1 A - 1 Atom Theory  (7 Questions)
• Atom structure theory developed over the last 200 years:
  - A hard shell (Dalton)
  - Contain an evenly mix of protons and electrons which are deflected by and electric fields (Thomson)
  - Has a dense positive nucleus but is mostly empty space (Rutherford -Gold foil)
  - Electrons circle the nucleus like planets circling the sun (Bohr)
  - Electrons are found in regions known as orbitals. Wave-mechanical - Cloud theory

1. Which sequence represents a correct order of historical developments leading to the modern model of the atom?
   (1) the atom is a hard sphere —> most of the atom is empty space —> electrons exist in orbitals outside the nucleus
   (2) the atom is a hard sphere —> electrons exist in orbitals outside the nucleus —> most of the atom is empty space
   (3) most of the atom is empty space —> electrons exist in orbitals outside the nucleus —> the atom is a hard sphere
   (4) most of the atom is empty space —> the atom is a hard sphere —> electrons exist in orbitals outside the nucleus

2. The modern model of the atom is based on the work of
   (1) one scientist over a short period of time
   (2) one scientist over a long period of time
   (3) many scientists over a short period of time
   (4) many scientists over a long period of time

3. In the late 1800s, experiments using cathode ray tubes led to the discovery of the
   (1) electron
   (2) neutron
   (3) positron
   (4) proton

1A-2 Gold Foil  (18 Questions)
• The Rutherford gold foil experiment suggested that the atom is mostly empty space with a heavy dense positive nucleus that would repel or deflect positively charged alpha particles. (Remember, like charges repel one another)

4. Which conclusion was a direct result of the gold foil experiment?
   (1) An atom is mostly empty space with a dense, positively charged nucleus.
   (2) An atom is composed of at least three types of subatomic particles.
   (3) An electron has a positive charge and is located inside the nucleus.
   (4) An electron has properties of both waves and particles.

5. State one conclusion about atomic structure based on the observation that almost all alpha particles passed straight through the foil. [1]

6. Explain, in terms of charged particles, why some of the alpha particles were deflected. [1]

1 A - 3 Orbitals  (10 Questions)
• The modern wave-mechanical or cloud theory states that electrons are most likely found surrounding the nucleus in regions called orbitals.

7. In the wave-mechanical model of the atom, orbitals are regions of the most probable locations of
   (1) protons
   (2) positrons
   (3) neutrons
   (4) electrons

8. In the electron cloud model of the atom, an orbital is defined as the most probable
   (1) charge of an electron
   (2) conductivity of an electron
   (3) location of an electron
   (4) mass of an electron
1B-1  Subatomic Particles  (34 Questions)

- Atoms have three subatomic particles:

<table>
<thead>
<tr>
<th>Subatomic Particle</th>
<th>Location</th>
<th>Charge</th>
<th>Mass (amu)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Proton</td>
<td>Nucleus</td>
<td>+1</td>
<td>1</td>
</tr>
<tr>
<td>Neutron</td>
<td>Nucleus</td>
<td>0</td>
<td>1</td>
</tr>
<tr>
<td>Electron</td>
<td>Orbitals</td>
<td>-1</td>
<td>0 (1/1836)</td>
</tr>
</tbody>
</table>

9. A proton has a charge that is opposite the charge of
- (1) an alpha particle
- (2) a neutron
- (3) an electron
- (4) a positron

10. Which total mass is the smallest?
- (1) the mass of 2 electrons
- (2) the mass of 2 neutrons
- (3) the mass of 1 electron plus the mass of 1 proton
- (4) the mass of 1 neutron plus the mass of 1 electron

11. The mass of 12 protons is approximately equal to
- (1) 1 atomic mass unit
- (2) 12 atomic mass units
- (3) the mass of 1 electron
- (4) the mass of 12 electrons

12. Which subatomic particles are located in the nucleus of an He-4 atom?
- (1) electrons and neutrons
- (2) electrons and protons
- (3) neutrons and protons
- (4) neutrons, protons, and electrons

13. Which particle has a mass that is approximately the same as the mass of a proton?
- (1) an alpha particle
- (2) a beta particle
- (3) a neutron
- (4) a positron

14. A neutron has a charge of
- (1) +1
- (2) +2
- (3) 0
- (4) -1

1B-2  Atomic Number  (14 Questions)

- All atoms of the same element must have the same number of protons which is equal to the atomic number. Same atomic number = same element, different atomic number = different elements. (See table S or the periodic table for atomic number.)

15. Each diagram below represents the nucleus of a different atom.

Which diagrams represent nuclei of the same element?
- (1) D and E, only
- (2) D, E, and Q
- (3) Q and R, only
- (4) Q, R, and E

16. The atomic number of an atom is always equal to the number of its
- (1) protons, only
- (2) neutrons, only
- (3) protons plus neutrons
- (4) protons plus electrons

17. Which quantity identifies an element?
- (1) atomic number
- (2) mass number
- (3) total number of neutrons in an atom of the element
- (4) total number of valence electrons in an atom of the element

18. The total mass of the protons in an atom of gold-198 is approximately
- (1) 79 atomic mass units
- (2) 119 atomic mass units
- (3) 198 atomic mass units
- (4) 277 atomic mass units

1B-3  Atoms are Neutral  (21 Questions)

- Atoms are neutral because they contain the same number of protons as electrons and the charge of the positive protons will cancel out the charge of the negative electrons. The net charge of the nucleus is equal to the atomic number (number of protons in nucleus). (See table S or the periodic table for atomic number.)

19. What is the charge of the nucleus in an atom of oxygen-17?
- (1) 0
- (2) -2
- (3) +8
- (4) +17

20. An atom of any element must contain
- (1) an equal number of protons and neutrons
- (2) an equal number of protons and electrons
- (3) more electrons than neutrons
- (4) more electrons than protons
21. What is the total number of electrons in an atom of Cu-65? [1]

22. Determine the total number of electrons in the boron atom. [1]

23. Determine the total charge of the boron nucleus. [1]

24. The diagram below represents the nucleus of an atom.

What are the atomic number and mass number of this atom?
(1) The atomic number is 9 and the mass number is 19.
(2) The atomic number is 9 and the mass number is 20.
(3) The atomic number is 11 and the mass number is 19.
(4) The atomic number is 11 and the mass number is 20.

25. Chlorine-37 can be represented as
(1) $^{37}_{17}$Cl
(2) $^{38}_{18}$Cl
(3) $^{35}_{17}$Cl
(4) $^{37}_{18}$Cl

26. The notation for the nuclide $^{137}_{55}$Cs gives information about
(1) mass number, only
(2) atomic number, only
(3) both mass number and atomic number
(4) neither mass number nor atomic number

27. What is the mass number of an atom that has six protons, six electrons, and eight neutrons?
(1) 6
(2) 12
(3) 14
(4) 20

28. What is the total number of neutrons in an atom of an element that has a mass number of 19 and an atomic number of 9?
(1) 9
(2) 10

29. The nucleus of an atom of cobalt-58 contains
(1) 27 protons and 31 neutrons
(2) 27 protons and 32 neutrons
(3) 59 protons and 60 neutrons
(4) 60 protons and 60 neutrons
30. Which diagram represents the nucleus of an atom of ¹⁵⁷ Al?

(1)  ![Diagram 1](image1.png)  
(2)  ![Diagram 2](image2.png)  
(3)  ![Diagram 3](image3.png)  
(4)  ![Diagram 4](image4.png)  

31. Which two particles make up most of the mass of a hydrogen-2 atom?

(1) electron and neutron  
(2) electron and proton  
(3) proton and neutron  
(4) proton and positron

32. The greatest composition by mass in an atom of ¹⁷⁸ O is due to the total mass of its:

(1) electrons  
(2) neutrons  
(3) positrons  
(4) protons

33. Which two particles make up most of the mass of a hydrogen-2 atom?

(1) electrons  
(2) neutrons  
(3) positrons  
(4) protons

34. Which two particles make up most of the mass of a hydrogen-2 atom?

(1) electrons  
(2) neutrons  
(3) positrons  
(4) protons

35. The most common isotope of chromium has a mass number of 52. Which notation represents a different isotope of chromium?

(1) ⁵² Cr  
(2) ⁵⁴ Cr  
(3) ⁵⁶ Cr  
(4) ⁵⁸ Cr

36. Isotopes of an element must have different:

(1) atomic numbers  
(2) mass numbers  
(3) numbers of protons  
(4) numbers of electrons

37. Atoms of different isotopes of the same element differ in their total number of:

(1) electrons  
(2) neutrons  
(3) protons  
(4) valence electrons

38. What information is necessary to determine the atomic mass of the element chlorine?

(1) the atomic mass of each artificially produced isotope of chlorine, only  
(2) the relative abundance of each naturally occurring isotope of chlorine, only  
(3) the atomic mass and the relative abundance of each naturally occurring isotope of chlorine  
(4) the atomic mass and the relative abundance of each naturally occurring and artificially produced isotope of chlorine

39. The atomic mass of titanium is 47.88 atomic mass units. This atomic mass represents the:

(1) total mass of all the protons and neutrons in an atom of Ti  
(2) total mass of all the protons, neutrons, and electrons in an atom of Ti  
(3) weighted average mass of the most abundant isotope of Ti  
(4) weighted average mass of all the naturally occurring isotopes of Ti

**1C-3 Isotopes (19 Questions)**

- Isotopes are atoms of the same element with different number of neutrons. Isotopes always have the same atomic numbers but different mass numbers. ⁴⁰ K and ⁴⁰ K are isotopes as are ¹²⁷ Sn, ¹²⁸ Sn and ¹²⁹ Sn. Oxygen-16 and Nitrogen-16 are not isotopes, but Oxygen-17 and Oxygen-16. Both oxygen-16 and oxygen-17 atomic numbers are 8 and oxygen-17 has 9 neutrons while oxygen-16 has 8 neutrons.

33. The total number of protons, electrons, and neutrons in each of four different atoms are shown in the table below.

<table>
<thead>
<tr>
<th>Atom</th>
<th>Total Number of Protons</th>
<th>Total Number of Electrons</th>
<th>Total Number of Neutrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>6</td>
<td>6</td>
<td>7</td>
</tr>
<tr>
<td>D</td>
<td>6</td>
<td>6</td>
<td>8</td>
</tr>
<tr>
<td>X</td>
<td>7</td>
<td>7</td>
<td>8</td>
</tr>
<tr>
<td>Z</td>
<td>8</td>
<td>8</td>
<td>9</td>
</tr>
</tbody>
</table>

Which two atoms are isotopes of the same element?

(1) A and D  
(2) A and Z  
(3) X and D  
(4) X and Z

34. Which two notations represent different isotopes of the same element?

(1) ⁶ Be and ⁷ Be  
(2) ⁷ Li and ⁶ Li  
(3) ¹⁴ N and ¹⁴ C  
(4) ¹⁵ P and ¹⁶ S

**1D-1 Atomic Mass (9 Questions)**

- The atomic mass of an element is the weighted average of the masses of all naturally occurring isotopes. The percentages of each isotope may be quite different. The atomic mass will usually be closest to the mass number of the isotope with the greatest percentage.

38. What information is necessary to determine the atomic mass of the element chlorine?

(1) the atomic mass of each artificially produced isotope of chlorine, only  
(2) the relative abundance of each naturally occurring isotope of chlorine, only  
(3) the atomic mass and the relative abundance of each naturally occurring isotope of chlorine  
(4) the atomic mass and the relative abundance of each naturally occurring and artificially produced isotope of chlorine

39. The atomic mass of titanium is 47.88 atomic mass units. This atomic mass represents the:

(1) total mass of all the protons and neutrons in an atom of Ti  
(2) total mass of all the protons, neutrons, and electrons in an atom of Ti  
(3) weighted average mass of the most abundant isotope of Ti  
(4) weighted average mass of all the naturally occurring isotopes of Ti
Calculating Atomic Mass  (14 Questions)

• Steps to calculate the atomic mass are:
  - Step 1: Change the percentages into decimals by moving the decimals left by 2 places.
  - Step 1: Multiply the atomic mass of each isotope by its percentage converted to its decimal
  - Step 2: Add all of the products of step 1. That is the answer.

• Problem: Calculate the atomic mass of copper.

Naturally Occurring Isotopes of Copper

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Atomic Mass (atomic mass units, u)</th>
<th>Percent Natural Abundance (%)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cu-63</td>
<td>62.93</td>
<td>69.17</td>
</tr>
<tr>
<td>Cu-65</td>
<td>64.93</td>
<td>30.83</td>
</tr>
</tbody>
</table>

(62.93)(.6917) + (64.93)(.3083) = 63.54

Problem: Calculate the atomic mass of copper.

Base your answer to question 40 on the information below.

Naturally Occurring Isotopes of Sulfur

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Atomic Mass (atomic mass units, u)</th>
<th>Natural Abundance (%)</th>
</tr>
</thead>
<tbody>
<tr>
<td>$^{32}$S</td>
<td>31.97</td>
<td>94.93</td>
</tr>
<tr>
<td>$^{33}$S</td>
<td>32.97</td>
<td>0.76</td>
</tr>
<tr>
<td>$^{34}$S</td>
<td>33.97</td>
<td>4.29</td>
</tr>
<tr>
<td>$^{35}$S</td>
<td>35.97</td>
<td>0.02</td>
</tr>
</tbody>
</table>

Lithium Isotopes

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Atomic Mass (u)</th>
<th>Natural Abundance (%)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li-6</td>
<td>6.02</td>
<td>7.5</td>
</tr>
<tr>
<td>Li-7</td>
<td>7.02</td>
<td>92.5</td>
</tr>
</tbody>
</table>

1D-3 Locating Atomic Mass  (1 Questions)

• Use the periodic table. (If you don't know the symbol, look it up in Table S.)

