants

4. Which type of reaction occurs when solutions of silver nitrate and sodium sulfide?
   a. synthesis
   b. single replacement
   c. combustion
   d. acid-base
   e. double replacement

5. Which of the following equations is the balanced net ionic equation for the reaction between aqueous silver nitrate and aqueous sodium sulfide?
   a. $2Ag^+(aq) + 2NO_3^-(aq) + 2Na^+(aq) + S^{2-}(aq) \rightarrow 2Na^+(aq) + 2NO_3^-(aq) + Ag_2S(s)$
   b. $2Ag^+(aq) + S^{2-}(aq) \rightarrow Ag_2S(s)$
   c. $2AgNO_3(aq) + Na_2S(aq) \rightarrow 2NaNO_3(aq) + Ag_2S(s)$
   d. $2NO_3^-(aq) + 2Na^+(aq) \rightarrow 2Na^+(aq) + 2NO_3^-(aq)$
   e. $2NO_3^-(aq) + 2Na^+(aq) \rightarrow 2NaNO_3(aq)$

6. Which of the following ions is a spectator ion in the reaction between silver nitrate and sodium sulfide?
   a. Al^{3+}
   b. N_2
   c. S^{2-}
   d. Ag^+
   e. NO_3^-

7. What is the empirical formula of acetylene?
   a. CH
   b. C_8H_{18}
   c. C_4H_6
   d. CH_4
   e. C_2H_2

8. What is the molar mass of water?
   a. 2 g/mol
   b. 16 g/mol
   c. 32 g/mol
   d. 18 g/mol
   e. 36 g/mol

9. What is the SMALLEST number of moles of oxygen required to react fully with 6 moles of hydrogen?
   a. 1
   b. 5
   c. 3
   d. 2
   e. 6

10. What is the molar mass of carbon dioxide?
   a. 32 g/mol
   b. 44 g/mol
   c. 12 g/mol
   d. 16 g/mol
   e. 36 g/mol

11. Proust's law is also known as the law of
   a. multiple proportions
   b. partial pressures
   c. conservation of energy
   d. conservation of mass
   e. definite proportions

12. Joseph-Louis Proust established his law by studying water and
    a. copper carbonate
    b. calcium carbonate
    c. chromium carbonate
    d. cesium carbonate
    e. cobalt carbonate

13. Which of the following scientists rejected Joseph-Louis Proust's law?
    a. Claude Louis Berthollet
    b. Antoine Lavoisier
    c. John Dalton
    d. Thomas Thomson
    e. Joseph Priestley

14. Which of the following scientists supported Joseph-Louis Proust's law?
    a. Robert Boyle
    b. Henry Cavendish
    c. Carl Wilhelm Scheele
    d. Claude Louis Berthollet
    e. John Dalton

15. John Dalton proposed a theory on
    a. gas properties
    b. acidic substances
    c. combustion
    d. the relationship between volume and moles
    e. atomic behavior

1. For the reaction A + B ↔ C + D, the equilibrium constant is equal to
   a. $\frac{[A][B]}{[C][D]}$
   b. $\frac{[A][D]}{[C][B]}$
   c. $\frac{[C][D]}{[A][B]}$
   d. $\frac{[C][B]}{[A][D]}$
   e. $\frac{[A][C]}{[B][D]}$

2. The concentration of reactant A is represented as
   a. <A>
   b. *A*
   c. (A)
   d. °A°
   e. [A]

3. For the reaction A + B ↔ 2C equilibrium constant is equal to
   a. $\frac{[C]^2}{[A][B]}$
   b. $\frac{[C]^{1/2}}{[A][B]}$
   c. $\frac{[A][B]}{[C]}$
   d. $\frac{[C]}{[A][B]}$
   e. $\frac{[A][B]}{[C]^2}$

4. Which equilibrium constant represents a reaction that involves gases?
   a. K_p
   b. K_a
   c. K_c
   d. K_{sp}
   e. K_b

