

Chemistry

1. What is the branch of chemistry that is being applied in measuring the concentration of an air pollutant?
 - A) analytical chemistry
 - B) biochemistry
 - C) inorganic chemistry
 - D) organic chemistry
 - E) physical chemistry

2. Which state of matter has no definite shape or volume?
 - A) liquid
 - B) solid
 - C) vapor
 - D) steam
 - E) gas

3. What kind of change does NOT alter the composition or identity of the substance undergoing the change?
 - A) molecular
 - B) endothermic
 - C) exothermic
 - D) physical
 - E) chemical

4. What do we call a statement of observed behavior for which no exceptions have been found?
 - A) hypothesis
 - B) theory
 - C) law
 - D) model
 - E) result

5. What kind of change always results in the formation of new materials?
 - A) molecular
 - B) exothermic
 - C) endothermic
 - D) physical
 - E) chemical

6. Which one of the following is an example of an extensive property?

- A) density
- B) specific gravity
- C) mass
- D) hardness
- E) boiling temperature

7. Which one of the following is an example of a pure substance?

- A) ethyl alcohol
- B) sugar water
- C) salt and pepper
- D) milk
- E) sand

8. Air is a/an

- A) element
- B) compound
- C) mixture
- D) molecule
- E) pure substance

9. The speed of light is 186,000 miles per second. What is its speed in centimeters per second? [Use: 5,280 feet = 1 mile; 12 inches = 1 foot; 2.54 cm = 1 inch]

- A) 3.01×10^{11} cm/s
- B) 3.01×10^{10} cm/s
- C) 6.06×10^{12} cm/s
- D) 3×10^{11} cm/s
- E) 2.99×10^{10} cm/s

10. 1 kilometer equals how many millimeters?

- A) 10^{-6}
- B) 10^{-3}
- C) 10^3
- D) 10^4
- E) 10^6

11. Round off 0.052018 to three significant figures.

- A) 0.05
- B) 0.052
- C) 0.0520
- D) 0.05201
- E) 0.05202

12. Select the answer which best expresses the result of the following calculation:

$$1.86 + 246.4 - 79.9208 = ?$$

- A) 168
- B) 168.3
- C) 168.34
- D) 168.339
- E) 168.3392

13. The appropriate number of significant figures to be used in expressing the result of 51.6×3.1416 is

- A) 1
- B) 2
- C) 3
- D) 4
- E) 5

14. What Celsius temperature corresponds to -4.6°F ?

- A) -20°C
- B) -20.3°C
- C) -23.0°C
- D) -10.9°C
- E) -68.4°C

15. What Fahrenheit temperature corresponds to -40.0°C ?

- A) -8°F
- B) 16.8°F
- C) -36.9°F
- D) -40.0°F
- E) -1.94°F

16. What Kelvin temperature corresponds to 98.6°F?

- A) 310K
- B) 310.K
- C) 31.00K
- D) 132.0K
- E) 199K

17. Which temperature scale does not use a degree sign?

- A) Celsius
- B) Kelvin
- C) Centigrade
- D) Fahrenheit
- E) Absolute Zero

18. If the density of carbon tetrachloride is 1.59 g/mL, what is the volume, in L, of 4.21 kg of carbon tetrachloride?

- A) 0.149 L
- B) 0.378 L
- C) 2.65 L
- D) 6.69 L
- E) 6690 L

19. What is the specific gravity of an object that weighs 13.35 g and has a volume of 25.00 mL? The density of water under the same conditions is 0.980 g/mL.

- A) 1.335
- B) 0.545 g/mL
- C) 1.335 mL
- D) 0.545
- E) 0.980

20. Organic chemistry is the study of those chemical processes that are found in living systems.

Chemistry

1. What kind(s) of particles can be found in the nucleus of an atom?
 - A) protons
 - B) neutrons
 - C) electrons
 - D) protons and electrons
 - E) protons and neutrons

2. The total mass of the protons in any neutral atom is about _____ times the total mass of electrons in the atom.
 - A) 0.0005
 - B) 0.3
 - C) 1
 - D) 2
 - E) 2000

3. What is the quantity represented by the mass number minus the atomic number?
 - A) number of atoms
 - B) number of neutrons
 - C) number of electrons
 - D) number of protons
 - E) number of particles in the nucleus

4. Which isotope of hydrogen has two neutrons?
 - A) hydrogen-1
 - B) hydrogen-2
 - C) hydrogen-3
 - D) deuterium
 - E) H₂

5. Which of the following accounts for the fact that chlorine has an atomic mass of 35.45 amu rather than a whole number?
 - A) isotopes
 - B) electrons
 - C) protons
 - D) radioactivity
 - E) isomers

6. Who discovered that cathode rays consist of a stream of negative particles, electrons?
- A) Crookes
 - B) Thomson
 - C) Geiger
 - D) Rutherford
 - E) Bohr
7. Who discovered the existence of the atomic nucleus?
- A) Crookes
 - B) Thomson
 - C) Geiger
 - D) Rutherford
 - E) Bohr
8. In Geiger's experiment which led to the discovery of the atomic nucleus, what type of particle or ray was fired at the gold foil target?
- A) alpha
 - B) beta
 - C) gamma
 - D) neutrons
 - E) cathode rays
9. Which of the following statements relating to Bohr's model of the hydrogen atom, is incorrect?
- A) The lowest energy orbit has quantum number $n = 1$
 - B) The highest energy orbits are furthest from the nucleus
 - C) In a transition from the $n = 3$ to the $n = 1$ level, light is emitted
 - D) Energy differences between energy levels can be calculated from the wavelengths of the light absorbed or emitted
 - E) The greater the energy difference between two levels, the longer the wavelength of the light absorbed or emitted
10. Who proposed that electrons could behave like waves, as well as like particles?
- A) Thomson
 - B) Rutherford
 - C) Bohr
 - D) de Broglie
 - E) Heisenberg
11. In the calcium atom represented by the symbol $^{40}_{20}\text{Ca}$, there are 20 protons, 20 neutrons and 20 electrons.

12. All atoms of a particular element have identical properties.
13. An atom cannot be created, divided, destroyed or converted to any other type of atom.
14. The atomic number of an ion tells us the number of protons that are present.
15. In an atom gains one electron, it becomes a cation.
16. The first experimentally based theory of atomic structure was proposed by John Dalton.
17. Short wavelengths of electromagnetic radiation have more energy than long wavelengths.

Chemistry

1. The modern periodic table is arranged according to what property?
 - A) atomic number
 - B) mass number
 - C) atomic mass
 - D) neutron number
 - E) density

2. What do we call a complete horizontal row of elements on the periodic table?
 - A) group
 - B) period
 - C) family
 - D) representative elements
 - E) transition elements

3. What are all the elements in the A-groups often called?
 - A) transition elements
 - B) lanthanides
 - C) metals
 - D) non-metals
 - E) representative elements

4. Which of the following elements is a metalloid?
 - A) C
 - B) Ge
 - C) Pb
 - D) N
 - E) P

5. Where are the alkali metals located on the periodic table?
 - A) representative elements
 - B) transition metals
 - C) Group IA (1)
 - D) Group IIA (2)
 - E) Group IIIA (3)

6. How many valence electrons are in an atom of carbon?

- A) 8
- B) 6
- C) 4
- D) 1
- E) 0

7. What is the lowest energy sublevel of a principal level?

- A) d
- B) e
- C) f
- D) s
- E) p

8. How many sublevels are there in the third principal energy level?

- A) 3
- B) 2
- C) 1
- D) 0
- E) 4

9. How many orbitals are there in a p sublevel?

- A) 2
- B) 3
- C) 1
- D) 0
- E) 4

10. Which of the following correctly gives the electron capacity of a principal energy level in terms of the number n ?

- A) n
- B) $2n$
- C) $2n + 2$
- D) n^2
- E) $2n^2$

11. What is the electron configuration of sulfur, atomic number 16?

- A) $1s^2 1p^6 2s^2 2p^6$
- B) $1s^2 2s^2 2p^6 2d^6$
- C) $1s^2 2s^2 2p^6 3s^2 3p^4$
- D) $1s^2 2s^2 2p^6 3s^2 3d^4$
- E) $1s^2 2s^2 2p^6 3s^2 2d^4$