KEY

<table>
<thead>
<tr>
<th>Atomic Mass</th>
<th>12.011</th>
</tr>
</thead>
<tbody>
<tr>
<td>Symbol</td>
<td>C</td>
</tr>
<tr>
<td>Atomic Number</td>
<td>6</td>
</tr>
<tr>
<td>Electron Configuration</td>
<td>2-4</td>
</tr>
</tbody>
</table>

Selected Oxidation States

Relative atomic masses are based on $^{12}$C = 12 (exact)

Note: Numbers in parentheses are mass numbers of the most stable or common isotope.

42. In which list are the elements arranged in order of increasing atomic mass?

(1) Cl, K, Ar
(2) Fe, Co, Ni
(3) Te, I, Xe
(4) Ne, F, Na
1E-1 Energy Levels (Shells)  (44 Questions)

- The electrons configuration for each element in its ground state can be found in the periodic table. If the atom absorbs energy and becomes excited, electrons will move up energy levels (represented by jumping up an energy level in the Electron Configuration on the periodic table). Example Carbon becomes 2-3-1 (As electrons move up energy levels, they gain or absorb energy. As electrons move down energy levels, they lose energy by emitting bright-line spectra.)
- Since atoms are neutral the atomic number must equal to the number of electrons. (Protons = electrons)
- The maximum number of electrons for the first four energy levels and their normal filling order can be determined by examining the periodic table. Don't memorize - Use the Periodic Table!

<table>
<thead>
<tr>
<th>Energy Level (Shell)</th>
<th># of orbitals</th>
<th>Maximum # of Electrons</th>
<th>Filling Order of Electrons in Energy Levels</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 (n=1)</td>
<td>1</td>
<td>2</td>
<td>1st always</td>
</tr>
<tr>
<td>2 (n=2)</td>
<td>4</td>
<td>8</td>
<td>2nd always</td>
</tr>
<tr>
<td>3 (n=3)</td>
<td>9</td>
<td>18</td>
<td>First 8 in 3rd then 2 in the 4th then remaining 10 in 3rd energy level.</td>
</tr>
<tr>
<td>4 (n=4)</td>
<td>16</td>
<td>32</td>
<td></td>
</tr>
</tbody>
</table>

43. As an electron in an atom moves from the ground state to the excited state, the electron
   (1) gains energy as it moves to a higher energy level
   (2) gains energy as it moves to a lower energy level
   (3) loses energy as it moves to a higher energy level
   (4) loses energy as it moves to a lower energy level

44. An atom of oxygen is in an excited state. When an electron in this atom moves from the third shell to the second shell, energy is
   (1) emitted by the nucleus
   (2) emitted by the electron
   (3) absorbed by the nucleus
   (4) absorbed by the electron

45. Which atom in the ground state has an outermost electron with the most energy?
   (1) Cs
   (2) K
   (3) Li
   (4) Na

46. What is the total number of protons in an atom with the electron configuration 2-8-18-32-18-1?
   (1) 69
   (2) 79
   (3) 118
   (4) 197

47. Write one electron configuration for an atom of silicon in an excited state. [1]

48. A bromine atom in an excited state could have an electron configuration of
   (1) 2-8-18-6
   (2) 2-8-18-7
   (3) 2-8-17-7
   (4) 2-8-17-8

49. Which electron configuration represents an excited state for a potassium atom?
   (1) 2-8-7-1
   (2) 2-8-7-2
   (3) 2-8-8-1
   (4) 2-8-8-2

1E-2 Valence Electrons  (14 Questions)

- The outermost electrons are called the valence electrons. Carbon has 4.

50. How many valence electrons are in an atom of K-42 in the ground state? [1]

51. What is the total number of valence electrons in an atom of germanium in the ground state?
   (1) 8
   (2) 2
   (3) 14
   (4) 4

52. What is the total number of valence electrons in a sulfide ion in the ground state?
   (1) 8
   (2) 2
   (3) 16
   (4) 18

53. Which element has an atom in the ground state with a total of three valence electrons?
   (1) aluminum
   (2) lithium
   (3) phosphorus
   (4) scandium

54. In comparison to an atom of $^1\text{H}$F in the ground state, an atom of $^{12}\text{C}$ in the ground state has
   (1) three fewer neutrons
   (2) three fewer valence electrons
   (3) three more neutrons
   (4) three more valence electrons
When excited atoms emit energy, electrons move down energy levels until they return to their ground state. The energy is emitted as specific bands of light called bright-line spectra. Elements can be identified by their bright-line spectra since every element produces its own specific bright-line spectra.

55. Explain, in terms of both electrons and energy, how the bright-line spectrum of an element is produced. [1]

56. Identify all the elements in the mixture. [1]

57. The diagram below represents the bright-line spectra of four elements and a bright-line spectrum produced by a mixture of two of these elements.

Which two elements are in this mixture?
(1) barium and hydrogen  (3) helium and hydrogen
(2) barium and lithium  (4) helium and lithium

58. During a flame test, ions of a specific metal are heated in the flame of a gas burner. A characteristic color of light is emitted by these ions in the flame when the electrons
(1) gain energy as they return to lower energy levels
(2) gain energy as they move to higher energy levels
(3) emit energy as they return to lower energy levels
(4) emit energy as they move to higher energy levels

59. When the electrons of an excited atom return to a lower energy state, the energy emitted can result in the production of
(1) alpha particles  (3) protons
(2) isotopes  (4) spectra
1 F - 1 Pure Substances  (32 Questions)

- Pure substances have the same composition throughout regardless of the sample. Elements and compounds are pure substances.
  - Elements - They cannot be broken down by ordinary or chemical means. (Note: See table S and/or the periodic table for names, symbols and physical properties.)

- Particle model diagrams (Elements):

<table>
<thead>
<tr>
<th>Four Different Elements</th>
</tr>
</thead>
<tbody>
<tr>
<td>KEY</td>
</tr>
<tr>
<td>Ø = atom of A</td>
</tr>
<tr>
<td>• = atom of X</td>
</tr>
<tr>
<td>O = atom of B</td>
</tr>
<tr>
<td>□ = atom of Z</td>
</tr>
</tbody>
</table>

- Compounds are made up of two or more elements that combined chemically in a specific ratio. Compounds have different physical and chemical properties than the elements that compose them. Compounds can be changed back (separated) into their elements by chemical means.
  - Examples: H₂O (water)-2H:1O, NaCl (salt)-1Na:1Cl and C₆H₁₂O₆ (Sugar)-1C:2H:1O

- Particle model diagrams (Compounds):

<table>
<thead>
<tr>
<th>Four Different Compounds</th>
</tr>
</thead>
<tbody>
<tr>
<td>KEY</td>
</tr>
<tr>
<td>Ø = atom of A</td>
</tr>
<tr>
<td>• = atom of X</td>
</tr>
<tr>
<td>O = atom of B</td>
</tr>
<tr>
<td>□ = atom of Z</td>
</tr>
</tbody>
</table>

- Pure substances can be identified by their physical and chemical properties.
  - Physical properties can be observed without changing the chemical makeup of the substance. Physical properties include luster, density, color, solubility (ability to be dissolved) melting and boiling points.
  - Chemical properties involve reactions that cause chemical change. Reacts, changes, burns, decomposes and forms are all key-words that may indicate chemical change.

Base your answers to questions 60 through 62 on the particle diagrams below, which show atoms and/or molecules in three different samples of matter at STP.

![Particle Diagrams]

60. Which sample represents a pure substance? [1]

Answer: ________________

61. When two atoms of y react with one atom of z, a compound forms. Using the number of atoms shown in sample 2, what is the maximum number of molecules of this compound that can be formed? [1]

Answer: ________________

62. Explain why XXX does not represent a compound. [1]

63. An example of a physical property of an element is the element’s ability to
   (1) react with an acid
   (2) react with oxygen
   (3) form a compound with chlorine
   (4) form an aqueous solution

64. Which of these contains only one substance?
   (1) distilled water
   (2) sugar water
   (3) saltwater
   (4) rainwater

65. Which sample of matter is a single substance?
   (1) air
   (2) ammonia gas
   (3) hydrochloric acid
   (4) salt water

66. Which property could be used to identify a compound in the laboratory?
   (1) mass
   (2) melting point
   (3) temperature
   (4) volume
67. Which particle diagram represents a sample of one compound, only?

Key

- atom of one element
- atom of a different element

(1)  
(2)  
(3)  
(4)  

68. Which substance can be decomposed by chemical means?

(1) tungsten  (3) krypton  
(2) antimony  (4) methane  

69. A large sample of solid calcium sulfate is crushed into smaller pieces for testing. Which two physical properties are the same for both the large sample and one of the smaller pieces?

(1) mass and density  (3) solubility and density  
(2) mass and volume  (4) solubility and volume  

70. Matter that is composed of two or more different elements chemically combined in a fixed proportion is classified as

(1) a compound  (3) a mixture  
(2) an isotope  (4) a solution  

71. At STP, which physical property of aluminum always remains the same from sample to sample?

(1) mass  (3) length  
(2) density  (4) volume  

1F-2 Mixtures (13 Questions)

- Mixtures are made up of two or more substances not chemically combined. Unlike compounds, the substances that make up mixtures can be present in different ratios or percentages. These substances can be separated by physical means such as physically separating by hand, evaporation or drying, using a magnet, and/or dissolving one of the substances.
- Particle model diagrams (Mixtures):

72. Which particle diagram represents a mixture of element X and element Z only?

73. Which formula represents a mixture?

(1) C_6H_{12}O_6(f)  (3) LiCl(aq)  
(2) C_6H_{12}O_6(s)  (4) LiCl(s)  

74. Bronze contains 90 to 95 percent copper and 5 to 10 percent tin. Because these percentages can vary, bronze is classified as

(1) a compound  (3) a mixture  
(2) an element  (4) a substance  

75. Which particle diagram represents a mixture of an element and a compound?

76. Matter is classified as a

(1) substance, only  
(2) substance or as a mixture of substances  
(3) homogenous mixture, only  
(4) homogenous mixture or as a heterogeneous mixture  

77. Which must be a mixture of substances?

(1) solid  (3) gas  
(2) liquid  (4) solution  

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1F-3 Homogeneous vs Heterogeneous (4 Questions)

- Homogeneous matter is uniform in structure or composition throughout. Homogeneous matter include elements, compounds and solutions.
- Solutions are homogeneous mixtures of two or more substances, which may be solids, liquids, gases, or a combination of these. Examples include: coffee, soda, brass, air and any substance dissolved in water indicated by "(aq)" following the substance (NaCl(aq) - salt water).
- Heterogeneous matter is a mixture of dissimilar elements or parts not uniformly mixed or dispersed. Examples include concrete, chocolate chip cookies and vegetable soup.
- Particle model diagrams (Homogeneous vs. heterogeneous):

Base your answers to questions 78 and 79 on the diagram below concerning the classification of matter.

### Classification of Matter

- **Mixtures**
  - Homogeneous
  - Elements
  - Z

### Four Different Materials

<table>
<thead>
<tr>
<th>KEY</th>
<th>2 Elements Homogeneous Mixture</th>
<th>1 Element Homogeneous Pure Substance</th>
<th>1 compound Homogeneous Pure Substance</th>
<th>2 compounds &amp; 2 Elements Heterogeneous Mixture</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>● = atom of A</td>
<td>● = atom of X</td>
<td>● = atom of X</td>
<td>● = atom of Z</td>
</tr>
</tbody>
</table>

80. Which of these terms refers to matter that could be heterogeneous?

(1) element  
(2) mixture  
(3) compound  
(4) solution

Base your answer to question 81 on the information below.

Cold packs are used to treat minor injuries. Some cold packs contain NH₄NO₃(s) and a small packet of water at room temperature before activation. To activate this type of cold pack, the small packet must be broken to mix the water and NH₄NO₃(s). The temperature of this mixture decreases to approximately 2°C and remains at this temperature for 10 to 15 minutes.

81. Identify the type of mixture formed when the NH₄NO₃(s) is completely dissolved in the water. [1]

---

78. What type of mixture is represented by X [1]

79. Explain, in terms of particle arrangement, why NaCl(aq) is a homogeneous mixture. [1]
2A-1 Chemical Symbols (2 Questions)
- First letter always capitalized. If there is a second letter, it is always lower-case.
- If you don't know a symbol, look it up in Table S.

1. What is the total number of different elements present in NH₄NO₃?
   (1) 7 (2) 9 (3) 3 (4) 4

2. What is the total number of oxygen atoms in the formula MgSO₄ •7 H₂O?
   [The • represents seven units of H₂O attached to one unit of MgSO₄.]
   (1) 11 (2) 7 (3) 5 (4) 4

2A-2 Diatomic Molecules (Gases) (1 Questions)
- Some pure elements only exist in nature as a pair of atoms bonded together including O₂, H₂, N₂, F₂, Cl₂, Br₂, and I₂

3. Which two particle diagrams represent mixtures of diatomic elements?

<table>
<thead>
<tr>
<th>Key</th>
</tr>
</thead>
<tbody>
<tr>
<td>☐ = atom of one element</td>
</tr>
<tr>
<td>● = atom of another element</td>
</tr>
</tbody>
</table>

   (A) A and B  (B) A and C  (C) B and C  (D) B and D

2A-3 Molecular Formulas (See Questions 2A-4)
- Represents the number and kinds of atoms in a molecule of a covalent compound

2A-4 Empirical Formulas (24 Questions)
- The molecular formula reduced by its greatest common denominator.