5. Which K value belongs to the reaction that favors the products MOST?
   a. -1.6×10^{12}
   b. 1.6×10^{12}
   c. -1.4×10^{8}
   d. 1.4×10^{8}
   e. 1.2×10^{-5}

6. Molarity is a unit of
   a. mass
   b. entropy
   c. concentration
   d. enthalpy
   e. temperature

7. Which of the following compounds is excluded from equilibrium expressions?
   a. water
   b. hydronium
   c. hydroxide
   d. oxygen
   e. hydrogen cation

8. A concentration value in an equilibrium expression may also be known as
   a. free energy
   b. ideality
   c. entropy
   d. activity
   e. enthalpy

9. Which value is unitless?
   a. G
   b. E
   c. c
   d. K
   e. n

10. In aqueous solutions, the standard state concentration is a
    a. 100 M solution
    b. 10 M solution
    c. 5 M solution
    d. 50 M solution
    e. 1 M solution

11. The activity of a pure solid or liquid is equal to
    a. 10
    b. 0
    c. 1
    d. -10
    e. -1

12. K_p =
    a. $K_c(RT)^{\Delta n}$
    b. $-RT\ln(K_c)$
    c. $K_c \times K_{sp}$
    d. $\Delta G° + RT\ln(Q)$
    e. $\frac{[C]^c[D]^d}{[A]^a[B]^b}$

13. Vinegar is also known as aqueous
    a. formic acid
    b. acetic acid
    c. carbonic acid
    d. benzoic acid
    e. oxalic acid

14. Which of the following values represents the STRONGEST acid or base?
    a. $K_b = 3.6 \times 10^{11}$
    b. $K_a = 2.4 \times 10^{8}$
    c. $K_a = 1$
    d. $K_a = 1.7 \times 10^{-15}$
    e. $K_b = .9 \times 10^{-13}$

15. Which K_{sp} value indicates the LEAST soluble compound?
    a. 1.8×10^{-5}
    b. -1×10^{-10}
    c. -1.8×10^{-5}
    d. 1×10^{10}
    e. 1×10^{-10}

1. Kinetics is defined as the study of
    a. forces on and between molecules
    b. the motion of objects and their surroundings
    c. the relationship of physical properties and heat energy
    d. reaction rates and what affects them
    e. the chemical processes that cause electron motion

2. Rate laws show the relationship between the rate of reaction and
    a. temperature
    b. surface area
    c. collisions
    d. concentration
    e. mass

3. Speed =
    a. velocity/time
    b. distance/time
    c. acceleration/time
    d. acceleration/mass
    e. velocity/distance

4. The collision model is MOST like the
    a. atomic theory
    b. VSEPR model
    c. molecular theory
    d. kinetic theory
    e. molecular orbital theory

5. A chemist drawing a line tangent to the curve on a graph of concentration versus time is MOST likely trying to determine
    a. informal rate
    b. instantaneous rate
    c. stoichiometric rate
    d. average rate
    e. ionization rate

6. In which example will an Alka-Seltzer tablet dissolve FASTEST?
    a. when crushed up in hot water
    b. when intact in hot water
    c. when intact in cold water
    d. when crushed up in water that already contains Alka-Seltzer
    e. when crushed up in cold water

7. In kinetics, E_a is the abbreviation for
    a. equilibrium constant
    b. ionization energy
    c. absorption energy
    d. electron affinity
    e. activation energy

8. Activation energy is the energy
    a. released by a reaction
    b. required to transform reactants into products
    c. produced in an electrochemical reaction
    d. required to change the state of the reactants
    e. absorbed in a reaction

9. Why do catalysts affect reaction rates?
    a. They change the total difference in energy.
    b. They change the necessary activation energy.
    c. They change the surface areas of the reactants.
    d. They change the products that are created.
    e. They change the temperature of the reaction.