12. Which one of the following electron configurations is appropriate for a normal atom?

- A) $1s^1 2s^1$
- B) $1s^2 2s^1$
- C) $1s^2 2s^2 2p^8$
- D) $1s^2 2s^2 2p^4 3s^1$
- E) $1s^2 2s^2 2p^6 3d^1$

13. Which of the following elements is most likely to form a 3+ ion?

- A) Li
- B) K
- C) Al
- D) N
- E) Cu

14. Give the complete electronic configuration of a sodium ion.

- A) $1s^2 2s^2 2p^5$
- B) $1s^2 2s^2 2p^6$
- C) $1s^2 2s^2 2p^6 3s^1$
- D) $1s^2 2s^2 2p^6 3s^2$
- E) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$

15. Which of the following ions does not follow the octet rule?

- A) Na^+
- B) Ca^{2+}
- C) Al^{3+}
- D) N^{3-}
- E) Cl^{2-}

16. Which of the following atoms has the biggest size (radius)?

- A) Na
- B) Al
- C) Cl
- D) Rb
- E) I

17. Which of the following elements has the highest ionization energy?

- A) Li
- B) B
- C) O
- D) F
- E) Ne

18. Which of the following elements has the lowest ionization energy?

- A) Li
- B) B
- C) O
- D) F
- E) Ne

19. The electron affinity is

- A) the energy required to remove an electron from an isolated atom
- B) the force between two electrons in the same orbital
- C) the force between two ions of opposite charge
- D) the energy released when an isolated atom gains an electron
- E) the attraction of an atom for an electron in a chemical bond

20. Which one of the following elements has the highest electron affinity?

- A) Li
- B) K
- C) Kr
- D) O
- E) Cl

Chemistry

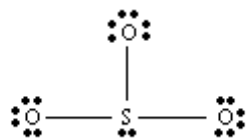
1. Draw the Lewis structure of the bromine atom.
2. How many dots are shown in the Lewis structure for the sulfur atom?
3. What are the two principal types of bonding called?
4. Name the two classes of element which are most likely to form an ionic compound if they are allowed to react with each other.
5. Draw the Lewis structure of the Pb^{2+} ion.
6. What constitutes a covalent bond between two atoms?
7. In what way is a polar covalent bond similar to a nonpolar covalent bond? In what way are they different?
8. What does it mean if an atom is said to have a high electronegativity?
9. The elements with the lowest electronegativities are found in the _____ region of the periodic table.
10. Who first assigned electronegativity values to many of the elements?
11. What do we call the three-dimensional arrangement of positive and negative ions in an ionic solid?
12. Predict the formula of the compound formed when ions of sodium and sulfur combine.
13. Predict the formula of the compound formed when ions of barium and nitrogen combine.

14. What is the name of Fe^{2+} in the Stock system?
15. What does the suffix “-ous” on the common names of ions mean?
16. What is the term used for ions that are composed of two or more atoms bonded together?
17. What is the formula of the sulfate ion?
18. What is the name of the ion HCO_3^- ?
19. What is the name of the ion NH_4^+ ?
20. Provide the name of Na_3PO_4 .
21. What is the name of Cu_2O in the Stock system?
22. Write the formula of sodium carbonate.
23. What kind of compound results when two or more different nonmetals share electrons?
24. What kind of bonding exists in substances which consist of discrete molecules?
25. Provide the formula of sulfur trioxide.
26. Write the formula of ammonia.
27. Provide the name of CCl_4 .
28. Provide the name of the compound whose formula is N_2O_5 .

29. At what temperature is a liquid converted into a gas?
30. What is the term that describes a solid with no regular structure?
31. What is the term that describes a compound that, when dissolved in water conducts an electric current?
32. What is the term that describes a compound that, when dissolved in water does not conduct an electric current?
33. What kind of bonding is present in substances which are nonelectrolytes?
34. How many bonding electrons are shown in the Lewis structure for the bicarbonate ion, HCO_3^- ?
35. Draw the Lewis structure of methylamine, CH_3NH_2 .

36. Draw the Lewis structure of hydrogen sulfide, H_2S .

37. What is wrong with the Lewis structure show below for sulfur trioxide, SO_3 ?



38. Ozone, O_3 , has two resonance forms. Draw them, given the skeletal arrangement
 $\text{O}-\text{O}-\text{O}$

39. What is defined as the amount of energy needed to break a bond holding two atoms together?

40. What is defined as the distance of separation of two nuclei in a covalent bond?

41. What do the letters VSEPR stand for?
42. If the shape of a molecule is trigonal planar, what are the values of the bond angles?
43. In the molecule AX_2 , the central atom A has two lone pairs of electrons in addition to the two bond pairs in the A—X bonds. What is the shape of this molecule?
44. The ammonia molecule, NH_3 , is polar. Why does this fact suggest that its shape is trigonal pyramidal, rather than trigonal planar?
45. Which of the following Lewis structures of neutral atoms is correct?
- A) $K\cdot$
- B) $\cdot Ba \cdot$
- C) $\cdot \overset{\cdot}{Al} \cdot$
- D) $\cdot \overset{\cdot}{Sn} \cdot$
- E) $\cdot \overset{\cdot}{O} \cdot$
46. Which of the following Lewis structures of ions is incorrect?
- A) $[Na]^+$
- B) $[Ca]^{2+}$
- C) $[Sn]^{2+}$
- D) $[\overset{\cdot\cdot}{\underset{\cdot\cdot}{N}}:]^{3-}$
- E) $[\overset{\cdot\cdot}{\underset{\cdot\cdot}{I}}:]^-$
47. Which of the following has the greatest electronegativity?
- A) H
- B) Cl
- C) O
- D) F
- E) Na

48. Which of the following has the greatest electronegativity?
- A) Si
 - B) P
 - C) Cl
 - D) Ar
 - E) Br
49. In the compound CH_3Cl the bond between carbon and chlorine is
- A) intermolecular
 - B) ionic
 - C) nonpolar covalent
 - D) polar covalent
50. Which one of the following is NOT true about elements that form cations?
- A) The atoms lose electrons in forming ions.
 - B) The elements are metals.
 - C) They are located to the left of the periodic table.
 - D) They have low ionization energies.
 - E) They have high electron affinities.
51. Assuming reactions between the following pairs of elements, which pair is most likely to form an ionic compound?
- A) copper and tin
 - B) chlorine and oxygen
 - C) cesium and iodine
 - D) carbon and chlorine
 - E) fluorine and iodine
52. What kind of bond results when electron transfer occurs between atoms of two different elements?
- A) ionic
 - B) covalent
 - C) nonpolar
 - D) single
 - E) double

53. What is the old name of Cu^+ ?

- A) cupric ion
- B) cuprous ion
- C) copper (I) ion
- D) copper (II) ion
- E) ferrous ion

54. Give the name of FeSO_4 in the Stock system.

- A) iron monosulfuric acid
- B) iron (II) sulfate
- C) iron (III) sulfate
- D) ferrous sulfate
- E) ferric sulfate

55. Assuming reactions between the following pairs of elements, which pair is most likely to form a covalent compound?

- A) lithium and iodine
- B) sodium and oxygen
- C) calcium and chlorine
- D) copper and tin
- E) carbon and oxygen