4. Which pair consists of a molecular formula and its corresponding empirical formula?
   (1) C₂H₂ and CH₃CH₃  (2) C₆H₆ and C₂H₂  (3) P₄O₁₀ and P₂O₅  (4) SO₂ and SO₃

5. Which pair of formulas correctly represents a molecular formula and its corresponding empirical formula?
   (1) C₂H₂ and CH  (2) C₃H₄ and CH₂  (3) C₄H₈ and CH  (4) C₃H₆ and C₂H₂

6. Write the empirical formula for the compound C₆H₁₂O₆.
   [1]

7. The compounds C₂H₄ and C₄H₈ have the same
   (1) freezing point at standard pressure  (2) boiling point at standard pressure
   (3) molecular formula  (4) empirical formula

8. Given the structural formula:
   HO–C–C–C–C–OH
   What is the empirical formula of this compound?
   (1) CH₃O  (2) C₂H₅O  (3) C₄H₁₀O₂  (4) C₆H₂₀O₄
2B-1 Identifying Compounds (4 Questions)
- Compounds contain 2 or more elements chemically combined.
- Compounds are always electrically neutral (no charge)

9. Which species represents a chemical compound?
(1) N₂   (3) Na   (2) NH₄⁺   (4) NaHCO₃

10. Which list of formulas represents compounds, only?
(1) CO₂, H₂O, NH₃   (3) H₂, Ne, NaCl   (2) H₂, N₂, O₂   (4) MgO, NaCl, O₂

2B-2 Ionic Charge (3 Questions)
- Protons have a +1 charge, electrons a -1 charge while neutrons are neutral (0).
- (# of protons) - (# of electrons) = Ionic Charge

11. What is the net charge on an ion that has 9 protons, 11 neutrons, and 10 electrons?
(1) 1+   (3) 1-   (2) 2+   (4) 2-

12. What is the net charge on an ion that has 8 protons, 9 neutrons, and 10 electrons?
(1) 1+   (3) 1-   (2) 2+   (4) 2-

2B-3 Counting Electrons (7 Questions)
- The atomic number is equal to the number of protons in an atom. Atoms are neutral having the same number of electrons. The following formula can be used to determine the number of electrons in an ion.
- (Atomic Number)- (Charge) = Number of Electrons (Remember minus(-) a minus(-) is a positive(+))

13. Which symbol represents a particle that has the same total number of electrons as S²⁻?
(1) O²⁻   (3) Se²⁻   (2) Si   (4) Ar

14. Which symbol represents a particle with a total of 10 electrons?
(1) O²⁻   (3) Al   (2) N³⁺   (4) Al³⁺

2B-4 Polyatomic Ions (3 Questions)
- A polyatomic ion is a group of atoms that are covalently bonded together and have a charge
- The polyatomic ions are listed in Table E

15. What is the name of the polyatomic ion in the compound Na₂O₂?
(1) hydroxide   (3) oxide   (2) oxalate   (4) peroxyde

16. Which polyatomic ion contains the greatest number of oxygen atoms?
(1) hydroxide   (3) oxide   (2) carbonate   (4) peroxyde

2B-5 Ion Electron Configuration (4 Questions)
- Ion electron configuration of an atom is similar to the electron configuration of its closest Nobel gas.
- This provides stability to the ion.

17. Which particle has the same electron configuration as a potassium ion?
(1) fluoride ion   (3) neon atom   (2) sodium ion   (4) argon atom

18. Which electron configuration is correct for a sodium ion?
(1) 2-7   (3) 2-8-1   (2) 2-8   (4) 2-8-2

2B-6 Ion Formation (4 Questions)
- Atoms become negative ions when they gain electrons. Atoms become positive ions when they lose electrons.

19. When a fluorine atom forms an F⁻ ion, the fluorine atom
(1) gains a proton   (3) loses a proton
(2) gains an electron   (4) loses an electron

20. Compared to a calcium atom, the calcium ion Ca²⁺ has
(1) more protons   (3) more electrons
(2) fewer protons   (4) fewer electrons
2C-1 Naming Binary Covalent Compounds  (1 Questions)
- The steps
  - The element with the lower electronegativity value is named first followed by element with the higher electronegativity value and ending in -ide.
  - If the first element named only has one atom, no prefix is used.
  - If the first element named has more than one atom, use the appropriate prefix in the below table.
  - If the element begins with a vowel, drop the the a or o at the end of the prefix.
  - The second element named should use the appropriate prefix in the below table followed by -ide.
  - If the element begins with a vowel, drop the the a or o at the end of the prefix.

<table>
<thead>
<tr>
<th>Number of Atoms</th>
<th>1</th>
<th>2</th>
<th>3</th>
<th>4</th>
<th>5</th>
</tr>
</thead>
<tbody>
<tr>
<td>Prefix</td>
<td>mono-</td>
<td>di-</td>
<td>tri-</td>
<td>tetra-</td>
<td>penta-</td>
</tr>
</tbody>
</table>

Base your answer to question 21 on the information below.

A scientist in a chemistry laboratory determined the molecular formulas for two compounds containing nitrogen and oxygen to be NO₂ and N₂O₅.

21. Write an IUPAC name for the compound N₂O₅. [1]

2C-2 Writing Binary Covalent Formulas  (2 Questions)

Base your answer to question 22 on the information below.

The particle diagrams below represent the reaction between two nonmetals, A₂ and Q₂.

Answer: ____________________________

22. Using the symbols A and Q, write the chemical formula of the product. [1]

2C-3 Naming Binary Ionic Compounds  (2 Questions)
- The steps
  - Positive charged ion (Usually a metal) is named first.
  - Negatively charged ion (Usually a non-metal) is named last with ending of -ide.

23. What is the IUPAC name for the compound ZnO?
   (1) zinc oxide   (2) zinc oxalate   (3) zinc peroxide   (4) zinc hydroxide

24. Write the chemical name of the compound Y₂O₃. [1]
2C-4 Writing Binary Ionic Formulas  (3 Questions)
- Use the oxidation numbers to determine the subscripts.
- The oxidation state of the two elements can be looked up in the periodic table.
- The sum of the oxidation numbers in a compound must be zero.
- The crisscross method may be used

25. Which substance has a chemical formula with the same ratio of metal ions to nonmetal ions as in potassium sulfide?
   (1) sodium oxide  (2) sodium chloride  (3) magnesium oxide  (4) magnesium chloride

26. Write the formula for the salt barium chloride. [1]

2C-5 Naming Stock Compounds  (13 Questions)
- Some metals have multiple common oxidation numbers. Compounds formed by these elements are distinguished by using Roman numerals after the metal that indicate the oxidation number.
- The proper oxidation number of the metal that is written in Roman Numerals can be determined by using the ratios of the elements within the compound and their subscripts.
- The sum of the oxidation numbers in a compound must be zero.
- The crisscross method may be used.

27. Which formula correctly represents the composition of iron (III) oxide?
   (1) FeO₃  (2) Fe₂O₃  (3) Fe₃O  (4) Fe₃O₂

28. What is the correct formula for iron (III) phosphate?
   (1) FeP  (2) Fe₃P₂  (3) FePO₄  (4) Fe₃(PO₄)₂

29. What is the formula of titanium(II) oxide?
   (1) TiO  (2) TiO₂  (3) Tl₂O  (4) Tl₂O₃

2C-6 Writing Stock Formulas  (4 Questions)
- Use the oxidation numbers to determine the subscripts.
- The oxidation state of the metal is given while the nonmetal can be looked up in the periodic table.
- The sum of the oxidation numbers in a compound must be zero.
- The crisscross method may be used

32. What is the chemical formula for tin(II) fluoride? [1]

33. Write the formula for arsenic (III) oxide. [1]

2C-7 Naming Polyatomic Compounds  (4 Questions)
- The steps:
  - If the polyatomic ion is negative, just name the positive ion first followed by the name of the polyatomic ion.
  - If the polyatomic ion is positive, just name the polyatomic ion first followed by the negative ion with ending of -ide.
  - If both the positive and negative ions are polyatomic, just name the positive polyatomic ion first followed by the negative polyatomic ion.

34. What is the correct IUPAC name for the compound NH₄Cl?
   (1) nitrogen chloride  (2) nitrogen chlorate  (3) ammonium chloride  (4) ammonium chlorate

35. What is the chemical formula for sodium sulfate?
   (1) Na₂SO₃  (2) Na₂SO₄  (3) NaSO₃  (4) NaSO₄
2C-8 Writing Polyatomic Formulas  (4 Questions)
- Electrons are shared.

36. Brand B antacid contains the acid-neutralizing agent sodium hydrogen carbonate. Write the chemical formula for sodium hydrogen carbonate. [1]

37. Write the formula for yttrium hydroxide. [1]

38. Write the chemical formula for sodium nitrate. [1]

2C-9 Electronegativity   (2 Questions)
- Electronegativity a measure of attraction for electrons bonded to another atom.

39. The strength of an atom’s attraction for the electrons in a chemical bond is the atom (1) electronegativity (3) heat of reaction (2) ionization energy (4) heat of formation

40. Which term indicates how strongly an atom attracts the electrons in a chemical bond?

2D-1 Chemical Change  (23 Questions)
- A chemical reaction occurred forming new substance(s) with new properties is(are) formed.
- Decomposes, combines, burns, reacts
- Elements cannot be broken down by chemical change (See C1F-1)
- Use the reference table if you are unsure if a substance is an element or not.

41. Which set of procedures and observations indicates a chemical change?

(1) Ethanol is added to an empty beaker and the ethanol eventually disappears.
(2) A solid is gently heated in a crucible and the solid slowly turns to liquid.
(3) Large crystals are crushed with a mortar and pestle and become powder.
(4) A cool, shiny metal is added to water in a beaker and rapid bubbling occurs.

42. Which process represents a chemical change?

(1) melting of ice (2) corrosion of copper (3) evaporation of water (4) crystallization of sugar

43. Which substance can be decomposed by chemical means?

(1) aluminum (2) octane (3) silicon (4) xenon

44. Which substance can not be broken down by a chemical reaction?

(1) ammonia (2) argon (3) methane (4) water

45. Which substance can not be decomposed by a chemical change?

(1) Ne (2) N₂O (3) HF (4) H₂O

2D-2 Physical Change  (3 Questions)
- Different form of the same substance.
- Melts, vaporizes, sublimes, freezes, condenses, filters, grinds, mixes, tears, rips

46. Which diagram represents a physical change, only?

47. Which equation represents a physical change?

(1) H₂O(s) + 6.01 kJ ➜ H₂O(l)
(2) 2H₂(g) + O₂(g) ➜ 2H₂O(g) + 483.6 kJ
(3) H₂(g) + I₂(g) + 53.0 kJ ➜ 2HI(g)
(4) N₂(g) + 2O₂(g) + 66.4 kJ ➜ 2NO₂(g)
2D-3 Chemical Property  (6 Questions)
- The ability of the substance to chemically react with another substance.
- New substances with new properties are formed.

48. Given the balanced particle-diagram equation:

\[ \text{\textcircled{O}} + \text{\textcircled{O}} \rightarrow \text{\textcircled{O}} \]

Which statement describes the type of change and the chemical properties of the product and reactants?
(1) The equation represents a physical change, with the product and reactants having different chemical properties.
(2) The equation represents a physical change, with the product and reactants having identical chemical properties.
(3) The equation represents a chemical change, with the product and reactants having different chemical properties.
(4) The equation represents a chemical change, with the product and reactants having identical chemical properties.

49. Which statement describes a chemical property of the element magnesium?
(1) Magnesium is malleable.
(2) Magnesium conducts electricity.
(3) Magnesium reacts with an acid.
(4) Magnesium has a high boiling point.

50. Which statement describes a chemical property of iron?
(1) Iron can be flattened into sheets.
(2) Iron conducts electricity and heat.
(3) Iron combines with oxygen to form rust.
(4) Iron can be drawn into a wire.

51. Which statement describes a chemical property of hydrogen gas?
(1) Hydrogen gas burns in air.
(2) Hydrogen gas is colorless.
(3) Hydrogen gas has a density of 0.000 09 g/cm\(^3\) at STP.
(4) Hydrogen gas has a boiling point of 20. K at standard pressure.

2D-4 Physical Property  (2 Questions)
- Properties associated with the physical aspects of a substance.

Base your answers to questions 52 and 53 on the information below.

Two sources of copper are cuprite, which has the IUPAC name copper(I) oxide, and malachite, which has the formula \( \text{Cu}_2\text{CO}_3(\text{OH})_2 \). Copper is used in home wiring and electric motors because it has good electrical conductivity. Other uses of copper not related to its electrical conductivity include coins, plumbing, roofing, and cooking pans. Aluminum is also used for cooking pans.

At room temperature, the electrical conductivity of a copper wire is 1.6 times greater than an aluminum wire with the same length and cross-sectional area. At room temperature, the heat conductivity of copper is 1.8 times greater than the heat conductivity of aluminum. At STP, the density of copper is 3.3 times greater than the density of aluminum.

52. Identify one physical property of copper that makes it a good choice for uses that are not related to electrical conductivity. [1]

53. Identify one physical property of aluminum that could make it a better choice than copper for a cooking pan. [1]

2E-1 Balancing Equations  (26 Questions)
- Use coefficients to equalize the reactants and products.