10. The vertical axis on a potential energy diagram is
    a. concentration
    b. reaction progress
    c. kinetic energy
    d. activation energy
    e. free energy

11. Potential energy diagrams can help determine whether a reaction
    a. involves metals or nonmetals
    b. results in an acidic or basic solution
    c. is endothermic or exothermic
    d. is a galvanic or electrolytic cell
    e. forms a precipitate

12. ΔG represents a change in
    a. free energy
    b. activation energy
    c. gas pressure
    d. enthalpy
    e. entropy

13. Reaction rate DECREASES in response to increasing
    a. catalysis
    b. concentration
    c. activation energy
    d. temperature
    e. surface area

14. How does the change in free energy compare between a catalyzed and uncatalyzed reactions?
    a. Catalyzed reactions have a higher change in free energy.
    b. Catalyzed reactions have a lower change in free energy if the reaction is exothermic but a higher change in free energy if the reaction is endothermic.
    c. Catalyzed and uncatalyzed reactions have the same change in free energy.
    d. Catalyzed reactions have a higher change in free energy if the reaction is exothermic but a lower change in free energy if the reaction is endothermic.
    e. Catalyzed reactions have a lower change in free energy.

15. Which substances act as catalysts in the body?
    a. carbohydrates
    b. proteins
    c. sugars
    d. nucleic acids
    e. enzymes

1. In a study that involves a solution by a beaker being measured for saturation, which of the following entities is the system?
    a. the air around the beaker
    b. the solution in the beaker
    c. the beaker
    d. the surface that supports the beaker
    e. the stir rod

2. Which of the following statements does NOT describe an endothermic reaction?
    a. Endothermic reactions require higher activation energies than exothermic reactions.
    b. Catalysis lowers the net change in energy of an endothermic reaction
    c. The surroundings of an endothermic reaction lose heat.
    d. Endothermic reactions have a net absorption of heat.
    e. The energy of the products of an endothermic reaction is greater than the energy of the reactants.

3. Which substance did Joseph Black discover?
    a. oxygen
    b. sodium
    c. carbon dioxide
    d. hydrogen
    e. carbon monoxide

4. Joseph Black became interested in studied heat by observing
    a. storms
    b. snow
    c. wind
    d. water level
    e. fire

5. Joseph Black experiments on heat took place in a room at
    a. 40°F
    b. 47°F
    c. 45°F
    d. 33°F
    e. 32°F

6. The liquid water in Joseph Black's experiments had a starting temperature of
    a. 40°F
    b. 39°F
    c. 32°F
    d. 33°F
    e. 47°F

7. The time taken by the ice in Joseph Black's experiments to reach its final temperature differed from the time taken for the water to reach its final temperature by a factor of
    a. 14
    b. 21
    c. 17
    d. 47
    e. 33

8. How many units of heat did the ice in Joseph Black's experiments absorb?
    a. 147
    b. 33
    c. 40
    d. 139
    e. 8

9. The heat released or absorbed by a substance when it changes its phase without changing temperature is known as
    a. heat of fusion
    b. specific heat capacity
    c. volumetric heat capacity
    d. heat of vaporization
    e. latent heat

10. Heat of vaporization can refer to the latent heat involved in
    a. condensation
    b. subliming
    c. depositing
    d. freezing
    e. melting

11. Who invented the steam engine?
    a. Pierre-Simon Laplace
    b. Joseph Black
    c. Germain Hess
    d. James Watt
    e. Antoine Lavoisier

12. With which historical period did the invention of the steam engine coincide?
    a. the Russian Revolution
    b. the English Revolution
    c. the Enlightenment
    d. the Commercial Revolution
    e. the Industrial Revolution

13. Who used the world's first ice-calorimeter?
    a. Joseph Black
    b. James Watt
    c. Pierre-Simon Laplace
    d. Germain Hess
    e. Joseph Priestley

14. State functions are determined by the
    a. initial and final states of a system
    b. pathway of the system's change
    c. theoretical behavior of a system
    d. number of factors a system depends upon
    e. means by which a system arrived at its current position