56. A double bond between two atoms, A and B

- A) is longer than a single bond between the same two atoms
- B) has a lower bond energy than a single bond between the same two atoms
- C) arises when two electrons are transferred from A to B
- D) consists of two electrons shared between A and B
- E) consists of four electrons shared between A and B

57. What is the correct formula of phosphorus pentachloride?

- A) PCl
- B) PCl_3
- C) PCl_5
- D) P_2Cl_5
- E) P_5Cl

58. What term describes the temperature at which a solid is converted into a liquid?

- A) critical point
- B) flash point
- C) sublimation point
- D) melting point
- E) boiling point

59. What term describes a solution of a compound in water that conducts an electric current?

- A) amorphous solution
- B) an electrolyte solution
- C) a nonelectrolyte solution
- D) superconducting solution
- E) isoelectric solution

60. How many bonding electrons are in CO_2 ?

- A) 1
- B) 2
- C) 3
- D) 4
- E) 8

61. How many nonbonding electrons are in CH_4 ?

- A) 0
- B) 1
- C) 2
- D) 3
- E) 8

62. How many valence electrons are in SO_4^{2-} ?

- A) 2
- B) 64
- C) 32
- D) 12
- E) 16

63. According to VSEPR theory, if the valence electrons on a central atom are 3 bond pairs and one nonbonding (lone) pair, the geometry (shape) at this atom will be

- A) linear
- B) bent (angular)
- C) trigonal planar
- D) trigonal pyramidal
- E) tetrahedral

64. In a Lewis structure, what do the dots represent?

Chemistry

1. How many iron atoms are present in one mole of iron?
2. How many grams of sulfur are found in 0.150 mol of sulfur? [Use atomic weight: S, 32.06 amu]
3. How many moles of sulfur are found in 1.81×10^{24} atoms of sulfur? [Use atomic weight: S, 32.06 amu]
4. How many atoms are present in a 7.31 g sample of copper? [Use atomic weight: Cu, 63.55 amu]
5. What is the mass, in grams, of 1.79 mol of helium, the gas commonly used to fill party balloons and lighter-than-air ships? [Use atomic weight: He, 4.00 amu]
6. An iodine sample contains 2.91×10^{22} atoms of iodine. What is its mass in grams? [Use atomic weight: I, 126.9]
7. When a solid compound is described as a “hydrate”, what does this mean?
8. What is the difference in meaning between “2O” and “O₂” when they occur in chemical equations?
9. 6.022×10^{23} molecules of a covalent compound is equal to how many moles of that compound?
10. What is the mass, in grams, of one mole of diatomic hydrogen? [Use molar mass: H, 1.0 g/mol]
11. How many molecules of water are there in 5.00 mol of water?
12. What law states that matter cannot be gained or lost during a chemical reaction?

13. What does the symbol “(aq)”, often found in chemical equations, mean?
14. In chemical equations, what are the meanings of the symbols s, l, and g, used in parentheses?
15. Balance the following equation: $\text{Ca(s)} + \text{HCl(g)} \rightarrow \text{CaCl}_2\text{(s)} + \text{H}_2\text{(g)}$
16. Balance the following equation: $\text{Na(s)} + \text{Cl}_2\text{(g)} \rightarrow \text{NaCl(s)}$
17. How many moles of oxygen gas are needed to react with hydrogen to form one mole of water?
 $2\text{H}_2\text{(g)} + \text{O}_2\text{(g)} \rightarrow 2\text{H}_2\text{O(l)}$
18. Iron reacts with oxygen to form iron(III) oxide (Fe_2O_3). How many grams of product will be formed from 5.00 grams of Fe? [Use atomic weights: Fe, 55.85 amu; O, 16.00 amu]
19. Magnesium hydroxide (Mg(OH)_2), as “Milk of Magnesia” can be used to neutralize excess stomach acid, represented by HCl(aq) :
 $\text{Mg(OH)}_2\text{(s)} + 2\text{HCl(aq)} \rightarrow \text{MgCl}_2\text{(aq)} + 2\text{H}_2\text{O(l)}$
When 5.00 g of HCl are combined with an excess of Mg(OH)_2 , what mass of MgCl_2 can be produced? [Use atomic weights: H, 1.01 amu; O, 16.00 amu; Mg, 24.31 amu; Cl, 35.45 amu]
20. Which two states of matter are the least compressible?
21. What device is used to measure atmospheric pressure?
22. The pascal (Pa) is a unit for expressing what quantity?
23. What experimental quantity measures force per unit area?
24. State Boyle’s Law.

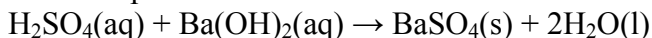
25. State Charles's Law.
26. State Avogadro's Law.
27. Consider 1.00 L of air in a patient's lungs at 37.0°C and 1.00 atm pressure. What volume would this air occupy if it were at 25.0°C under a pressure of 5.00×10^2 atm (a typical pressure in a compressed air cylinder)?
28. If a gas sample is at STP (Standard Temperature and Pressure), what are its temperature and pressure?
29. Calculate the density of oxygen gas (O₂) at STP, in g/L. [Use molar mass: O₂, 32.0 g/mol]
30. What is the volume (L) occupied by a mole of an ideal gas, if the pressure is 626 mmHg and the temperature is 25.0°C?
31. A gas sample is prepared in which the components have the following partial pressures: nitrogen, 555 mmHg; oxygen, 149 mmHg; water vapor, 13 mmHg; argon, 7 mmHg. What is the total pressure of this mixture?
32. Choose the best one from each pair in the following sentence: The closest approach to ideal gas behavior will be shown by (NH₃ or H₂?) at (low or high?) pressure and (low or high?) temperature.
33. What is meant by viscosity?
34. What experimental quantity is a measure of the attractive forces between molecules at the surface of a liquid?
35. What process is responsible for the formation of dew on the grass early in the morning?

36. How can pure water be made to boil at a temperature above the 100°C?
37. Who postulated the existence of temporary dipole attraction among nonpolar molecules?
38. What are the general structural requirements for a compound to display hydrogen bonding?
39. Why is hydrogen bonding more extensive in water than in hydrogen fluoride?
40. Name the four main types of crystalline solid, and give an example of each.
41. Classify the following reaction as precipitation, acid-base or oxidation-reduction:
$$\text{Ce}^{4+}(\text{aq}) + \text{Fe}^{2+}(\text{aq}) \rightarrow \text{Ce}^{3+}(\text{aq}) + \text{Fe}^{3+}(\text{aq})$$
42. Classify the following reaction as precipitation, acid-base or oxidation-reduction:
$$\text{H}_2\text{SO}_4(\text{aq}) + 2\text{KOH}(\text{aq}) \rightarrow \text{K}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O}(\text{aq})$$
43. Classify the following reaction as precipitation, acid-base or oxidation-reduction:
$$\text{Na}_2\text{S}(\text{aq}) + \text{CuSO}_4(\text{aq}) \rightarrow \text{Na}_2\text{SO}_4(\text{aq}) + \text{CuS}(\text{s})$$
44. Complete the products and balance the following equation for an acid-base reaction:
$$\text{HCl}(\text{aq}) + \text{KOH}(\text{aq}) \rightarrow$$
45. Complete the products and balance the following equation for a precipitation reaction:
$$\text{FeSO}_4(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow$$
46. Write a balanced equation for the reaction between zinc metal and iron(III) ions to form zinc(II) ions and iron(II) ions (Symbols: zinc = Zn; iron = Fe).
47. Name the two products formed when octane (C₈H₁₈) burns completely in excess oxygen gas.