54. Balance the equation below, using the smallest whole-number coefficients. [1]

\[ \text{C}_6\text{H}_{12}(g) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(g) \]

55. Balance the equation for the reaction of \( \text{Fe}_2\text{O}_3 \) and \( \text{CO} \), using the smallest whole-number coefficients. [1]

\[ \text{Fe}_2\text{O}_3 + \text{CO} \rightarrow \text{Fe} + \text{CO}_2 \]
56. Which chemical equation is correctly balanced?
   (1) H₂(g) + O₂(g) → H₂O(g)
   (2) N₂(g) + H₂(g) → NH₃(g)
   (3) 2NaCl(s) → Na(s) + Cl₂(g)
   (4) 2KCl(s) → 2K(s) + Cl₂(g)

57. A balanced equation representing a chemical reaction can be written using
   (1) chemical formulas and mass numbers
   (2) chemical formulas and coefficients
   (3) first ionization energies and mass numbers

2E-2 Conservation (11 Questions)
- Mass, energy and charge are conserved in every chemical reaction.
- What is to the left (type, number and total charge) of the arrow must be to the right of the arrow

58. Which equation shows a conservation of mass?
   (1) Na + Cl₂ → NaCl
   (2) Al + Br₂ → AlBr₃
   (3) Mg + 2AgNO₃(aq) → Mg(NO₃)₂(aq) + 2Ag(s)
   (4) PCl₅ → PCl₃ + Cl₂

59. Which equation shows conservation of both mass and charge?
   (1) Cl₂ + Br⁻ → Cl⁻ + Br₂
   (2) Cu + 2Ag⁺ → Cu²⁺ + Ag
   (3) Zn + Cr³⁺ → Zn²⁺ + Cr²⁺
   (4) Ni + Pb²⁺ → Ni²⁺ + Pb

60. Which equation shows conservation of atoms?
   (1) H₂ + O₂ → H₂O
   (2) H₂ + O₂ → 2H₂O

61. Which quantities must be conserved in all chemical reactions?
   (1) mass, charge, density
   (2) mass, charge, energy
   (3) charge, volume, density
   (4) charge, volume, energy

2F-1 Single Replacement (8 Questions)
- Form A + BX ↔ AX + B or Form A + XB ↔ B + XA
- All are redox reactions.
- Single replacement reactions spontaneously occur when a more reactive substance replaces a less reactive substance in the compound.
- In the above example, element A is must be more reactive than element B.
- See table J (Activity Series)

62. Given the reaction:
   Mg(s) + 2AgNO₃(aq) → Mg(NO₃)₂(aq) + 2Ag(s)
   Which type of reaction is represented?
   (1) single replacement
   (2) double replacement
   (3) synthesis
   (4) decomposition

63. Which balanced equation represents a single-replacement reaction?
   (1) Mg + 2AgNO₃ → Mg(NO₃)₂ + 2Ag
   (2) 2Mg + O₂ → 2MgO
   (3) MgCO₃ → MgO + CO₂
   (4) MgCl₂ + 2AgNO₃ → 2AgCl + Mg(NO₃)₂

64. Based on Reference Table J, which two reactants react spontaneously?
   (1) Mg(s) + ZnCl₂(aq)
   (2) Cu(s) + FeSO₄(aq)
   (3) Pb(s) + ZnCl₂(aq)
   (4) Co(s) + NaCl(aq)

65. Which reaction occurs spontaneously?
   (1) Cl₂(g) + 2NaBr(aq) → Br₂(ℓ) + 2NaCl(aq)
   (2) Cl₂(g) + 2NaF(aq) → F₂(g) + 2NaCl(aq)
   (3) I₂(s) + 2NaBr(aq) → Br₂(ℓ) + 2NaI(aq)
   (4) I₂(s) + 2NaF(aq) → F₂(g) + 2NaI(aq)

2F-2 Double Replacement (2 Questions)
- Form AB + CD → AD + CB
- This will only occur if one of the substances precipitates out of solution, a gas is given off, or a molecular compound such as water is formed.

66. Which equation represents a double replacement reaction?
   (1) 2Na + 2H₂O → 2NaOH + H₂
   (2) CaCO₃ → CaO + CO₂
   (3) LiOH + HCl → LiCl + H₂O
   (4) CH₄ + 2O₂ → CO₂ + 2H₂O

67. Given the balanced equation:
   AgNO₃(aq) + NaCl(aq) → NaNO₃(aq) + AgCl(s)
   This reaction is classified as
   (1) synthesis
   (2) decomposition
   (3) single replacement
   (4) double replacement
2F-3 Synthesis (3 Questions)
- Form A + B → AB
- Two reactants and one product.

68. Given the balanced equation representing a reaction:

4Al(s) + 3O_2(g) → 2Al_2O_3(s)

Which type of chemical reaction is represented by this equation?
(1) double replacement
(2) single replacement
(3) substitution
(4) synthesis

69. In which type of reaction do two or more substances combine to produce a single substance?
(1) synthesis
(2) decomposition
(3) single replacement
(4) double replacement

2F-4 Decomposition (10 Questions)
- Form AB → A+B
- One reactant and two products.

70. Two different samples decompose when heated. Only one of the samples is soluble in water. Based on this information, these two samples are
(1) both the same element
(2) two different elements
(3) both the same compound
(4) two different compounds

71. Given the balanced equation: 2KClO_3 → 2KCl + 3O_2
Which type of reaction is represented by this equation?
(1) synthesis
(2) decomposition
(3) single replacement
(4) double replacement

72. Which equation represents a decomposition reaction?
(1) CaCO_3(s) → CaO(s) + CO_2(g)
(2) Cu(s) + 2AgNO_3(aq) → 2Ag(s) + Cu(NO_3)_2(aq)
(3) 2H_2(g) + O_2(g) → 2H_2O(ℓ)
(4) KOH(aq) + HCl(aq) → KCl(aq) + H_2O(ℓ)

73. Given the equation:
CaCO_3(s) → CaO(s) + CO_2(g)
Name the type of reaction this equation represents. [1]
Ans.____________________

2F-5 General Types of Reactions (2 Questions)

74. Which list consists of types of chemical formulas?
(1) atoms, ions, molecules
(2) metals, nonmetals, metalloids
(3) empirical, molecular, structural
(4) synthesis, decomposition, neutralization

75. Which list includes three types of chemical reactions?
(1) condensation, double replacement, and sublimation
(2) condensation, solidification, and synthesis
(3) decomposition, double replacement, and synthesis
(4) decomposition, solidification, and sublimation

2G-1 Find the missing product (or Reactant) (3 Questions)
- What is to the left (type, number and total charge) of the arrow must be to the right of the arrow
- Determine what is missing and write the chemical formula that equalizes the equation

76. Given the equation:

x + Cl_2 → C_2H_5Cl + HCl

Which molecule is represented by X?
(1) C_2H_4
(2) C_2H_6
(3) C_3H_6
(4) C_3H_8

77. Given the incomplete equation:

4Fe + 3O_2 → 2X

Which compound is represented by X?
(1) FeO
(2) Fe_2O_3
(3) Fe_3O_2
(4) Fe_3O_4

78. Given the incomplete equation for the combustion of ethane:

2C_2H_6 + 7O_2 → 4CO_2 + 6 ____

What is the formula of the missing product?
(1) CH_3OH
(2) HCOOH
(3) H_2O
(4) H_2O_2
Base your answer to question 79 on the information below.

In an experiment, 2.54 grams of copper completely reacts with sulfur, producing 3.18 grams of copper(I) sulfide.

79. Determine the total mass of sulfur consumed. [1]

______________ g

80. Given the balanced equation representing a reaction:

\[ \text{CaO(s)} + \text{CO}_2(g) \rightarrow \text{CaCO}_3(s) + \text{heat} \]

What is the total mass of \text{CaO(s)} that reacts completely with 88 grams of \text{CO}_2(g) to produce 200 grams of \text{CaCO}_3(s)?

(1) 56 g  (3) 112 g
(2) 88 g  (4) 288 g

81. Given the balanced equation representing a reaction:

\[ 2\text{NaCl(ℓ)} \rightarrow 2\text{Na(ℓ)} + \text{Cl}_2(g) \]

A 1170.-gram sample of \text{NaCl(ℓ)} completely reacts, producing 460 grams of \text{Na(ℓ)}. What is the total mass of \text{Cl}_2(g) produced?

(1) 355 g  (3) 1420. g
(2) 710. g  (4) 1630. g

82. Given the balanced equation representing a reaction:

\[ 2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O} \]

What is the mass of \text{H}_2\text{O} produced when 10.0 grams of \text{H}_2 reacts completely with 80.0 grams of \text{O}_2?

(1) 70.0 g  (3) 180. g
(2) 90.0 g  (4) 800. g
3A-1 Formula Mass (2 Questions)

• Formula mass is the sum of all of the atomic masses of the atoms present in the formula in amu.
  - Example: What is the formula mass of (NH₄)₃PO₄?
    N  3 x 14.0 = 42
    H 12 x  1.0 = 12
    P  1 x 30.0 = 31
    O  4 x 16.0 = 64
    149 amu
    or
    (3x14.0)+(12x1.0)+(1x30.0)+(4x16) = 149 amu

1. The sum of the atomic masses of the atoms in one molecule of C₃H₆Br₂ is called the
   (1) formula mass  (3) percent abundance
   (2) isotopic mass  (4) percent composition

2. In the space below, show a correct numerical setup for calculating the formula mass of glucose, C₆H₁₂O₆. [1]

3A-2 Gram Formula Mass (13 Questions)

• The gram-formula mass is the formula mass expressed in grams. This is equivalent to the mass of one mole of the compound. The gram formula mass is often expressed a g/mol. In the problem above, the gram-formula mass of (NH₄)₃PO₄ = 149 g/mol

Base your answer to question 3 on the information below.

Glycine, NH₂CH₂COOH, is an organic compound found in proteins. Acetamide, CH₃CONH₂, is an organic compound that is an excellent solvent. Both glycine and acetamide consist of the same four elements, but the compounds have different functional groups.

3. In the space below, calculate the gram-formula mass of glycine. Your response must include both a numerical setup and the calculated result. [2]

   __________________________ g/mol

Base your answer to question 4 on the information below.

₂H₂O₂ → ₂H₂O + O₂

4. Determine the gram-formula mass of H₂O₂. [1]
   __________________________ g/mol

5. What is the gram-formula mass of Ca₃(PO₄)₂?
   (1) 248 g/mol  (3) 279 g/mol
   (2) 263 g/mol  (4) 310. g/mol

6. What is the gram-formula mass of (NH₄)₃PO₄?
   (1) 112 g/mol  (3) 149 g/mol
   (2) 121 g/mol  (4) 242 g/mol

7. The gram-formula mass of NO₂ is defined as the mass of
   (1) one mole of NO₂  (3) two moles of NO
   (2) one molecule of NO₂  (4) two molecules of NO

   __________________________ g/mol
3B-1 Percent by Mass  (23 Questions)

- %composition by mass = \( \frac{\text{mass of part}}{\text{mass of whole}} \times 100 \)

Base your answer to question 8 on the information below.

A scientist in a chemistry laboratory determined the molecular formulas for two compounds containing nitrogen and oxygen to be \( \text{NO}_2 \) and \( \text{N}_2\text{O}_5 \).

8. In the space provided below, show a correct numerical setup for calculating the percent composition by mass of oxygen in \( \text{NO}_2 \). [1]

9. The percent by mass of hydrogen in \( \text{NH}_3 \) is equal to
   - (1) \( \frac{17}{1} \times 100 \)
   - (2) \( \frac{17}{3} \times 100 \)
   - (3) \( \frac{1}{17} \times 100 \)
   - (4) \( \frac{3}{17} \times 100 \)

10. What is the percent by mass of oxygen in \( \text{H}_2\text{SO}_4 \)?
    [formula mass = 98]
    - (1) 16%
    - (2) 33%
    - (3) 65%
    - (4) 98%

11. What is the percent by mass of oxygen in propanal, \( \text{CH}_3\text{CH}_2\text{CHO} \)?
    - (1) 10.0%
    - (2) 27.6%
    - (3) 38.1%
    - (4) 62.1%

12. The percentage by mass of Br in the compound \( \text{AlBr}_3 \) is closest to
    - (1) 10.0%
    - (2) 25%
    - (3) 75%
    - (4) 90.0%

3B-2 Percent by Mass Hydrate  (10 Questions)

- Heating the hydrate will drive off the mass of water. This mass can be used to find the % of water in the hydrate.
- %composition of water = \( \frac{\text{mass of water}}{\text{mass of whole}} \times 100 \)

Base your answers to questions 13 and 14 on the information below.

Gypsum is a mineral that is used in the construction industry to make drywall (sheetrock). The chemical formula for this hydrated compound is \( \text{CaSO}_4 \cdot 2 \text{H}_2\text{O} \). A hydrated compound contains water molecules within its crystalline structure. Gypsum contains 2 moles of water for each 1 mole of calcium sulfate.

13. What is the gram formula mass of \( \text{CaSO}_4 \cdot 2 \text{H}_2\text{O} \)? [1]
   Answer: ___________________________ g/mol

14. a In the space provided below show a correct numerical setup for calculating the percent composition by mass of water in this compound. [1]
   b Record your answer. [1]
   Answer: ___________________________
3B-3 Percent Rearranged (5 Questions)

- mass of part = \( \frac{\% \text{composition by mass} \times \text{mass of whole}}{100} \)
- mass of whole = \( \frac{\text{mass of part}}{\% \text{composition by mass}} \times 100 \)

17. A sample of a substance containing only magnesium and chlorine was tested in the laboratory and was found to be composed of 74.5% chlorine by mass. If the total mass of the sample was 190.2 grams, what was the mass of the magnesium?

(1) 24.3 g (3) 70.9 g
(2) 48.5 g (4) 142 g

Base your answer to question 18 on the information below.