15. Which of the following quantities is a state function?
    a. heat
    b. arc length
    c. total distance
    d. volume
    e. work

1. Enthalpy is defined as the
   a. disorder of a system
   b. free energy of a system
   c. energy content of a system
   d. average kinetic energy of a system
   e. activity of a system

2. In calorimetry, the change in temperature of a reaction is compared to the value
   a. 2.09 J
   b. 0.918 J
   c. 1.01 J
   d. 4.18 J
   e. 2.03 J

3. Hess's Law is also known as the
   a. law of definite proportions
   b. second law of thermodynamics
   c. third law of thermodynamics
   d. law of conservation of energy
   e. law of constant heat summation

4. Subtracting $CO + \frac{1}{2}O_2 \rightarrow CO_2$ from the equation $C + O_2 \rightarrow CO_2$ yields the equation
   a. $CO \rightarrow C + \frac{1}{2}O_2$
   b. $2C + O_2 \rightarrow 2CO$
   c. $C + \frac{1}{2}O_2 \rightarrow CO$
   d. $2CO_2 \rightarrow \frac{1}{2}O_2 + C + CO$
   e. $C + CO + 3/2\ O_2 \rightarrow 2CO_2$

5. A negative enthalpy value indicates that a reaction is
   a. disordered
   b. ordered
   c. exothermic
   d. nonspontaneous
   e. endothermic

6. Which symbol represents entropy?
   a. H
   b. K
   c. E
   d. G
   e. S

7. All natural processes
   a. decrease the universe's entropy
   b. occur spontaneously in the forward direction
   c. occur spontaneously in the reverse direction
   d. increase the universe's entropy
   e. occur spontaneously at only low temperatures

8. If a system has decreased in entropy, the
   a. reaction is endothermic
   b. system increases in entropy
   c. reaction is exothermic
   d. surroundings increase in entropy
   e. surroundings decrease in entropy

9. Which list correctly ranks the states of matter in order of lowest to highest entropy?
   a. liquid, solution, gas, solid
   b. gas, solid, liquid, solution
   c. gas, solution, liquid, solid
   d. solid, liquid, solution, gas
   e. solution, gas, solid, liquid

10. In the equation for Gibbs free energy, temperature MUST be measured in
    a. kilojoules
    b. Fahrenheit
    c. Kelvin
    d. Celsius
    e. calories

11. Which expression is equivalent to ΔG?
    a. $\Delta H - T\Delta S$
    b. $\Sigma S_{products} - \Sigma S_{reactants}$
    c. $\Sigma H_{products} - \Sigma H_{reactants}$
    d. $mc\Delta T$
    e. $\Delta S - T\Delta H$

12. When is a reaction ALWAYS nonspontaneous?
    a. when ΔH is negative and ΔS is positive
    b. when both ΔH and ΔS are 0
    c. when both ΔH and ΔS are positive
    d. when both ΔH and ΔS are negative
    e. when ΔH is positive and ΔS is negative

13. Which ΔH and ΔS values correlate to a reaction that is spontaneous at all temperatures?
    a. $\Delta H = 110$ kJ/mol, $\Delta S = -33.7$ kJ/mol
    b. $\Delta H = -2815.8$ kJ/mol, $\Delta S = 0.6415$ kJ/mol
    c. $\Delta H = 393$ kJ/mol, $\Delta S = -131$ kJ/mol
    d. $\Delta H = 76.6$ kJ/mol, $\Delta S = 3890$ kJ/mol
    e. $\Delta H = -174.1$ kJ/mol, $\Delta S = -199$ kJ/mol

14. If ΔG is negative,
    a. the reaction proceeds spontaneously at low temperatures
    b. the reaction proceeds spontaneously at high temperatures
    c. the reverse reaction proceeds spontaneously in all directions
    d. the reverse reaction proceeds spontaneously at low temperatures
    e. the reaction proceeds spontaneously at all temperatures