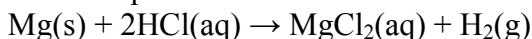
48. Classify the following reaction as decomposition, combination, single-replacement or double-replacement:



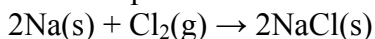
49. Classify the following reaction as decomposition, combination, single-replacement or double-replacement:



50. Classify the following reaction as decomposition, combination, single-replacement or double-replacement:



51. Classify the following reaction as decomposition, combination, single-replacement or double-replacement:



52. What would you call a liquid that displayed the Tyndall effect?

53. A solution contains 1.65g of NaOH in a total volume of 150.0mL. What is its concentration expressed as % (W/V)?

54. Define “molarity”.

55. How many moles of KNO₃ are contained in one liter of 0.2M KNO₃ solution?

56. How many moles of KCl are present in 50.0mL of a 0.552M solution?

57. If 5.20g of HCl is added to enough distilled water to form 3.0L of solution, what is the molarity of the solution? [Use molecular weight: HCl, 36.46 amu]

58. Calculate the molarity of a solution if 300.0mL of it contains 16.8g of KNO₃. [Use formula weight: KNO₃, 101.11 amu]

59. What is the molarity of 50.0mL of a 0.660M NaOH solution after it has been diluted to 450.0mL?
60. In one sentence, explain what is meant by a colligative property.
61. An aqueous solution is warmed from 20°C to 30°C. Does this change in temperature affect either the molarity or the molality of the solution? Explain.
62. How much more will the freezing point of water be lowered by adding one mole of sodium chloride rather than one mole of glucose, C₆H₁₂O₆?
63. What type of membranes allow solvent molecules to pass through but do not allow solute molecules to pass through?
64. What happens when a hypotonic solution is separated from a hypertonic solution by an osmotic membrane?
65. What term do we use to describe a solution of lower osmolarity compared to one with a higher osmolarity?
66. Describe what happens when red blood cells are placed in a hypertonic solution. Give the name of the process which occurs.
67. What happens to red blood cells that are placed in an isotonic solution?
68. Calculate the osmolarity of a $2.0 \times 10^{-3}M$ Na₃PO₄ solution. Na₃PO₄ is an ionic compound and produces an electrolytic solution.
69. Calculate the osmotic pressure of a $6.0 \times 10^{-2}M$ solution of NaCl at 20°C (293K).
70. What liquid is referred to as the “universal solvent”?

71. In what way are dialysis and osmosis similar? In what way are they different?

72. Name two compounds which can more through the membrane used in hemodialysis.

Chemistry

1. The average mass of one atom of iron is 55.85 amu. What is the mass of Avogadro's number of atoms?
 - A) 55.85 centigrams
 - B) 55.85 g
 - C) 55.85 kg
 - D) 55.85 atoms
 - E) 55.85 formula units

2. What is the weight, in grams, of one mole of hydrogen atoms? [Use atomic weight: H, 1.01 amu]
 - A) 1.01 g
 - B) 2.02 g
 - C) 2.52 g
 - D) 6.02×10^{23} g
 - E) 1.81×10^{24} g

3. How many grams of sulfur make up 3.01 mol of sulfur? [Use atomic weight: S, 32.06 amu]
 - A) 1.81×10^{24} g
 - B) 32.06 g
 - C) 3.01 g
 - D) 0.150 g
 - E) 96.5 g

4. How many moles are there in one ounce (28.4 g) of pure gold? [Use atomic weight: au, 197.0 amu]
 - A) 1.97×10^2 mol
 - B) 6.94 mol
 - C) 0.144 mol
 - D) 0.0721 mol
 - E) 5.08×10^{-3} mol

5. How many atoms of sulfur are present in 155 g of sulfur? [Use atomic weight: S, 32.06 amu]
 - A) 2.91×10^{24} atoms
 - B) 6.02×10^{23} atoms
 - C) 3.01×10^{23} atoms
 - D) 2.91×10^{23} atoms
 - E) 2.01×10^{23} atoms

6. How many iron atoms are present in 3.01 mol of iron?
- A) 1.81×10^{23}
 - B) 6.02×10^{23}
 - C) 3.01×10^{23}
 - D) 1.81×10^{24}
 - E) 58.5
7. What is the formula weight of carbon dioxide? [Use atomic weights: C, 12.01 amu; O, 16.00 amu]
- A) 28.01 amu
 - B) 28.01 g
 - C) 44.01 amu
 - D) 44.01 g
 - E) 44.01 mol
8. Aspirin is the common name for acetyl salicylic acid, $C_9H_8O_4$. A tablet has 0.325g of aspirin. How many moles is this? [Use formula weight: aspirin, 180.2 amu]
- A) 1.80×10^{-6} mol
 - B) 1.80×10^{-3} mol
 - C) 0.554 mol
 - D) 554 mol
 - E) 1.96×10^{23} mol
9. How many grams are there in 0.0200 mol of nicotine, a yellow liquid? [Use formula weight: nicotine, 162.2 amu]
- A) 1.23×10^{-4} g
 - B) 0.308 g
 - C) 3.24 g
 - D) 32.4 g
 - E) 8.11×10^3 g
10. How many molecules are there in 0.0200 mol of nicotine, a yellow liquid?
- A) 0.0400 molecules
 - B) 1.20×10^{22} molecules
 - C) 2.41×10^{22} molecules
 - D) 6.02×10^{23} molecules
 - E) 1.20×10^{23} molecules

11. How many molecules are there in 0.325 g of aspirin? [Use formula weight: aspirin, 180.2 amu]

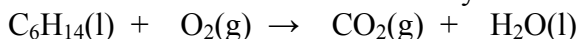
- A) 1.09×10^{21} molecules
- B) 2.17×10^{21} molecules
- C) 1.96×10^{23} molecules
- D) 3.91×10^{23} molecules
- E) 1.85×10^{24} molecules

12. What number will be found in front of "Al" when the following equation is balanced with smallest whole number coefficients?



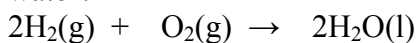
- A) 0.5 mol
- B) 1 mol
- C) 2 mol
- D) 4 mol
- E) 8 mol

13. Which of the choices is the correctly balanced form of the following equation?



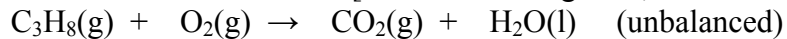
- A) $\text{C}_6\text{H}_{14}(\text{l}) + 13\text{O}_2(\text{g}) \rightarrow 6\text{CO}_2(\text{g}) + 7\text{H}_2\text{O}(\text{l})$
- B) $\text{C}_6\text{H}_{14}(\text{l}) + 19\text{O}(\text{g}) \rightarrow 6\text{CO}_2(\text{g}) + 7\text{H}_2\text{O}(\text{l})$
- C) $\text{C}_6\text{H}_{14}(\text{l}) + 19\text{O}_2(\text{g}) \rightarrow 6\text{CO}_2(\text{g}) + 7\text{H}_2\text{O}(\text{l})$
- D) $2\text{C}_6\text{H}_{14}(\text{l}) + 19\text{O}_2(\text{g}) \rightarrow 12\text{CO}_2(\text{g}) + 14\text{H}_2\text{O}(\text{l})$
- E) $\text{C}_{12}\text{H}_{28}(\text{l}) + 38\text{O}(\text{g}) \rightarrow 12\text{CO}_2(\text{g}) + 14\text{H}_2\text{O}(\text{l})$

14. How many moles of hydrogen gas are needed to react with oxygen to form one mole of water?



- A) 0.5 mol
- B) 1 mol
- C) 2 mol
- D) 4 mol
- E) 6 mol

15. Calculate the mass in grams of oxygen needed to react with 1.000 mol of C_3H_8 to form carbon dioxide and water. [Use atomic weight: O, 16.00 amu]



- A) 80.0 g
- B) 32.00 g
- C) 40.0 g
- D) 320.0 g
- E) 160.0 g

16. Iron reacts with oxygen to form iron(III) oxide (Fe_2O_3). How many grams of product will be formed from 125.5 g of Fe? [Use atomic weights: Fe, 55.85 amu; O, 16.00 amu]

- A) 59.8 g
- B) 178.8 g
- C) 89.7 g
- D) 358.5 g
- E) 159.7 g

17. Consider the hypothetical reaction: $3A_2 + 2B \rightarrow C + 2D$

How many moles of D can be formed from 5.0 mol of A_2 and excess B?