At STP, iodine, I\(_2\), is a crystal, and fluorine, F\(_2\), is a gas. Iodine is soluble in ethanol, forming a tincture of iodine. A typical tincture of iodine is 2% iodine by mass.

18. Determine the total mass of I\(_2\) in 25 grams of this typical tincture of iodine. [1]

3C-1 Moles from gram-formula (14 Questions)

- 1 mole of any substance = the gram-formula mass of that substance
- number of moles = \( \frac{\text{given mass (g)}}{\text{gram formula mass}} \)

Base your answer to question 19 on the information below.

The compound 1,2-ethanediol can be mixed with water. This mixture is added to automobile radiators as an engine coolant. The cooling system of a small van contains 6690 grams of 1,2-ethanediol. Some properties of water and 1,2-ethanediol are given in the table below.

<table>
<thead>
<tr>
<th>Property</th>
<th>Water ((H_2O))</th>
<th>1,2-ethanediol ((CH_2OHCH_2OH))</th>
</tr>
</thead>
<tbody>
<tr>
<td>gram-formula mass (g/mol)</td>
<td>18.0</td>
<td>62.0</td>
</tr>
<tr>
<td>boiling point at standard pressure (°C)</td>
<td>100.0</td>
<td>197.2</td>
</tr>
</tbody>
</table>

19. In the space below calculate the total number of moles of 1,2-ethanediol in the small van’s cooling system. Your response must include both a correct numerical setup and the calculated result. [2]

Base your answer to question 20 on the information below.

Vitamin C, also known as ascorbic acid, is water soluble and cannot be produced by the human body. Each day, a person’s diet should include a source of vitamin C, such as orange juice. Ascorbic acid has a molecular formula of \(C_6H_8O_6\) and a gram-formula mass of 176 grams per mole.

20. The gram-formula mass of NO\(_2\) is defined as the mass of

(1) one mole of NO\(_2\) (3) two moles of NO
(2) one molecule of NO\(_2\) (4) two molecules of NO

Base your answer to question 21 on the information below.

21. Determine the number of moles of vitamin C in an orange that contains 0.071 gram of vitamin C. [1]

______________________ mol

Base your answer to question 22 on the information below.

The decomposition of sodium azide, NaN\(_3\)(s), is used to inflate airbags. On impact, the NaN\(_3\)(s) is ignited by an electrical spark, producing N\(_2\)(g) and Na(s). The N\(_2\)(g) inflates the airbag.

22. What is the total number of moles present in a 52.0-gram sample of NaN\(_3\)(s) (gram-formula mass = 65.0 gram/mole)? [1]

______________________ mol
3C-2 Mole Rearranged (7 Questions)

- gram formula mass = \( \frac{\text{given mass (g)}}{\text{number of moles}} \)
- mass (g) = number of moles \( \times \) gram formula mass

23. The gram-formula mass of a compound is 48 grams. The mass of 1.0 mole of this compound is
   (1) 1.0 g  (3) 48 g
   (2) 4.8 g  (4) 480 g

24. Calculate the total mass of propane consumed during the lantern test. Your response must include both a correct numerical setup and the calculated result. [2]

25. A 1.0-mole sample of krypton gas has a mass of
   (1) 19 g  (3) 39 g
   (2) 36 g  (4) 84 g

26. Determine the mass of 5.20 moles of \( \text{C}_6\text{H}_{12} \) (gram-formula mass = 84.2 grams/mole). [1]

3D-1 Finding Molecular Formulas (6 Questions)

- If you are given the empirical formula and the gram-formula mass, the molecular formula can be determined by:
  - Find the gram-formula mass for the empirical formula
  - Divide the gram-formula mass of the compound by the gram-formula empirical formula determined in the previous step. The result must be a whole number (Integer)!
  - Multiply each subscript to determine the molecular formula.

27. What is the molecular formula of a compound that has a molecular mass of 54 and the empirical formula \( \text{C}_2\text{H}_3 \).
   (1) \( \text{C}_2\text{H}_3 \)  (3) \( \text{C}_6\text{H}_9 \)
   (2) \( \text{C}_4\text{H}_6 \)  (4) \( \text{C}_8\text{H}_{12} \)

28. A substance has an empirical formula of \( \text{CH}_2 \) and a molar mass of 56 grams per mole. The molecular formula for this compound is
   (1) \( \text{CH}_2 \)  (3) \( \text{C}_4\text{H}_8 \)
   (2) \( \text{C}_4\text{H}_6 \)  (4) \( \text{C}_8\text{H}_{12} \)

29. A compound has a molar mass of 90 grams per mole and the empirical formula \( \text{CH}_2\text{O} \). What is the molecular formula of this compound?
   (1) \( \text{CH}_2\text{O} \)  (3) \( \text{C}_2\text{H}_6\text{O}_3 \)
   (2) \( \text{C}_2\text{H}_4\text{O}_2 \)  (4) \( \text{C}_4\text{H}_8\text{O}_4 \)

30. A compound has a gram-formula mass of 56 grams per mole. What is the molecular formula for this compound?
   (1) \( \text{CH}_2 \)  (3) \( \text{C}_3\text{H}_6 \)
   (2) \( \text{C}_2\text{H}_4 \)  (4) \( \text{C}_4\text{H}_8 \)

31. A compound has the empirical formula \( \text{CH}_2\text{O} \) and a gram-formula mass of 60 grams per mole. What is the molecular formula of this compound?
   (1) \( \text{CH}_2\text{O} \)  (3) \( \text{C}_3\text{H}_6\text{O} \)
   (2) \( \text{C}_2\text{H}_4\text{O}_2 \)  (4) \( \text{C}_4\text{H}_8\text{O}_4 \)
3E-1 Ratios by Equations  (5 Questions)

- Coefficients can represent the mole ratios of reactants and products of a balance equation. To determine the ratio between two coefficients, list the two coefficients in the order written in the question and separate them by colons. Reduce to lowest terms if appropriate.

32. Given the reaction:

\[ 2\text{C}_2\text{H}_6 + 7\text{O}_2 \rightarrow 4\text{CO}_2 + 6\text{H}_2\text{O} \]

What is the ratio of moles of CO\(_2\) produced to moles of C\(_2\)H\(_6\) consumed?

- (1) 1 to 1
- (2) 2 to 1
- (3) 3 to 2
- (4) 7 to 2

33. Given the reaction:

\[ \text{N}_2(g) + 3\text{H}_2(g) \rightleftharpoons 2\text{NH}_3(g) \]

What is the mole-to-mole ratio between nitrogen gas and hydrogen gas?

- (1) 1:2
- (2) 1:3
- (3) 2:2
- (4) 2:3

34. Given the balanced equation representing a reaction:

\[ \text{F}_2(g) + \text{H}_2(g) \rightarrow 2\text{HF}(g) \]

What is the mole ratio of H\(_2\)(g) to HF(g) in this reaction?

- (1) 1:1
- (2) 1:2
- (3) 2:1
- (4) 2:3

35. Given the balanced equation representing the reaction between propane and oxygen:

\[ \text{C}_3\text{H}_8 + 5\text{O}_2 \rightarrow 3\text{CO}_2 + 4\text{H}_2\text{O} \]

According to this equation, which ratio of oxygen to propane is correct?

- (1) 5 grams O\(_2\) to 1 gram C\(_3\)H\(_8\)
- (2) 5 moles O\(_2\) to 1 mole C\(_3\)H\(_8\)
- (3) 10 grams O\(_2\) to 11 gram C\(_3\)H\(_8\)
- (4) 10 moles O\(_2\) to 11 mole C\(_3\)H\(_8\)

3E-2 Moles by Equations  (19 Questions)

- Coefficients can represent the mole ratios of reactants and products of a balance equation. Therefore, the following equation can be used to determine the # of moles when a quantity other than the coefficient is used or produced.

\[
\frac{\text{Moles 1}}{\text{Coefficient 1}} = \frac{\text{Moles 2}}{\text{Coefficient 2}}
\]

Base your answer to question 36 on the balanced equation below.

\[ \text{C}_5\text{H}_{12}(g) + 8\text{O}_2(g) \rightarrow 5\text{CO}_2(g) + 6\text{H}_2\text{O}(g) \]

36. a Using your balanced equation, show a correct numerical setup for calculating the total number of moles of H\(_2\)O(g) produced when 5.0 moles of O\(_2\)(g) are completely consumed. [1]

36. b Record your answer. [1]

Answer: ________________ mol H\(_2\)O

Base your answer to question 37 on the information below.

Some dry chemicals can be used to put out forest fires. One of these chemicals is NaHCO\(_3\). When NaHCO\(_3\)(s) is heated, one of the products is CO\(_2\)(g), as shown in the balanced equation below.

\[ 2\text{NaHCO}_3(s) + \text{heat} \rightarrow \text{Na}_2\text{CO}_3(s) + \text{H}_2\text{O}(g) + \text{CO}_2(g) \]

37. Determine the total number of moles of CO\(_2\)(g) produced when 7.0 moles of NaHCO\(_3\)(s) is completely reacted. [1]

_______________ mol

38. Given the balanced equation representing a reaction:

\[ 2\text{C}_2\text{H}_6 + 7\text{O}_2 \rightarrow 4\text{CO}_2 + 6\text{H}_2\text{O} \]

Determine the total number of moles of oxygen that react completely with 8.0 moles of C\(_2\)H\(_6\). [1]

_______________ mol
3F-1 Density (15 Questions)

- density = \( \frac{\text{mass}}{\text{volume}} \)
- The density of elements can be found in Table S.

Base your answer to question 39 on the information below.

A student prepared two mixtures, each in a labeled beaker. Enough water at 20.°C was used to make 100 milliliters of each mixture. Information about Two Mixtures at 20.°C

<table>
<thead>
<tr>
<th>Composition</th>
<th>Mixtue 1</th>
<th>Mixtue 2</th>
</tr>
</thead>
<tbody>
<tr>
<td>Student Observations</td>
<td>NaCl in H₂O</td>
<td>Fe filings in H₂O</td>
</tr>
<tr>
<td>Other Data</td>
<td>• colorless liquid • no visible solid on bottom of beaker</td>
<td>• colorless liquid • black solid on bottom of beaker</td>
</tr>
<tr>
<td></td>
<td>• mass of NaCl(s) dissolved = 2.3 g</td>
<td>• mass of Fe(s) = 15.9 g • density of Fe(s) = 7.87 g/cm³</td>
</tr>
</tbody>
</table>

39. Determine the volume of the Fe filings used to produce mixture 2. [1]

________________________ cm³

40. At STP, a 7.49-gram sample of an element has a volume of 1.65 cubic centimeters. The sample is most likely

(1) Ta  (2) Tc
(3) Te  (4) Ti

41. An inflated airbag has a volume of 5.00 x 10⁻⁴ cm³ at STP. The density of N₂(g) at STP is 0.00125 g/cm³. What is the total number of grams of N₂(g) in the airbag? [1]

________________________ g

3F-2 Percent Error (10 Questions)

- %error = \( \frac{\text{measured value} - \text{accepted value}}{\text{accepted value}} \) x 100

44. A student measures the mass and volume of a piece of aluminum. The measurements are 25.6 grams and 9.1 cubic centimeters. The student calculates the density of the aluminum. What is the percent error of the student’s calculated density of aluminum?

(1) 1%  (2) 2%
(3) 3%  (4) 4%

45. A student calculated the percent by mass of water in a hydrate as 14.2%. A hydrate is a compound that contains water as part of its crystal structure. If the accepted value is 14.7%, the student’s percent error was

(1) \( \frac{0.5}{14.2} \) x 100  (2) \( \frac{0.5}{14.2} \) x 100
(3) \( \frac{0.5}{14.7} \) x 100  (4) \( \frac{0.5}{14.7} \) x 100
46. Based on data collected during a laboratory investigation, a student determined an experimental value of 322 joules per gram for the heat of fusion of H₂O. Calculate the student's percent error. Your response must include a correct numerical setup and the calculated result. [2]

\[
\text{Percent Error} = \frac{|\text{Experimental Value} - \text{Theoretical Value}|}{\text{Theoretical Value}} \times 100\%
\]

\[
\frac{|322 \text{ J/g} - 333 \text{ J/g}|}{333 \text{ J/g}} \times 100\% = \frac{11 \text{ J/g}}{333 \text{ J/g}} \times 100\% = 3.3\%
\]

3F-3 Miscellaneous Math (8 Questions)

- Metric conversions
- Significant figures

Base your answer to question 47 on the information below.

In a titration experiment, a student uses a 1.4 M HBr(aq) solution and the indicator phenolphthalein to determine the concentration of a KOH(aq) solution. The data for trial 1 is recorded in the table below.

<table>
<thead>
<tr>
<th>Buret Readings</th>
<th>HBr(aq)</th>
<th>KOH(aq)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial volume (mL)</td>
<td>7.50</td>
<td>11.00</td>
</tr>
<tr>
<td>Final volume (mL)</td>
<td>22.90</td>
<td>33.10</td>
</tr>
<tr>
<td>Volume used (mL)</td>
<td>15.40</td>
<td>22.10</td>
</tr>
</tbody>
</table>

47. In a second trial of this experiment, the molarity of KOH(aq) was determined to be 0.95 M. The actual molarity was 0.83 M. What is the percent error in the second trial? [1]

Answer: _______________________

Base your answer to question 51 on the information below.

The balanced equation below represents the reaction between magnesium metal and hydrochloric acid to produce aqueous magnesium chloride and hydrogen gas.

\[\text{Mg(s) + 2HCl(aq) } \rightarrow \text{ MgCl}_2(aq) + \text{ H}_2(g)\]

A piece of Mg(s) has a volume of 0.0640 cubic centimeters. This piece of Mg(s) reacts completely with HCl(aq) to produce H₂(g). The H₂(g) produced has a volume of 112 milliliters and a pressure of 1.00 atmosphere at 298 K.