15. Which reaction is spontaneous at only low temperatures?
    a. $3O_2 \rightarrow 2O_3$
    b. $2H_2O_2 \rightarrow 2H_2O + O_2$
    c. $2HgO \rightarrow 2Hg + O_2$
    d. $4Fe + 3O_2 \rightarrow 2Fe_2O_3$
    e. $NH_3 + HCl \rightarrow NH_4Cl$

1. Which of the following properties is MOST directly related to both equilibrium constants and electrode potentials?
   a. enthalpy
   b. entropy
   c. temperature
   d. pressure
   e. free energy

2. At equilibrium,
   a. $\Delta G - E = 0$
   b. $\Delta G = E = 1$
   c. $\Delta G = E = 0$
   d. $\Delta G + E = 0$
   e. $\Delta G = E = -1$

3. If a reaction has a negative ΔG, galvanic cell voltage MUST be
   a. equal to the change in free energy
   b. negative
   c. positive
   d. zero
   e. the opposite of the change in free energy

4. A positive galvanic cell voltage indicates that a reaction
   a. does not proceed
   b. is at equilibrium
   c. is spontaneous in the forward direction
   d. is spontaneous in the forward direction at specific temperatures
   e. is spontaneous in the reverse diagram

5. For prediction and comparison purposes, it is best to use ΔG and E values for the reactants and products in their
   a. liquid states
   b. solid states
   c. solvent states
   d. gaseous states
   e. standard states

6. From which two values can the value of K be DIRECTLY determined?
   a. ΔH and ΔG
   b. ΔG and ΔS
   c. ΔG and E
   d. ΔS and E
   e. ΔH and ΔS

7. Which of the following values is considered a large equilibrium constant?
   a. -0.7
   b. 0
   c. 0.7
   d. 1.2
   e. -1.2

8. Which equation do chemists use to change from free energy to equilibrium constants?
   a. $\Delta G° = -nFE°_{cell}$
   b. $\Delta G° = T\Delta S - \Delta H$
   c. $\Delta G° = -RT\ln K$
   d. $\Delta G_{rxn}° = -$
   e. $\Delta G° = \Delta H - T\Delta S$

9. Which of the following equations is BEST for changing from free energy to standard cell potential?
   a. $\Delta G° = -nFE°_{cell}$
   b. $E = E°_{cell} -$
   c. $E°_{cell} = E°_{cathode} - E°_{anode}$
   d. $\Delta G° = \Delta H - T\Delta S$
   e. $\Delta G° = -RT\ln K$

10. For the signs of the K and E values to be correct,
    a. all other values must be absolute values
    b. all temperature values must be measured in Celsius
    c. both sides of the equation must have negative signs
    d. there must be a negative sign on one side of an equation
    e. one side of the equation must include a natural logarithm

11. Given a galvanic cell, which two values are ALWAYS positive?
    a. K and E
    b. E and ΔH
    c. ΔS and K
    d. E and ΔG
    e. K and $E°_{cell}$

12. If K is larger than 1, the reaction favors the

    a. reactants only at high temperatures
    b. products only at high temperatures
    c. products
    d. products and reactants equally
    e. reactants

13. What can be said about a reaction if $E°_{cell}$ is negative?

    a. It favors the formation of reactants at all temperatures.
    b. It favors reactants and products equally at all temperatures.
    c. It favors the formation of products at all temperatures.
    d. It favors the formation of reactants only at low temperatures.
    e. It favors the formation of products only at low temperatures.