- A) 1.7 mol
- B) 3.3 mol
- C) 6.7 mol
- D) 7.5 mol
- E) 10 mol

Chemistry

- Express one standard atmosphere of pressure in units of mm Hg.
 - 14.7 mm Hg
 - 380 mm Hg
 - 760 mm Hg
 - 0.333 mm Hg
 - 30 mm Hg

- Which one of the following pressures is NOT equal to one atmosphere?
 - 76 cm Hg
 - 14.7 lb/in²
 - 30 inches of Hg
 - 101kPa
 - 76 torr

- A sample of oxygen occupies 1.00 L. If the temperature remains constant, and the pressure on the oxygen is tripled, what is the new volume?
 - 3.00 L
 - 1.50 L
 - 0.667 L
 - 0.500 L
 - 0.333 L

- A given mass of oxygen at room temperature occupies a volume of 500.0 mL at 1.50 atm pressure. What pressure must be applied to compress the gas to a volume of only 150.0 mL?
 - 500 atm
 - 150 atm
 - 5.00 atm
 - 1.50 atm
 - 0.500 atm

- A balloon filled with helium has a volume of 1.00×10^3 L at 20°C. What would be the balloon's volume at 30°C, if the pressure surrounding the balloon remains constant?
 - 6.7×10^2 L
 - 9.70×10^2 L
 - 1.03×10^3 L
 - 1.11×10^3 L
 - 1.50×10^3 L

6. What is the volume occupied by one mole of helium at 0°C and 1 atm pressure?
- A) 1.0 L
 - B) 22.4 L
 - C) 4.0 L
 - D) 40.0 L
 - E) 12.2 L
7. A helium filled weather balloon is launched from the ground where the pressure is 752 mmHg and the temperature is 21°C. Under these conditions its volume is 75.01. When it has climbed to an altitude where the pressure is 89 mmHg and the temperature is 0°C, what is its volume?
- A) 0.00 L
 - B) 8.24 L
 - C) 9.56 L
 - D) 588 L
 - E) 682 L
8. How many moles of gas are there in a gas-filled balloon which has a volume of 67.0 L at a pressure of 742 mmHg and a temperature of 25.0°C?
- A) 2.24 mol
 - B) 2.67 mol
 - C) 2.81 mol
 - D) 31.9 mol
 - E) 1.71×10^3 mol
9. Carbon dioxide acts as a greenhouse gas by
- A) absorbing visible radiation
 - B) absorbing ultraviolet radiation
 - C) absorbing infrared radiation
 - D) storing solar energy
 - E) trapping sunlight during photosynthesis
10. Which of the following statements conflicts with the kinetic molecular theory of gases?
- A) There are no forces between gas particles.
 - B) Gas particles occupy a negligible volume compared with the volume of their container.
 - C) The average kinetic energy of the gas particles is proportional to the absolute temperature.
 - D) Gas particles lose energy only when they collide with the walls of the container.
 - E) Gas particles are in constant, random motion.

11. What quantity is directly proportional to the kinetic energy of the particles in a gas?
- A) distance between molecules
 - B) absolute temperature
 - C) atomic mass
 - D) formula mass
 - E) volume of the individual particles
12. Of the following gases, which will behave most like an ideal gas?
- A) H_2
 - B) HF
 - C) NH_3
 - D) CH_3Cl
 - E) CO
13. At the membrane barrier in lung tissue between the blood and the surrounding atmosphere, what is the relationship between the partial pressure of atmospheric oxygen to that of the oxygen present in the blood?
- A) equal
 - B) proportional
 - C) zero
 - D) lower
 - E) higher
14. Surface tension
- A) increases with increasing temperature
 - B) is unaffected by temperature
 - C) is higher for nonpolar substances than for polar ones
 - D) is lowered by surfactants
 - E) is the same as viscosity
15. Which is the term that describes a liquid changing to a vapor at a temperature less than its boiling point?
- A) evaporation
 - B) sublimation
 - C) dissociation
 - D) condensation
 - E) supercooling
16. The density of a gas is proportional to its molecular weight.

17. Approximately 99% of the total pressure of dry air is due to molecules of N_2 and O_2 .
18. The boiling point of a liquid is dependent on the atmospheric pressure.
19. Polar compounds generally have higher boiling points than nonpolar compounds of similar molecular weight.
20. Ionic compounds tend to have higher melting points than molecular compounds.
21. What device is used to measure atmospheric pressure?
22. The pascal (Pa) is a unit for expressing what quantity?
23. What experimental quantity measures force per unit area?
24. State Boyle's Law.
25. State Charles's Law.
26. State Avogadro's Law.
27. Consider 1.00 L of air in a patient's lungs at 37.0°C and 1.00 atm pressure. What volume would this air occupy if it were at 25.0°C under a pressure of 5.00×10^2 atm (a typical pressure in a compressed air cylinder)?
28. If a gas sample is at STP (Standard Temperature and Pressure), what are its temperature and pressure?
29. Calculate the density of oxygen gas (O_2) at STP, in g/L. [Use molar mass: O_2 , 32.0 g/mol]

30. What is the volume (L) occupied by a mole of an ideal gas, if the pressure is 626 mmHg and the temperature is 25.0°C?
31. A gas sample is prepared in which the components have the following partial pressures: nitrogen, 555 mmHg; oxygen, 149 mmHg; water vapor, 13 mmHg; argon, 7 mmHg. What is the total pressure of this mixture?
32. What is meant by viscosity?
33. What experimental quantity is a measure of the attractive forces between molecules at the surface of a liquid?
34. What process is responsible for the formation of dew on the grass early in the morning?
35. How can pure water be made to boil at a temperature above 100°C?
36. What are the general structural requirements for a compound to display hydrogen bonding?
37. Why is hydrogen bonding more extensive in water than in hydrogen fluoride?
38. Name the four main types of crystalline solid, and give an example of each.
39. Which two states of matter are the least compressible?

Chemistry

1. Classify the following reaction as precipitation, acid-base or oxidation-reduction:
$$\text{Ce}^{4+}(\text{aq}) + \text{Fe}^{2+}(\text{aq}) \rightarrow \text{Ce}^{3+}(\text{aq}) + \text{Fe}^{3+}(\text{aq})$$
2. Complete the products and balance the following equation for an acid-base reaction:
$$\text{HCl}(\text{aq}) + \text{KOH}(\text{aq}) \rightarrow$$
3. Name the two products formed when octane (C_8H_{18}) burns completely in excess oxygen gas.
4. Classify the following reaction as decomposition, combination, single-replacement or double-replacement:
$$\text{Mg}(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{MgCl}_2(\text{aq}) + \text{H}_2(\text{g})$$
5. Outline a simple method to distinguish between a beaker containing a true solution and one containing a colloidal suspension.
6. A solution contains 1.65g of NaOH in a total volume of 150.0mL. What is its concentration expressed as % (W/V)?
7. Define “molarity”.
8. How many moles of KNO_3 are contained in one liter of 0.2M KNO_3 solution?
9. What is the molarity of 50.0mL of a 0.660M NaOH solution after it has been diluted to 450.0mL?
10. In one sentence, explain what is meant by a colligative property.
11. An aqueous solution is warmed from 20°C to 30°C. Does this change in temperature affect either the molarity or the molality of the solution? Explain.