48. The volume of the piece of Mg(s) is expressed to what number of significant figures? [1]

49. A sample of an element has a mass of 34.261 grams and a volume of 3.8 cubic centimeters. To which number of significant figures should the calculated density of the sample be expressed?

(1) 5  (3) 3
(2) 2  (4) 4

50. Which quantity is equal to 50 kilojoules?

(1) 0.05 J  (3) 5 x 10³ J
(2) 500 J  (4) 5 x 10⁴ J

Base your answer to question 51 on the information below.

The atomic and ionic radii for sodium and chlorine are shown in the table below. Atomic and Ionic Radii

<table>
<thead>
<tr>
<th>Particle</th>
<th>Radius (pm)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Na atom</td>
<td>190.</td>
</tr>
<tr>
<td>Na⁺ ion</td>
<td>102</td>
</tr>
<tr>
<td>Cl atom</td>
<td>97</td>
</tr>
<tr>
<td>Cl⁻ ion</td>
<td>181</td>
</tr>
</tbody>
</table>

51. Convert the radius of an Na⁺ ion to meters. [1]

____________________ m

52. Which quantity of heat is equal to 200. joules?

(1) 20.0 kJ  (3) 0.200 kJ
(2) 2.00 kJ  (4) 0.0200 kJ
4A-1 Phases of Matter (26 Questions)

<table>
<thead>
<tr>
<th>Phase</th>
<th>Molecular Attraction</th>
<th>Symbol</th>
<th>Characteristics</th>
</tr>
</thead>
<tbody>
<tr>
<td>Solid</td>
<td>Strong</td>
<td>(s)</td>
<td>definite volume &amp; shape</td>
</tr>
<tr>
<td>Liquid</td>
<td>Moderate</td>
<td>(l)</td>
<td>definite volume but not a definite shape</td>
</tr>
<tr>
<td>Gas</td>
<td>Little</td>
<td>(g)</td>
<td>neither have a definite volume or shape - they take the shape of their container!</td>
</tr>
</tbody>
</table>

1. Given the particle diagram:

At 101.3 kPa and 298 K, which element could this diagram represent?
(1) Rn  (3) Ag  (2) Xe  (4) Kr

2. Which diagram best represents a gas in a closed container?
(1)  (2)  (3)  (4)

3. In which material are the particles arranged in a regular geometric pattern?
(1) CO₂(g)  (3) H₂O(l)  (2) NaCl(aq)  (4) C₁₂H₂₂O₁₁(s)

4. At STP, fluorine is a gas and bromine is a liquid because, compared to fluorine, bromine has
   (1) stronger covalent bonds
   (2) stronger intermolecular forces
   (3) weaker covalent bonds
   (4) weaker intermolecular forces

5. Given the particle diagram representing four molecules of a substance:

Which particle diagram best represents this same substance after a physical change has taken place?
(1)  (2)  (3)  (4)
6. Given the graph below that represents the uniform cooling of a sample of lauric acid starting as a liquid above freezing point.

![Cooling Curve for a Substance](image)

a Which line segment represents a phase change, only? [1]

Ans. _______________________

b What is the melting point of lauric acid? [1]

Ans. _______________________

c At which point do the particles of lauric acid have the highest average kinetic energy? [1]

Ans. _______________________

d Name the phase change that takes place during this 10-minute cooling time. [1]

Ans. _______________________

Base your answers to questions 7 through 9 on the information below.

Starting as a gas at 206°C, a sample of a substance is allowed to cool for 16 minutes. This process is represented by the cooling curve below.

7. What is the melting point of this substance? [1]

____________________ °C

8. At what time do the particles of this sample have the lowest average kinetic energy? [1]

____________________

9. Using the key below, draw two particle diagrams to represent the two phases of the sample at minute 4. Your response must include at least six particles for each diagram. [1]

<table>
<thead>
<tr>
<th>Key</th>
</tr>
</thead>
<tbody>
<tr>
<td>○ = particle of the substance</td>
</tr>
</tbody>
</table>

One phase of the sample at minute 4

A different phase of the sample at minute 4
4B-2 Phase Change Phenomena (11 Questions)

- **Endothermic** (Absorb energy) - Potential energy increases
  - Solid to Liquid (Heat of Fusion) (Melting)
  - Liquid to Gas (Heat of Vaporization)(Vaporization)
  - Solid to Gas (Sublimation)
- **Exothermic** (Give off energy)
  - Liquid to Solid (Heat of Fusion)(Freezing)
  - Gas to Liquid (Heat of Vaporization)(Condensation)

10. The solid and liquid phases of water can exist in a state of equilibrium at 1 atmosphere of pressure and a temperature of
   (1) 0°C   (2) 100°C   (3) 273°C   (4) 373°C

11. Which phase change is an exothermic process?
   (1) CO$_2$(s) $\rightarrow$ CO$_2$(g)   (2) NH$_3$(g) $\rightarrow$ NH$_3$(l)
   (3) Cu(s) $\rightarrow$ Cu(ℓ)   (4) Hg(ℓ) $\rightarrow$ Hg(g)

4C-1 Temperature Scales (12 Questions)

- $K = ^{\circ}C + 273$

13. Which kelvin temperature is equivalent to –24°C?
   (1) 226 K   (2) 249 K
   (3) 273 K   (4) 297 K

14. At which temperature would atoms of a He(g) sample have the greatest average kinetic energy?
   (1) 25°C   (2) 37°C
   (3) 273 K   (4) 298 K

Base your answer to question 15 on the information below.

A method used by ancient Egyptians to obtain copper metal from copper (I) sulfide ore was heating the ore in the presence of air. Later, copper was mixed with tin to produce a useful alloy called bronze.

15. Convert the melting point of the metal obtained from copper (I) sulfide ore to degrees Celsius. [1]

__________________________ °C

4D-1 Temperature (39 Questions)

- Temperature is a measure of the average kinetic energy
- Heat Energy flows from hot to cold

16. In which sample of water do the molecules have the highest average kinetic energy?
   (1) 20. mL at 100°C   (2) 40. mL at 80°C
   (3) 60. mL at 60°C   (4) 80. mL at 40°C

17. Which sample of ethanol has particles with the highest average kinetic energy?
   (1) 10.0 mL of ethanol at 25°C
   (2) 10.0 mL of ethanol at 55°C
   (3) 100.0 mL of ethanol at 35°C
   (4) 100.0 mL of ethanol at 45°C

18. An iron bar at 325 K is placed in a sample of water. The iron bar gains energy from the water if the temperature of the water is
   (1) 65 K   (2) 45 K
   (3) 65°C   (4) 45°C

19. What occurs when a 35-gram aluminum cube at 100°C is placed in 90. grams of water at 25°C in an insulated cup?
   (1) Heat is transferred from the aluminum to the water, and the temperature of the water increases.
   (2) Heat is transferred from the aluminum to the water, and the temperature of the water decreases.
   (3) Heat is transferred from the water to the aluminum, and the temperature of the water increases.
   (4) Heat is transferred from the water to the aluminum, and the temperature of the water decreases.
4D-2 Heat (22 Questions)

- \( q = mC\Delta T \), \( q = mH_f \) and \( q = mH_v \) where \( q \) heat, \( m \) mass, \( \Delta T \) change in temperature, \( C \) Specific heat capacity, \( H_f \) Heat of fusion, and \( H_v \) Heat of vaporization

Base your answers to questions 20 and 21 on the information below.

A 5.00-gram sample of liquid ammonia is originally at 210. K. The diagram of the partial heating curve below represents the vaporization of the sample of ammonia at standard pressure due to the addition of heat. The heat is not added at a constant rate.

21. Determine the total amount of heat required to vaporize this 5.00-gram sample of ammonia at its boiling point. [1]

22. How much heat energy must be absorbed to completely melt 35.0 grams of \( \text{H}_2\text{O} \) (s) at 0°C?

   (1) 9.54 J  
   (2) 146 J  
   (3) 11 700 J  
   (4) 79 100 J

23. What amount of heat is required to completely melt a 29.95-gram sample of \( \text{H}_2\text{O} \) (s) at 0°C?

   (1) 334 J  
   (2) 2260 J  
   (3) 1.00 x 10³ J  
   (4) 1.00 x 10⁴ J

24. What is the total number of joules released when a 5.00-gram sample of water changes from liquid to solid at 0°C?

   (1) 334 J  
   (2) 1670 J  
   (3) 2260 J  
   (4) 11 300 J

Base your answer to question 25 on the information below.

At a pressure of 101.3 kilopascals and a temperature of 373 K, heat is removed from a sample of water vapor, causing the sample to change from the gaseous phase to the liquid phase. This phase change is represented by the equation below.

\( \text{H}_2\text{O}(g) \rightarrow \text{H}_2\text{O}(l) + \text{heat} \)

25. Determine the total amount of heat released by 5.00 grams of water vapor during this phase change. [1]

26. What is the total amount of heat absorbed by 100.0 grams of water when the temperature of the water is increased from 30.0°C to 45.0°C?

   (1) 418 J  
   (2) 6270 J  
   (3) 12 500 J  
   (4) 18 800 J

---

4E-1 Gas Behavior (29 Questions)

- The Kinetic Molecular Theory (KMT)
- Gas particles exert pressure when they collide with their container's walls.
- All gases have the same number of particles if they occupy the same volume at the same pressure and temperature. The density and mass of the gas may be different, but the number of particles will be the same.
- As he volume decreases, the gas molecules are concentrated and will exert more pressure on the walls of the container.
If the temperature of a gas that is confined in a container increases, the pressure increases. As the temperature increases, the volume increase assuming the container is able to maintain a constant pressure by expanding. The higher the temperature, the greater the velocity of the particles.

Base your answer to question 27 on the information below.

A weather balloon has a volume of 52.5 liters at a temperature of 295 K. The balloon is released and rises to an altitude where the temperature is 252 K.

27. How does this temperature change affect the gas particle motion?

28. A sample of oxygen gas is sealed in container X. A sample of hydrogen gas is sealed in container Z. Both samples have the same volume, temperature, and pressure. Which statement is true?
   1. Container X contains more gas molecules than container Z.
   2. Container X contains fewer gas molecules than container Z.
   3. Containers X and Z both contain the same number of gas molecules.
   4. Containers X and Z both contain the same mass of gas.

29. At 1 atmosphere and 298 K, 1 mole of H₂O(ℓ) molecules and 1 mole of C₂H₅OH(ℓ) molecules both have the same
   1. vapor pressure
   2. average kinetic energy
   3. mass
   4. density

30. At the same temperature and pressure, which sample contains the same number of moles of particles as 1 liter of O₂(g)?
   1. 1 L Ne(g)
   2. 2 L N₂(g)
   3. 0.5 L SO₂(g)
   4. 1 L H₂O (ℓ)

31. A sealed, rigid 1.0-liter cylinder contains He gas at STP. An identical sealed cylinder contains Ne gas at STP. These two cylinders contain the same number of
   1. atoms
   2. electrons
   3. ions
   4. protons

32. Standard pressure is equal to
   1. 1 atm
   2. 1 kPa
   3. 273 atm
   4. 273 kPa

4E-2 Gas problems (21 Problems)

Base your answer to question 33 on the diagram below, which shows a piston confining a gas in a cylinder.

33. The gas volume in the cylinder is 6.2 milliliters and its pressure is 1.4 atmospheres. The piston is then pushed in until the gas volume is 3.1 milliliters while the temperature remains constant. In the space below, calculate the pressure, in atmospheres, after the change in volume. Show all work. [2]

34. The temperature of a 2.0-liter sample of helium gas at STP is increased to 27°C and the pressure is decreased to 80 kPa. What is the new volume of the helium sample?
   1. 1.4 L
   2. 2.0 L
   3. 2.8 L
   4. 4.0 L

35. A gas occupies a volume of 40.0 milliliters at 20°C. If the volume is increased to 80.0 milliliters at constant pressure, the resulting temperature will be equal to
   1. 20°C X 80.0 mL
   2. 20°C X 40.0 mL
   3. 293 K X 80.0 mL
   4. 293 K X 40.0 mL

36. A rigid cylinder with a movable piston contains a 2.0-liter sample of neon gas at STP. What is the volume of this sample when its temperature is increased to 30°C while its pressure is decreased to 900 kilopascals?
   1. 2.5 L
   2. 2.0 L
   3. 1.6 L
   4. 0.22 L
**Ideal Gases (26 Questions)**

- Gas particles do not occupy volume
- There is no attraction between particles
- Hydrogen and helium act most ideal.
  - Mass and polarity can affect the behavior of the gas.
  - Most ideal at low pressure and high temperature.

37. The concept of an ideal gas is used to explain
   (1) the mass of a gas sample
   (2) the behavior of a gas sample
   (3) why some gases are monatomic
   (4) why some gases are diatomic

38. Under which conditions of temperature and pressure would helium behave most like an ideal gas?
   (1) 50 K and 20 kPa
   (2) 50 K and 600 kPa
   (3) 750 K and 20 kPa
   (4) 750 K and 600 kPa

39. The kinetic molecular theory assumes that the particles of an ideal gas
   (1) are in random, constant, straight-line motion
   (2) are arranged in a regular geometric pattern
   (3) have strong attractive forces between them
   (4) have collisions that result in the system losing energy

40. Which statement describes the particles of an ideal gas according to the kinetic molecular theory?
   (1) The gas particles are arranged in a regular geometric pattern.
   (2) The gas particles are in random, constant, straight-line motion.
   (3) The gas particles are separated by very small distances, relative to their sizes.
   (4) The gas particles are strongly attracted to each other.