14. If $\Delta G°$ is negative, the reaction favors the

    a. products and reactants equally
    b. reactants
    c. reactants only at low temperatures
    d. products only at low temperatures
    e. products

15. At equilibrium,

    a. $K = E°_{cell}$
    b. $K = 1$
    c. $K = 0$
    d. $K = \Delta G°$
    e. $K = -1$

1. Combustion reactions involve substances reacting with
   a. carbon
   b. carbon dioxide
   c. oxygen
   d. hydrogen
   e. water

2. Antoine Lavoisier's experiments led to a better understanding of
   a. decomposition reactions
   b. combustion reactions
   c. precipitation reactions
   d. single replacement reactions
   e. oxidation-reduction reactions

3. Which of the following pH values indicates the MOST acidic reaction?
   a. 7
   b. 10
   c. 5
   d. 2
   e. 13

4. Which of the following formulas represents a hydronium ion?
   a. H_3O^+
   b. O^{2-}
   c. OH^-
   d. H^-
   e. OH^+

5. When a substance acts as either an acid or a base, it is considered
   a. triprotic
   b. isoelectric
   c. amphoteric
   d. diprotic
   e. organic

6. A buret is MOST likely to be used in
   a. calorimetry
   b. distillation
   c. catalysis
   d. titrations
   e. electrochemical cells

7. The equivalence point is also called the
   a. supercritical point
   b. triple point
   c. curie point
   d. critical point
   e. endpoint

8. Which type of reactions produce salts?
   a. oxidation-reduction reactions
   b. decomposition reactions
   c. precipitation reactions
   d. combustion reactions
   e. acid-base reactions

9. Electrochemistry relies on
   a. oxidation-reduction reactions
   b. acid-base reactions
   c. precipitation reactions
   d. decomposition reactions
   e. combustion reactions

10. The MOST useful version of the Nernst equation is
    a. $\Delta G° = -nFE°$
    b. $\Delta G° = -RT\ln K$
    c. $\Delta G° = \Delta H - T\Delta S$
    d. $E = E° - (0.0592 \log Q)/n$
    e. $E°_{cell} = E°_{red} - E°_{ox}$

11. For a reaction $aA + bB \rightarrow cC + dD$, K =
    a. $\frac{[A]^a[B]^b}{[C]^c[D]^d}$
    b. $\frac{[B]^b[D]^d}{[C]^c[A]^a}$
    c. $\frac{[B]^b[C]^c}{[A]^a[D]^d}$
    d. $\frac{[C]^c[D]^d}{[A]^a[B]^b}$
    e. $\frac{[A]^a[C]^c}{[B]^b[D]^d}$

12. Which K value is used for a precipitation reaction?

    a. K_b
    b. K_c
    c. K_p
    d. K_a
    e. K_{sp}

13. A reaction favors the reactants if K is less than

    a. -10
    b. 1
    c. -1
    d. 0
    e. 10

14. Which of the following quantities is NOT a state function?

    a. work
    b. temperature
    c. pressure
    d. entropy
    e. volume

15. If $\Delta G° = 0$, a reaction

    a. is endothermic
    b. favors the products
    c. favors the reactants
    d. is at equilibrium
    e. is exothermic

1. Acids have an excess of
   a. O^-
   b. H^-
   c. H^+
   d. OH^-
   e. O^{2-}

2. Alchemy was MOST popular during the
   a. Enlightenment
   b. Age of Discovery
   c. Neoclassical Period
   d. Gilded Age
   e. Middle Ages

3. Which of the following particles is an alpha particle?
   a. ${}^{4}_{1}He^{2+}$
   b. ${}^{2}_{2}He^{2+}$
   c. ${}^{4}_{2}He^{2+}$
   d. ${}^{2}_{4}He^{2+}$
   e. ${}^{4}_{1}He^{2+}$

4. An atomic mass unit may also be called a(n)
   a. Rutherford
   b. Thompson
   c. Bohr
   d. Dalton
   e. Pauling

5. Which of the following numbers is termed Avogadro's number?
   a. 1.381×10^{-23}
   b. 6.022×10^{23}
   c. 2.988×10^{8}
   d. 6.626×10^{-34}
   e. 8.314×10^{17}