12. What type of membranes allow solvent molecules to pass through but do not allow solute molecules to pass through?
13. Describe what happens when red blood cells are placed in a hypertonic solution. Give the name of the process which occurs.
14. Calculate the osmolarity of a $2.0 \times 10^{-3} M$ Na_3PO_4 solution. Na_3PO_4 is an ionic compound and produces an electrolytic solution.
15. Calculate the osmotic pressure of a $6.0 \times 10^{-2} M$ solution of NaCl at 20°C (293K).
16. What liquid is referred to as the “universal solvent”?
17. In what way are dialysis and osmosis similar? In what way are they different?
18. Which one of the following processes cannot separate solute from solvent in a true solution?
- A) chromatography
 - B) filtration
 - C) boiling
 - D) adsorption
 - E) evaporation
19. The solubility of gases in liquids is highest at
- A) low temperature and low pressure
 - B) low temperature and high pressure
 - C) high temperature and low pressure
 - D) high temperature and high pressure
 - E) high pressure; temperature is immaterial
20. Calculate the concentration (% W/V) of NaCl solution that was made by dissolving 15.0g of sodium chloride in enough water to make 300.0mL of solution.
- A) 50.0% (W/V)
 - B) 0.0500% (W/V)
 - C) 0.356% (W/V)
 - D) 35.6% (W/V)
 - E) 5.00% (W/V)

21. How many milliliters of $0.1250M$ KCl solution contain 2.330g of KCl?
[Use formula weight: KCl, 74.55 amu]
- A) 20.50 mL
 - B) 26.95 mL
 - C) 25.00 mL
 - D) 1.500 mL
 - E) 250.0 mL
22. Which one of the following is NOT a colligative property of a solution?
- A) vapor pressure lowering
 - B) density
 - C) boiling point elevation
 - D) freezing point depression
 - E) osmotic pressure
23. If the concentration of Mg^{2+} in solution is $3.0 \times 10^{-3}M$, what is its concentration expressed in meq/L?
- A) 6.0 meq/L
 - B) 3.0 meq/L
 - C) 1.5 meq/L
 - D) 6.0×10^{-6} meq/L
 - E) 1.5×10^{-6} meq/L
24. Nitrates and acetates are generally soluble.
25. The usual products of an acid-base reaction are a salt and water.
26. In normal room lighting, the eye cannot distinguish a true solution from a colloidal one.
27. The solubility of gases in liquids increases with increasing temperature.
28. Colligative properties depend only on the concentration of solute particles, not on their identity.
29. Osmosis is the process that regulates the sodium/potassium ratio in living cells.

Chemistry

1. Complete in not more than ten words: Thermodynamics deals with...
2. Name, in order, the three thermodynamic quantities represented by H , S , and G (as in ΔH° , ΔS° , and ΔG°)
3. Explain what is meant by “system” and surroundings” in thermodynamics.
4. What is another name for the first law of thermodynamics?
5. What is the name and symbol of the thermodynamic quantity (function) which is equal to the heat absorbed or liberated in a chemical reaction at constant pressure?
6. How many small calories are equivalent to one nutritional Calorie?
7. What process is the main source of energy in the human body?
8. In one sentence, define a catalyst.
9. What is the name given to a biological catalyst?
10. Write the equilibrium constant expression for the reaction:
$$2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \leftrightarrow 2\text{NO}_2(\text{g})$$
11. In the industrial synthesis of ammonia, the equilibrium constant expression may be written as:
$$K_{\text{eq}} = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3}$$

Calculate the value of this equilibrium constant, if the equilibrium concentrations in the reaction mixture at 600°C are: $[\text{N}_2] = 0.830 \text{ M}$, $[\text{H}_2] = 2.49 \text{ M}$, $[\text{NH}_3] = 7.62 \text{ M}$
12. State LeChatelier’s principle.

13. When a cold-pack is activated, a chemical reaction occurs and the temperature of the pack contents drops sharply. Which of the following is a correct description of the reaction occurring in the pack?
- A) The reaction is exothermic; $\Delta H^\circ > 0$
 - B) The reaction is exothermic; $\Delta H^\circ < 0$
 - C) The reaction is endothermic; $\Delta H^\circ > 0$
 - D) The reaction is endothermic; $\Delta H^\circ < 0$
 - E) None of the above statements is correct.
14. What is the measure of the randomness or disorder of a chemical system?
- A) energy
 - B) calorimetry
 - C) entropy
 - D) enthalpy
 - E) free energy
15. In general, which state of matter has the lowest entropy?
- A) solid
 - B) liquid
 - C) gas
 - D) plasma
 - E) supercritical fluid
16. Under normal conditions, which of the following would have the highest entropy?
- A) oxygen gas
 - B) liquid water
 - C) ice
 - D) solid sodium chloride
 - E) solid iron
17. A granola bar contains 185 nutritional Calories. How many kilojoules is this?
- A) 7.74×10^{-1} kJ
 - B) 4.42×10^{-2} kJ
 - C) 442 kJ
 - D) 774 kJ
 - E) 1.85×10^5 kJ

18. Ethylene glycol has a specific heat of $0.578 \text{ cal/(g } ^\circ\text{C)}$. If 23.2 g of ethylene glycol absorbs 75.6 cal of heat energy, what will the temperature increase be?

- A) 0.177°C
- B) 1.88°C
- C) 5.64°C
- D) $1.01 \times 10^3^\circ\text{C}$
- E) $3.03 \times 10^3^\circ\text{C}$

19. Which one of the following actions can alter the activation energy of a reaction?

- A) changing the temperature
- B) changing the concentration of reactants
- C) changing the concentration of products
- D) changing the size of the reaction vessel
- E) adding a catalyst

20. One effect of a catalyst being added to a reaction mixture is

- A) to increase the rate of collisions between reactants
- B) to slow down the rate of the reverse reaction
- C) to raise the temperature of the mixture
- D) to provide a new pathway for the reaction
- E) none of the above

21. In kinetics, the order of a reaction

- A) is the inverse of the entropy of the system
- B) is measured experimentally
- C) can be deduced from the balanced equation for the reaction
- D) depends on the rate constant
- E) depends on the concentrations of reactants

22. When a chemical reaction reaches equilibrium

- A) the forward reaction stops
- B) the reverse reaction stops
- C) one of the reactants has been completely consumed
- D) the equilibrium constant has reached its maximum value
- E) the rates of the forward and reverse reactions are equal

23. Consider the reversible reaction: $A(g) \leftrightarrow 2B(g)$
At equilibrium, the concentration of A is 0.381 M and that of B is 0.154 M. What is the value of the equilibrium constant, K_{eq} ?

- A) 0.0622
- B) 0.404
- C) 1.06
- D) 2.47
- E) 16.1

24. Consider the reversible reaction: $N_2(g) + 3H_2(g) \leftrightarrow 2NH_3(g)$

A) $\frac{[2NH_3]}{[N_2][3H_2]}$

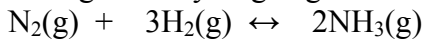
B) $\frac{2[NH_3]}{[N_2]3[H_2]}$

C) $\frac{[NH_3]^2}{[N_2][H_2]^3}$

D) $\frac{[N_2][H_2]^3}{[NH_3]^2}$

E) $\frac{[N_2]3[H_2]}{2[NH_3]}$

25. The reaction below is at equilibrium. Use LeChatelier's principle to predict the effect of adding more hydrogen gas to the equilibrium reaction mixture.



- A) The equilibrium position will remain unchanged.
- B) The equilibrium position will shift to the right.
- C) The equilibrium position will shift to the left.
- D) The equilibrium constant will increase.
- E) All of the nitrogen gas will be used up.

26. Exothermic reactions are often, but not always, spontaneous.

27. Endothermic reactions are never spontaneous.
28. A reaction that leads to an increase in the entropy of the system is always spontaneous.
29. A reaction that leads to a decrease in the enthalpy of the system is always spontaneous.
30. A reaction that leads to a decrease in the free energy of the system is always spontaneous.
31. According to the second law of thermodynamics, a system and its surroundings spontaneously tend to towards increasing order.
32. In general, a liquid state will have a higher entropy than a solid state.
33. A catalyst increases the equilibrium constant for a reaction.