41. Which temperature change would cause a sample of an ideal gas to double in volume while the pressure is held constant?
   (1) from 400. K to 200. K
   (2) from 200. K to 400. K
   (3) from 400. °C to 200. °C
   (4) from 200. °C to 400. °C

**Separating by Filtration (4 Questions)**

- If one of the substances is insoluble, the mixture can be separated by making a solution with the soluble substance and then using a piece of filter paper to remove the insoluble substance.

42. At room temperature, a mixture of sand and water can be separated by
   (1) ionization
   (2) combustion
   (3) filtration
   (4) sublimation

43. Given a mixture of sand and water, state one process that can be used to separate water from the sand. [1]

**Separating by Temperature (13 Questions)**

- If the substances have different boiling or freezing points, the two or more substances can be separated by distillation, boiling, vaporization (same as boiling), condensation, freezing or by melting.

44. Which property makes it possible to separate the oxygen and the nitrogen from a sample of liquefied air?
   (1) boiling point
   (2) conductivity
   (3) hardness
   (4) electronegativity

45. Which physical property makes it possible to separate the components of crude oil by means of distillation?
   (1) melting point
   (2) conductivity
   (3) solubility
   (4) boiling point

46. Which sample of matter can be separated into different substances by physical means?
   (1) LiCl(aq)
   (2) LiCl(s)
   (3) NH₃(g)
   (4) NH₃(l)

47. The laboratory process of distillation does **not** involve
   (1) changing a liquid to vapor
   (2) changing a vapor to liquid
   (3) liquids with different boiling points
   (4) liquids with the same boiling points

48. Which two physical properties allow a mixture to be separated by chromatography?
   (1) hardness and boiling point
   (2) density and specific heat capacity
   (3) malleability and thermal conductivity
   (4) solubility and molecular polarity
5A-1 Allotropes  (16 Questions)
• The different forms an element can exist in the same phase.
  - The different forms have different chemical and physical properties.
• Allotropes of 3 Elements

<table>
<thead>
<tr>
<th>Key</th>
</tr>
</thead>
<tbody>
<tr>
<td>● = atom of element X</td>
</tr>
<tr>
<td>○ = atom of element Y</td>
</tr>
<tr>
<td>◼ = atom of element Z</td>
</tr>
</tbody>
</table>

1. Which statement describes oxygen gas, \( \text{O}_2(g) \), and ozone gas, \( \text{O}_3(g) \)?

(1) They have different molecular structures, only.
(2) They have different properties, only.
(3) They have different molecular structures and different properties.
(4) They have the same molecular structure and the same properties.

2. The carbon atoms in graphite and the carbon atoms in diamond have different
   (1) atomic numbers
   (2) atomic masses
   (3) electronegativities
   (4) structural arrangements

3. Which diagram represents a mixture of two different molecular forms of the same element?

<table>
<thead>
<tr>
<th>Key</th>
</tr>
</thead>
<tbody>
<tr>
<td>● = atom of element X</td>
</tr>
<tr>
<td>○ = atom of element Z</td>
</tr>
</tbody>
</table>

   (1) ![Diagram 1]  (2) ![Diagram 2]  (3) ![Diagram 3]  (4) ![Diagram 4]

5B-1 Periodic Groups  (14 Questions)
• Elements in the same family (group) have similar chemical properties. They have the same # of valence electrons.

Base your answer to question 4 on the information below.

Elements with atomic numbers 112 and 114 have been produced and their IUPAC names are pending approval. However, an element that would be put between these two elements on the Periodic Table has not yet been produced. If produced, this element will be identified by the symbol Uut until an IUPAC name is approved.

4. Identify one element that would be chemically similar to Uut. [1]

   ____________________________

5. Which list includes elements with the most similar chemical properties?

(1) Br, Ga, Hg  (2) Cr, Pb, Xe  (3) O, S, Se  (4) N, O, F

6. The chemical properties of calcium are most similar to the chemical properties of

(1) Ar  (2) K  (3) Mg  (4) Sc

7. Which element has chemical properties that are most similar to the chemical properties of sodium?

(1) beryllium  (2) calcium  (3) lithium  (4) magnesium

5B-2 Metals-Nonmetals & PT  (24 Questions)
• Metals are to the left of the periodic table, nonmetals are to the right and metalloids are between.
• Metals are malleable and can conduct electricity.
• Metalloids include B, Si, Ge, As, Sb and Te, all next to the staircase separating the metals from the nonmetals.

8. Which list of elements contains two metalloids?

(1) Si, Ge, Po, Pb  (2) As, Bi, Br, Kr  (3) Si, P, S, Cl  (4) Po, Sb, I, Xe

9. Which element has both metallic and nonmetallic properties?

(1) Rb  (2) Rn  (3) Si  (4) Sr

10. The element in Period 4 and Group 14 of the Periodic Table would be classified as a

(1) metal  (2) metalloid  (3) nonmetal  (4) noble gas

11. Which Group 14 element is a metalloid?

(1) tin  (2) silicon  (3) lead  (4) carbon
5B-3 Valence Electrons & PT  (35 Questions)
• The valence electrons can be used to determine an unknown element in a compound.
- The total positive valence number must equal the total negative valence number in each compound.
- Subscripts X the valence number = total valance for each element in a compound.
• Most stable valance arrangements are those of the Nobel gases.
• Reason for similar properties within each group or family - they have the same valence numbers

12. Which set of symbols represents atoms with valence electrons in the same electron shell?
(1) Ba, Br, Bi (3) 0, S, Te
(2) Sr, Sn, I (4) Mn, Hg, Cu

13. In the formula X₂O₅, the symbol X could represent an element in Group
(1) 1 (3) 15
(2) 2 (4) 18

14. In the ground state, which atom has a completely filled valence electron shell?
(1) C (3) Ne
(2) V (4) Sb

15. Identify the element in Period 3 of the Periodic Table that reacts with oxygen to form an ionic compound represented by the formula X₂O. [1]

16. Magnesium and calcium have similar chemical properties because a magnesium atom and a calcium atom have the same
(1) atomic number
(2) mass number
(3) total number of electron shells
(4) total number of valence electrons

5B-4 Nobel Gases & PT  (5 Questions)
• Nobel gases are not normally reactive and are found in Group 18
- They are not normally reactive because their valence energy level is full (He - 2 electrons, Other noble gases - 8 electrons)

17. Which list of elements contains a metal, a metalloid, a nonmetal, and a noble gas?
(1) Be, Si, Cl, Kr (3) K, Fe, B, F
(2) C, N, Ne, Ar (4) Na, Zn, As, Sb

18. Which element is a noble gas?
(1) krypton (3) antimony
(2) chlorine (4) manganese

5C-1 Electronegativity & PT-S  (13 Questions)
• How strong is the nucleus of one atom is attracted to electrons bonded to another atom. The electronegativity for each element is found in Table S. In general:
- Electronegativity increases as one moves across the periodic table from left to right.
- Electronegativity decreases as one moves down a group.

19. Which of the following atoms has the greatest tendency to attract electrons?
(1) barium (3) boron
(2) beryllium (4) bromine

20. Which of the following elements has the highest electronegativity?
(1) H (3) Al
(2) K (4) Ca

21. Which atom has the weakest attraction for the electrons in a bond with an H atom?
(1) Cl atom (3) O atom
(2) F atom (4) S atom

22. What is the most likely electronegativity value for a metallic element?
(1) 1.3 (3) 3.4
(2) 2.7 (4) 4.0

5C-2 Phase-temperature & PT-S  (25 Questions)
• Each element has its own melting point and boiling point at standard pressure. They are listed in Table S.
- Solids exist below the melting point
- Liquids exist between the melting & boiling point
- Gases exist above the boiling point

23. At STP, which element is solid, brittle, and a poor conductor of electricity?
(1) Al (3) Ne
(2) K (4) S

24. Which element is a liquid at 758 K and standard pressure?
(1) gold (3) platinum
(2) silver (4) thallium
25. Which element is a liquid at STP?
   (1) argon (3) chlorine
   (2) bromine (4) sulfur

26. At standard pressure, which element has a freezing point below standard temperature?
   (1) In (3) Hf
   (2) Ir (4) Hg

27. Which element is a liquid at 305 K and 1.0 atm?
   (1) magnesium (3) gallium
   (2) fluorine (4) iodine

28. Which Group 15 element exists as diatomic molecules at STP?
   (1) phosphorus (3) bismuth
   (2) nitrogen (4) arsenic

5C-3 Density & PT-S (4 Questions)
- Density is the measure of mass/volume. Each Element density is listed in Table S.
  - A typical unit of density is g/cm³.

29. Which element has the greatest density at STP?
   (1) barium (3) magnesium
   (2) beryllium (4) radium

30. A 10.0-gram sample of which element has the smallest volume at STP?
   (1) aluminum (3) titanium
   (2) magnesium (4) zinc

5C-4 Ionization & PT-S (9 Questions)
- The Ionization energy is the amount of energy needed to remove the most loosely bounded electron of a neutral atom in its gaseous state. The Ionization energy for each element is found in Table S. The lower the ionization energy, the easier for the electron to be lost to another atom. In general:
  - Ionization energy increases as one moves across the periodic table from left to right.
  - Ionization energy decreases as one moves down a group.

31. In the ground state, each atom of an element has two valence electrons. This element has a lower first ionization energy than calcium. Where is this element located on the Periodic Table?
   (1) Group 1, Period 4 (3) Group 2, Period 3
   (2) Group 2, Period 5 (4) Group 3, Period 4

32. Samples of four Group 15 elements, antimony, arsenic, bismuth, and phosphorus, are in the gaseous phase. An atom in the ground state of which element requires the least amount of energy to remove its most loosely held electron?
   (1) As (3) P
   (2) Bi (4) Sb

5C-5 Atomic Radius & PT-S (8 Questions)
- The atomic radius is a measure of the size of an atom. Each element atomic radius is listed in Table S. In general:
  - Atomic Radius decreases as one moves across the periodic table from left to right.
  - Increase pull from extra protons causes the electrons to move in closer,
  - Atomic Radius increases as one moves down a group.
  - Valence electrons are shielded by inner electrons.

33. Which grouping of circles, when considered in order from the top to the bottom, best represents the relative size of the atoms of Li, Na, K, and Rb, respectively?

34. How do the atomic radius and metallic properties of sodium compare to the atomic radius and metallic properties of phosphorus?
   (1) Sodium has a larger atomic radius and is more metallic.
   (2) Sodium has a larger atomic radius and is less metallic.
   (3) Sodium has a smaller atomic radius and is more metallic.
   (4) Sodium has a smaller atomic radius and is less metallic.
Base your answers to questions 35 through 37 on the data in Reference Table S.

35. On the data table below, record the boiling points for He, Ne, Ar, Kr, and Xe. [1]

<table>
<thead>
<tr>
<th>Symbol</th>
<th>Atomic Number</th>
<th>Boiling Point (K)</th>
</tr>
</thead>
<tbody>
<tr>
<td>He</td>
<td>2</td>
<td></td>
</tr>
<tr>
<td>Ne</td>
<td>10</td>
<td></td>
</tr>
<tr>
<td>Ar</td>
<td>18</td>
<td></td>
</tr>
<tr>
<td>Kr</td>
<td>36</td>
<td></td>
</tr>
<tr>
<td>Xe</td>
<td>54</td>
<td></td>
</tr>
</tbody>
</table>

36. On the grid below, plot the boiling point versus the atomic number for He, Ne, Ar, Kr, and Xe. Circle and connect the points. [1]

Example:

37. Based on your graph, describe the trend in the boiling points of these elements as the atomic number increases. [1]

5E-1 Z & Periodic Table (11 Questions)

- The periodic table is arranged according to increasing atomic number

38. The elements in the Periodic Table are arranged in order of increasing (1) atomic number (3) mass number (2) atomic radius (4) neutron number

39. The elements on the Periodic Table are arranged in order of increasing (1) boiling point (3) atomic number (2) electronegativity (4) atomic mass

5E-2 Electronegativity Trends & PT (5 Questions)

- Electronegativity is a measure of the strength a nucleus of one atom is attracted to electrons bonded to another atom. The electronegativity for each element is found in Table S. In general:
  - Electronegativity increases as one moves across the periodic table from left to right.
  - Electronegativity decreases as one moves down a group.

40. Which general trend is found in Period 3 as the elements are considered in order of increasing atomic number? (1) increasing atomic radius (2) increasing electronegativity (3) decreasing atomic mass (4) decreasing first ionization energy

41. As the elements in Group 17 are considered in order of increasing atomic number, the chemical reactivity of each successive element decreases remains the same (1) decreases (2) increases
5E-3 Ionization Trends & PT  (7 Questions)

- The ionization energy is the amount of energy needed to remove the most loosely bounded electron of a neutral atom in its gaseous state. The ionization energy for each element is found in Table S. The lower the ionization energy, the easier for the electron to be lost to another atom. In general:
  - Ionization energy increases as one moves across the periodic table from left to right.
  - Ionization energy decreases as one moves down a group.

42. Which general trend is found in Period 2 on the Periodic Table as the elements are considered in order of increasing atomic number?
   (1) decreasing atomic mass
   (2) decreasing electronegativity
   (3) increasing atomic radius
   (4) increasing first ionization energy

43. State the general trend in first ionization energy for the elements in Group 2 as these elements are considered in order from top to bottom in the group. [1]

5E-4 Atomic Radius Trends & PT  (10 Questions)

- The atomic radius is a measure of the size of an atom. Each element atomic radius is listed in Table S. In general:
  - Atomic Radius decreases as one moves across the periodic table from left to right.
  - Increase pull from extra protons causes the electrons to move in closer.
  - Atomic Radius increases as one moves down a group.
  - Valence electrons are shielded by inner electrons.