6. A binary substance has
   a. 2 oxidation states
   b. 2 hydrogens
   c. 2 elements
   d. 2 oxygens
   e. a charge of +2

7. Ductility describes a metal's ability to
   a. be hammered into sheets
   b. create sound when struck
   c. be pulled into wires
   d. conduct electricity
   e. create magnetic fields

8. Who discovered the electron?
   a. Ernest Rutherford
   b. Niels Bohr
   c. John Dalton
   d. Erwin Schrödinger
   e. J. J. Thompson

9. Enthalpy has the symbol
   a. Q
   b. G
   c. H
   d. E
   e. S

10. Which of the following equations represents the ideal gas law?
    a. $P_1V_1 = P_2V_2$
    b. $\frac{P_1}{T_1} = \frac{P_2}{T_2}$
    c. $\frac{V_1}{T_1} = \frac{V_2}{T_2}$
    d. $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
    e. $PV = nRT$

11. Water boils at
    a. 212 K
    b. 298 K
    c. 273 K
    d. 100 K
    e. 373 K

12. The units of molarity are
    a. moles/mole
    b. kilograms/kilogram
    c. grams/liter
    d. moles/kilogram
    e. moles/liter

13. Steel is composed of MOSTLY iron and
    a. carbon
    b. chromium
    c. nickel
    d. molybdenum
    e. cobalt

14. Which of the following acids is triprotic?
    a. HNO_3
    b. HNO_2
    c. HF
    d. H_2SO_4
    e. H_3PO_4

15. Volts are units of
    a. current
    b. free energy
    c. electrical potential
    d. resistance
    e. charge

1. Which of the following pairs of elements will have the MOST similar properties?
   a. silicon and germanium
   b. magnesium and manganese
   c. boron and sulfur
   d. lead and mercury
   e. oxygen and indium

2. On the periodic table, elements are arranged in order of
   a. number of protons
   b. electron affinity
   c. atomic radius
   d. electronegativity
   e. discovery

3. Group 18 elements are also known as the
   a. lanthanoids
   b. noble gases
   c. halogens
   d. metalloids
   e. actinoids

4. Which two elements are diatomic?
   a. nitrogen and iodine
   b. nobelium and helium
   c. neon and argon
   d. chlorine and xenon
   e. oxygen and oganesson

5. Which of the following pairs of elements are CLOSEST to each other on the periodic table?
   a. tennessine and californium
   b. antimony and astatine
   c. strontium and seaborgium
   d. holmium and hafnium
   e. tantalum and tungsten

6. Which types of elements lie between the post-transition metals and the alkaline earth metals?
   a. noble gases
   b. transition metals
   c. alkali metals
   d. halogens
   e. metalloids

7. How many elements are on the periodic table?
   a. 76
   b. 154
   c. 108
   d. 118
   e. 99

8. Which of the following elements is an alkali metal?
   a. Cs
   b. Cn
   c. Cu
   d. Co
   e. Ca

9. Which type of element is phosphorous?
   a. halogen
   b. post-transition metal
   c. diatomic nonmetal
   d. polyatomic nonmetal
   e. metalloid

10. Which elemental symbol belongs to manganese?
    a. Md
    b. Mn
    c. Mg
    d. Mo
    e. Mc

11. Which element is part of the lowest group of the periodic table?
    a. Rb
    b. Kr
    c. Ba
    d. Be
    e. Sc

12. Which element is part of period 7 on the periodic table?
    a. yttrium
    b. copernicium
    c. ytterbium
    d. ruthenium
    e. antimony

13. Which element has the largest mass?
    a. He
    b. Fr
    c. Og
    d. Lr
    e. F

14. Where in the periodic table are MOST synthetic elements located?
    a. period 6
    b. period 8
    c. period 7
    d. period 3
    e. period 1

15. To which family does hydrogen belong?
    a. diatomic nonmetals
    b. polyatomic nonmetals
    c. halogens
    d. alkali metals
    e. noble gases