Chemistry

1. What is another name for a protonated water molecule?
2. Explain briefly what is meant by a “conjugate acid-base pair” in the Bronsted-Lowry theory. Give an example of such a pair.
3. What is the conjugate acid of NH_3 ?
4. What is the conjugate base of HNO_3 ?
5. What is the fundamental difference between a strong acid and a weak acid?
6. Complete the following equation for the dissociation of acetic acid in water, so as to illustrate unambiguously that acetic acid is a weak acid.
 $\text{CH}_3\text{COOH}(\text{aq}) + \text{H}_2\text{O}(\text{l})$
7. What is meant by the “auto-ionization of water”?
8. In a chemical context, what is meant by the term “neutralization”?
9. What do we call a substance which is used to show changes in pH by its change in color?
10. What is the name of the process in which we carefully measure the volume of a solution of known concentration needed to neutralize a solution of unknown concentration?
11. What is meant by a diprotic acid? Give an example.
12. What is meant by the “buffer capacity” of a solution?
13. Briefly explain what is meant by a “voltaic cell” and an “electrolytic cell” emphasizing how they relate to one another.

14. In a voltaic cell, oxidation occurs at the _____ while reduction occurs at the _____.

15. Which one of the following is NOT generally true of a base?

- A) tastes bitter
- B) feels slippery
- C) increases $[H^+]$ in water
- D) is corrosive
- E) causes many metal ions to precipitate

16. What is the hydronium ion concentration of pure water at 25°C?

- A) 7.0×10^7
- B) 7.0×10^{-7}
- C) 1.0×10^7
- D) 1.0×10^{-7}
- E) 1.0×10^{-14}

17. What is the hydronium ion concentration of a solution with a pH of 6.0?

- A) 1×10^6
- B) 1×10^{-6}
- C) 6×10^1
- D) 6×10^{-1}
- E) 6×10^{-14}

18. What is the pH of a $1.0 \times 10^{-4} M$ solution of KOH?

- A) 4.00
- B) 6.00
- C) 7.00
- D) 10.00
- E) 14.00

19. What is the pH of a solution that has $[H_3O^+] = 6.0 \times 10^{-3} M$?

- A) 1.7×10^{-12}
- B) 2.22
- C) 3.60
- D) 5.12
- E) 11.78

20. If 14.8mL of 0.100M NaOH solution are needed to react with 25.0mL of a unknown HCl solution, what is the molar concentration of the HCl solution?

- A) 0.592 M
- B) 5.87 M
- C) 0.0592 M
- D) 0.0692 M
- E) 1.25 M

21. The reaction of an acid with a base will produce a salt. What is the other product?

- A) a hydronium ion
- B) water
- C) a buffer
- D) a metal
- E) a hydroxide ion

22. What particle is transferred from one reactant to another in oxidation-reduction reactions?

- A) proton
- B) electron
- C) hydronium ion
- D) hydrogen ion
- E) hydride ion

23. Which one of the following substances is not a good oxidizing agent?

- A) hydrogen peroxide
- B) sodium hypochlorite
- C) chlorine
- D) ozone
- E) carbon dioxide

24. All strong bases are metal hydroxides.

25. A pH of 0 is considered neutral.

26. Carbonic acid and the bicarbonate ion form the main buffer system found in blood.

27. Emphysema can cause blood acidosis.

28. The conversion of ferrous ion to ferric ion is a reduction reaction.
29. The conversion of Fe^{2+} to Fe^{3+} is a reduction reaction.
30. Metals such as sodium are good reducing agents.
31. A rechargeable battery can operate either as a voltaic cell or as an electrolytic cell.
32. In electrolysis, electrical energy is used to drive a non-spontaneous chemical reaction.
33. Which theory describes an acid as a proton donor and a base as a proton acceptor?

Chemistry

1. In the general symbol A_ZX for a nucleus, which of the three letters represents the atomic number?
2. What is the mass number of an alpha particle?
3. What is the mass number of a beta particle?
4. Write the complete symbol for a beta particle, in the form A_ZX .
5. Which type of radiation emitted by radioactive nuclei is similar in mass to a helium atom?
6. Which type of radiation emitted by radioactive nuclei is negatively charged?
7. Which type of radiation emitted by radioactive nuclei is a form of electromagnetic radiation?
8. Which type of radiation emitted by radioactive nuclei has no mass?
9. Which type of radiation emitted by radioactive nuclei is the most penetrating?
10. Which type of radiation emitted by radioactive nuclei is the slowest moving and least penetrating?
11. What may happen to a molecule if it is hit by gamma radiation?
12. What product nucleus would result from the alpha decay of radium-226?
13. The isotope ${}^{226}_{88}\text{Ra}$ decays to ${}^{222}_{86}\text{Rn}$ by emitting radiation. Name the type of radiation.

14. The isotope $^{63}_{28}\text{Ni}$ decays to $^{63}_{29}\text{Cu}$ by emitting radiation. Name the type of radiation.
15. The isotope $^{99\text{m}}_{43}\text{Tc}$ decays to $^{99}_{43}\text{Tc}$ by emitting radiation. Name the type of radiation.
16. Give the complete nuclear symbol for the isotope formed when the isotope $^{16}_7\text{N}$ undergoes beta decay.
17. Give the complete nuclear symbol for X in the following equation for radioactive decay.
 $^{14}_6\text{C} \rightarrow ^{14}_7\text{N} + \text{X}$
18. Give the complete nuclear symbol for X in the following equation for radioactive decay.
 $^{238}_{92}\text{U} \rightarrow \text{X} + ^4_2\text{He}$
19. What term is used to describe a radioactive isotope which decays by emitting only a gamma ray?
20. What is meant by the “binding energy” of a nucleus?
21. What fraction of the initial amount of a radioactive isotope still remains after four half-lives?
22. The half-life of tritium (^3_1H) is 12 years. How long does it take for 16.0ng of tritium to decay to the point where 2.0ng remains?
23. In Einstein’s equation, $E = mc^2$, what do E, m, and c represent?
24. What is the nuclear process that produces energy in commercial nuclear power plants?
25. What kind of reactor produces its own fuel in the process of providing electrical energy?
26. What is the identity of the radioactive isotope involved in radiocarbon dating?

27. In what important way do cancer cells differ from normal cells?
28. What term is used to describe radioactive substances which are used as probes to study internal organs?
29. In what part of the body does iodine tend to concentrate?
30. Name any two radioactive isotopes commonly used in nuclear medicine.
31. What type of disease can be conveniently studied using xenon-133 as a tracer?
32. What device uses magnetic and electric fields to create high-energy charged particles?
33. What is the term that describes the amount of radiation attributable to our everyday surroundings?
34. What is a film badge?
35. Which radioactive element is found in some indoor air?
36. What term represents the dosage of toxic material needed to kill 50% of the exposed population in 30 days?
37. The isotope iodine-131 is used in studies of the
- A) heart
 - B) lung
 - C) liver
 - D) thyroid
 - E) kidney
38. The sun's source of energy is nuclear fission.

Answer Key

1. A
2. E
3. D
4. C
5. E
6. C
7. A
8. C
9. E
10. E
11. C
12. B
13. C
14. B
15. D
16. B
17. B
18. C
19. D
20. False

Answer Key

1. E
2. E
3. B
4. C
5. A
6. B
7. D
8. A
9. E
10. D
11. True
12. False
13. False
14. True
15. False
16. True
17. True

Answer Key

1. A
2. B
3. E
4. B
5. C
6. C
7. D
8. A
9. B
10. E
11. C
12. B
13. C
14. B
15. E
16. D
17. E
18. A
19. D
20. E

Answer Key

1. MOLECULE
2. six
3. ionic bonding and covalent bonding
4. metal, nonmetal
5. MOLECULE
6. a shared pair of electrons
7. In each case, the bond consists of an electron pair shared between the bonded atoms. The difference is that the sharing is unequal in the case of the polar covalent bond, equal in a nonpolar covalent bond.
8. The atom has a strong attraction for shared electron pairs (electrons in covalent bonds).
9. bottom left
10. Pauling
11. crystal lattice
12. Na_2S
13. Ba_3N_2
14. iron(II) ion
15. lower positive charge
16. polyatomic
17. SO_4^{2-}
18. hydrogen carbonate
19. ammonium
20. sodium phosphate
21. copper(I) oxide
22. Na_2CO_3
23. covalent
24. covalent
25. SO_3
26. NH_3
27. carbon tetrachloride
28. dinitrogen pentoxide
29. boiling point
30. amorphous
31. electrolyte
32. nonelectrolyte
33. covalent
34. ten
35. MOLECULE
36. MOLECULE
37. The structure shows 26 valence electrons, but there should only be 24.
38. MOLECULE
39. bond energy
40. bond length
41. Valence Shell Electron Pair Repulsion
42. 120°

43. bent or angular
44. If the molecule were trigonal planar, the symmetry would result in a nonpolar molecule.
The centers of positive and negative charge would coincide.
45. C
46. C
47. D
48. C
49. D
50. E
51. C
52. A
53. B
54. B
55. E
56. E
57. C
58. D
59. B
60. E
61. A
62. C
63. D
64. valence electrons

Answer Key

1. B
2. A
3. E
4. C
5. A
6. D
7. C
8. B
9. C
10. B
11. A
12. D
13. D
14. D
15. B
16. E
17. B
18. B

Answer Key

1. C
2. E
3. E
4. C
5. C
6. B
7. D
8. B
9. C
10. D
11. B
12. A
13. E
14. D
15. A
16. True
17. True
18. True
19. True
20. True
21. barometer
22. pressure
23. pressure
24. The volume of a gas is inversely proportional to the pressure, if the number of moles (or mass) and the temperature of the gas are kept constant.
25. The volume of a gas is directly proportional to the absolute temperature, if the number of moles (or mass) and the pressure of the gas are kept constant.
26. Equal volumes of any ideal gas, at the same temperature and pressure, contain the same number of moles.
27. $1.92 \times 10^{-3}\text{L}$
28. temperature = 0.0°C or 273K , pressure = 1.00atm
29. 1.43 g/L
30. 29.7 L
31. 724 mmHg
32. It is the resistance of a liquid to flow.
33. surface tension
34. condensation
35. by raising the pressure to more than one atmosphere
36. It must have hydrogen atoms bonded to small, electronegative atoms such as N, O, or F.

37. Each water molecule has two δ^+ sites and two δ^- sites, all of which can be used for hydrogen bonding. The hydrogen fluoride has three δ^- but only one δ^+ site, and the shortage of the latter limits the number of hydrogen bonds which can form per molecule to half the number in water.
38. ionic solid, NaCl; covalent solid, diamond; molecular solid, ice; metallic solid, iron
39. solids and liquids

Answer Key

1. oxidation-reduction
2. $\text{HCl(aq)} + \text{KOH(aq)} \rightarrow \text{KCl(aq)} + \text{H}_2\text{O(l)}$
3. carbon dioxide and water
4. single-replacement
5. Direct a narrow beam of light horizontally through the two beakers. The colloidal suspension will scatter light (the Tyndall effect) making the beam visible as it passes through; the true solution will show no scattering.
6. 1.10% (W/V)
7. It is the concentration of a solution expressed as the number of moles of solute per liter of solution.
8. 0.2 mol
9. 0.0733 M
10. It is a solution property which depends on the concentration of solute particles rather than on their identity.
11. The volume of the solution is likely to increase, causing a decrease in the molarity. Since mass is not affected by temperature, the molality will stay the same.
12. semipermeable
13. The cells lose water, by osmosis, to the hypertonic solution, and they collapse. The process is known as crenation.
14. 8.0×10^{-3} Osm
15. 2.9 atm
16. Water
17. Both involve the selective movement of small molecules through a membrane, from a solution of high concentration of those molecules to one of lower concentration. They differ in that osmosis involves only movement of solvent (water) molecules, whereas in dialysis, solute molecules can also pass through the membrane.
18. B
19. B
20. E
21. E
22. B
23. A
24. True
25. True
26. True
27. False
28. True
29. False

Answer Key

1. energy changes in chemical reactions
2. enthalpy, entropy, free energy
3. The system is the reaction or process being studied, the surroundings are the remainder of the universe.
4. law of conservation of energy
5. enthalpy change, ΔH°
6. 1000 cal
7. combustion of carbohydrates
8. It is a substance which speeds up a chemical reaction without being consumed to the process.
9. enzyme
10. $K_{\text{eq}} = \frac{[\text{NO}_2]^2}{[\text{NO}]^2[\text{O}_2]}$
11. 4.53
12. If a stress is applied to a system at equilibrium, the equilibrium will shift in such a way as to minimize that stress.
13. C
14. C
15. A
16. A
17. D
18. C
19. E
20. D
21. B
22. E
23. A
24. C
25. B
26. True
27. False
28. False
29. False
30. True
31. False
32. True
33. False

Answer Key

1. hydronium ion
2. According to this theory, an acid donates a proton and in the process becomes a base. Thus, a conjugate acid-base pair is a pair of species which differ by one proton. e.g., H₂O and OH⁻
3. NH₄⁺
4. NO₃
5. A strong acid dissociates completely in solution. A weak acid dissociates only partially, forming relatively fewer hydronium ions than a strong acid.
6. CH₃COOH(aq) + H₂O(l) ↔ H₃O⁺(aq) + CH₃COO⁻(aq)
7. It is the transfer of a proton from one water molecule to another, producing a hydronium ion and a hydroxide ion.
8. The reaction between an acid and a base.
9. an indicator
10. titration
11. A diprotic acid is capable of donating two protons. e.g., H₂SO₄
12. Buffer capacity refers to the amount of added acid or base which the buffer solution can neutralize without undergoing a substantial change in pH.
13. A voltaic cell utilizes a spontaneous oxidation-reduction reaction to produce electrical energy; an electrolytic cell reverses this process, using electrical energy to drive a non-spontaneous oxidation-reduction reaction.
14. anode, cathode
15. C
16. D
17. B
18. D
19. B
20. C
21. B
22. B
23. E
24. True
25. False
26. True
27. True
28. False
29. False
30. True
31. True
32. True
33. Bronsted-Lowry theory

Answer Key

1. Z
2. 4
3. 0
4. ${}^0_{-1}\text{e}$
5. alpha particle
6. beta particle
7. gamma ray
8. gamma ray
9. gamma ray
10. alpha particle
11. loss of electrons, or ionization
12. radon-222
13. alpha particle
14. beta particle
15. gamma ray
16. ${}^{16}_8\text{O}$
17. ${}^0_{-1}\text{e}$
18. ${}^{234}_{90}\text{Th}$
19. metastable isotope
20. It is the energy responsible for holding the protons and neutrons together in the nucleus.
21. one sixteenth (0.061)
22. 36 years
23. E = energy; m = mass; c = speed of light
24. fission
25. breeder reactor
26. C-14 or carbon -14
27. They undergo much more rapid cell division.
28. tracers or radioactive tracers
29. thyroid gland
30. (from) iodine-131, technetium-99m, thallium-201, xenon-133, barium-131, chromium-51
31. pulmonary disease
32. particle accelerator
33. background level of radiation
34. It is a badge worn by radiation workers, containing a film which measures their cumulative radiation does.
35. radon
36. LD₅₀
37. D
38. False