44. As atomic number increases within Group 15 on the Periodic Table, atomic radius
   (1) decreases, only
   (2) increases, only
   (3) decreases, then increases
   (4) increases, then decreases

45. Which characteristics both generally decrease when the elements in Period 3 on the Periodic Table are considered in order from left to right?
   (1) nonmetallic properties and atomic radius
   (2) nonmetallic properties and ionization energy
   (3) metallic properties and atomic radius
   (4) metallic properties and ionization energy

5E-5 Metallic Properties & PT  (6 Questions)

- The further left and down on the table, the greater the metallic properties.
  - In general, the metals and nonmetals are separated by the staircase that starts at boron and ends at astatine.

46. The elements located in the lower left corner of the Periodic Table are classified as
   (1) metals
   (2) nonmetals
   (3) metalloids
   (4) noble gases

47. Which characteristics both generally decrease when the elements in Period 3 on the Periodic Table are considered in order from left to right?
   (1) nonmetallic properties and atomic radius
   (2) nonmetallic properties and ionization energy
   (3) metallic properties and atomic radius
   (4) metallic properties and ionization energy

5F-1 Ion Radius Size  (24 Questions)

- If the atom looses an electron (becomes +), the ion is smaller than the atom. If the atom gains an electron (becomes -), the ion is larger than the atom.

48. What occurs when an atom loses an electron?
   (1) The atom's radius decreases and the atom becomes a negative ion.
   (2) The atom's radius decreases and the atom becomes a positive ion.
   (3) The atom's radius increases and the atom becomes a negative ion.
   (4) The atom's radius increases and the atom becomes a positive ion.

49. When an atom becomes a positive ion, the radius of the atom
   (1) decreases
   (2) increases

50. Compared to a phosphorus atom, a P^3– ion has
   (1) more electrons and a larger radius
   (2) more electrons and a smaller radius
   (3) fewer electrons and a larger radius
   (4) fewer electrons and a smaller radius
5F-2 Ion Electron number (7 Questions)

• If an atom gains electrons, it becomes a negative ion. If an atom loses electrons, the atom becomes a positive ion.

Base your answer to question 51 on the information below.

Potassium ions are essential to human health. The movement of dissolved potassium ions, $\text{K}^+ (\text{aq})$, in and out of a nerve cell allows that cell to transmit an electrical impulse.

51. What is the total number of electrons in a potassium ion? [1]

Ans: __________________________ electrons

52. Which change occurs when a barium atom loses two electrons?

(1) It becomes a negative ion and its radius decreases.
(2) It becomes a negative ion and its radius increases.
(3) It becomes a positive ion and its radius decreases.
(4) It becomes a positive ion and its radius increases.

53. An atom of an element forms a $2^+\text{ion}$. In which group on the Periodic Table could this element be located?

(1) 1 (3) 13
(2) 2 (4) 17
6A-1 Bond Energy (14 Questions)

• When bonds break, energy is absorbed
  - Heat is a reactant - endothermic.
• When bonds form, energy is released
  - Heat is a product - exothermic.
  - The greater the energy released, the more stable the compound formed.

1. What occurs when an atom of chlorine and an atom of hydrogen become a molecule of hydrogen chloride?
   (1) A chemical bond is broken and energy is released.
   (2) A chemical bond is broken and energy is absorbed.
   (3) A chemical bond is formed and energy is released.
   (4) A chemical bond is formed and energy is absorbed.

2. The balanced equation below represents a molecule of bromine separating into two bromine atoms.
   \( \text{Br}_2 \rightarrow \text{Br} + \text{Br} \)
   What occurs during this change?
   (1) Energy is absorbed and a bond is formed.
   (2) Energy is absorbed and a bond is broken.
   (3) Energy is released and a bond is formed.
   (4) Energy is released and a bond is broken.

6A-2 Octet formation (5 Questions)

• Atoms are most stable when the outer electron energy level contains 8 electrons (Octet).
  - Exceptions - light elements (H, He, Li & Be) outer energy level will contain 2.

3. An atom of argon rarely bonds to an atom of another element because an argon atom has
   (1) 8 valence electrons
   (2) 2 electrons in the first shell
   (3) 3 electron shells
   (4) 22 neutrons

4. An atom in the ground state has a stable valence electron configuration. This atom could be an atom of
   (1) Al
   (2) Cl
   (3) Na
   (4) Ne

6B-1 Lewis Dot Atoms (19 Questions)

• The atom symbols are drawn surrounded with dots representing its valence electrons.
  - The number of valence electrons can be found in the periodic table.
  - Examples:
    \( \text{Na} \quad \text{Mg} \quad \text{Al} \quad \text{Si} \quad \text{P} \quad \text{S} \quad \text{Cl} \quad \text{Ar} \)

5. Which Lewis electron-dot diagram represents a boron atom in the ground state?
   (1) \( \cdot \text{B} \)
   (2) \( \cdot \text{B} \)
   (3) \( \cdot \text{B} \)
   (4) \( \cdot \text{B} \)

6. In the box below, draw a Lewis electron-dot diagram for an atom of boron. [1]

7. Which Lewis electron-dot structure is drawn correctly for the atom it represents?
   (1) \( \cdot \text{N} \)
   (2) \( \cdot \text{F} \)
   (3) \( \cdot \text{O} \)
   (4) \( \cdot \text{Ne} \)

8. In the space below, draw a Lewis electron-dot diagram for an atom of sulfur-33. [1]
6B-2 Lewis Dot Ions  (6 Questions)

- Ions are drawn with brackets with the charged indicated in the upper right corner.
  - Example: Fluorine ion - [:F:]

9. Given the equation:

\[
\text{[:F:] + 1e}^{-} \rightarrow [:F:]^{-}
\]

This equation represents the formation of a
(1) fluoride ion, which is smaller in radius than a fluorine atom
(2) fluoride ion, which is larger in radius than a fluorine atom
(3) fluorine atom, which is smaller in radius than a fluoride ion
(4) fluorine atom, which is larger in radius than a fluoride ion

Base your answer to question 10 on the information below.

When a person perspires (sweats), the body loses many sodium ions and potassium ions. The evaporation of sweat cools the skin.

After a strenuous workout, people often quench their thirst with sports drinks that contain NaCl and KCl. A single 250.-gram serving of one sports drink contains 0.055 gram of sodium ions.

10. In the space below, draw a Lewis electron-dot diagram for one of the positive ions lost by the body as a person perspires. [1]

6B-3 Lewis Dot Molecules  (19 Questions)

- Bonding between atoms may be represented by lines. Each line represents a shared pair (two) of electrons.
  - Example 1: Oxygen (O₂) -

- Example 2: Methane (CH₄) -

11. Given the structural formula:

H – C≡C – H

What is the total number of electrons shared in the bond between the two carbon atoms?
(1) 6 (3) 3
(2) 2 (4) 4

12. Given the Lewis electron-dot diagram:

\[
\text{H} \quad \text{H} \quad \text{H} \quad \text{H}
\]

Which electrons are represented by all of the dots?
(1) the carbon valence electrons, only
(2) the hydrogen valence electrons, only
(3) the carbon and hydrogen valence electrons
(4) all of the carbon and hydrogen electrons

Base your answer to question 13 on your knowledge of chemical bonding and on the Lewis electron-dot diagrams of H₂S, CO₂, and F₂ below.

13. Which atom, when bonded as shown, has the same electron configuration as an atom of argon? [1]

Answer: ______________________

14. Given a formula for oxygen:

\[
\text{O} = \text{O}
\]

What is the total number of electrons shared between the atoms represented in this formula?
(1) 1 (3) 8
(2) 2 (4) 4
6C-1 Ionic Bond Traits  (17 Questions)

- Electrons are transferred, forming ions that have strong attraction between each other.
- Ionic compounds generally have high melting and boiling points, are hard, do not conduct electricity in their solid phase, but will conduct electricity in their liquid phase (melted) and when dissolved in water.

15. A substance that does not conduct electricity as a solid but does conduct electricity when melted is most likely classified as
   (1) an ionic compound  (2) a molecular compound  (3) a metal  (4) a nonmetal

16. Which type of bond results when one or more valence electrons are transferred from one atom to another?
   (1) a hydrogen bond  (2) an ionic bond  (3) a nonpolar covalent bond  (4) a polar covalent bond

17. Which substance contains bonds that involved the transfer of electrons from one atom to another?
   (1) CO₂  (2) NH₃  (3) KBr  (4) Cl₂

6C-2 Covalent Bond Traits  (23 Questions)

- Electrons are shared between the atoms of the compound, forming molecules.
  - Each bond contains a pair (two) of electrons
- Covalent compounds (Molecular compounds) are generally soft with low melting and boiling points and they do not conduct electricity.

19. An oxygen molecule contains a double bond because the two atoms of oxygen share a total of
   (1) 1 electron  (2) 2 electrons  (3) 3 electrons  (4) 4 electrons

20. Covalent bonds are formed when electrons are transferred from one atom to another
   (1) captured by the nucleus  (2) transferred from one atom to another  (3) mobile within a metal  (4) shared between two atoms

6C-3 Metallic Bond Traits  (7 Questions)

- The metallic bond is a crystalline lattice surrounded by a sea of moveable electrons.
  - Metals are generally hard substances, but malleable (can be bent or banged into a different shape while in the solid phase) with high melting and boiling points, and are excellent conductors of electricity because their electrons are able to flow (move).

23. A solid substance is an excellent conductor of electricity. The chemical bonds in this substance are most likely
   (1) ionic, because the valence electrons are shared between atoms  (2) ionic, because the valence electrons are mobile
   (3) metallic, because the valence electrons are stationary  (4) metallic, because the valence electrons are mobile

18. A chemist performs the same tests on two homogeneous white crystalline solids, A and B. The results are shown in the table above in the next column.

<table>
<thead>
<tr>
<th></th>
<th>Solid A</th>
<th>Solid B</th>
</tr>
</thead>
<tbody>
<tr>
<td>Melting Point</td>
<td>High, 801°C</td>
<td>Low, decomposes at 186°C</td>
</tr>
<tr>
<td>Solubility in H₂O (grams per 100.0 g H₂O at 0°C)</td>
<td>36.7</td>
<td>3.2</td>
</tr>
<tr>
<td>Electrical Conductivity (in aqueous solution)</td>
<td>Good conductor</td>
<td>Nonconductor</td>
</tr>
</tbody>
</table>

The results of these tests suggest that
   (1) both solids contain only ionic bonds
   (2) both solids contain only covalent bonds
   (3) solid A contains only covalent bonds and solid B contains only ionic bonds
   (4) solid A contains only ionic bonds and solid B contains only covalent bonds

21. Which characteristic is a property of molecular substances?
   (1) good heat conductivity  (2) good electrical conductivity  (3) low melting point  (4) high melting point

22. Which formula represents a molecular compound?
   (1) Kr  (2) LiOH  (3) N₂O₄  (4) NaI

24. Describe one appropriate laboratory test that can be used to determine the malleability of a solid sample of an element at room temperature. [1]
**6D-1 Ionic Bond Pick (7 Questions)**

- Ionic bonds are usually between a metal and nonmetal.
- The difference in electronegativity between two elements in the compound is greater than 1.7.

<table>
<thead>
<tr>
<th>Bond Type</th>
<th>Melting and boiling points</th>
<th>Hardness</th>
<th>Conductivity</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ionic</td>
<td>High</td>
<td>Hard</td>
<td>No</td>
</tr>
</tbody>
</table>

* Free ions carry current. When ionic compounds dissolve in water, the ions separate and are free to move. Likewise, when ionic compounds melt, the now free ions can carry current.

25. Which type of bond is found in sodium bromide?
   - (1) covalent
   - (2) hydrogen
   - (3) ionic
   - (4) metallic

26. Which formula represents an ionic compound?
   - (1) H₂
   - (2) CH₄
   - (3) CH₂OH
   - (4) NH₄Cl

27. An ionic compound is formed when there is a reaction between the elements
   - (1) strontium and chlorine
   - (2) hydrogen and chlorine
   - (3) nitrogen and oxygen
   - (4) sulfur and oxygen

28. Which element forms an ionic compound when it reacts with lithium?
   - (1) K
   - (2) Fe
   - (3) Kr
   - (4) Br

**6D-2 Covalent Bond Pick (11 Questions)**

- The difference in electronegativity between two elements in the compound is less than or equal to 1.7.

<table>
<thead>
<tr>
<th>Bond Type</th>
<th>Melting and boiling points</th>
<th>Hardness</th>
<th>Conductivity</th>
</tr>
</thead>
<tbody>
<tr>
<td>Covalent</td>
<td>Low</td>
<td>Soft**</td>
<td>No</td>
</tr>
</tbody>
</table>

** Most covalent compounds are soft. However Tungsten carbide, Silicon carbide and Diamonds are the three hardest substances and they are all have a crystalline lattice held together by covalent bonds.

- The two atoms in all diatomic gas molecules are held together by covalent bonds. Cl₂, F₂, Br₂, and I₂ share 1 pair of electrons (single bond) while O₂ shares two pairs of electrons (double bond) , and N₂ shares 3 pairs of electrons (triple bond).

29. Which two substances are covalent compounds?
   - (1) C₂H₁₂O₆(s) and KI(s)
   - (2) C₂H₁₂O₆(s) and HCl(g)
   - (3) KI(s) and NaCl(s)
   - (4) NaCl(s) and HCl(g)

30. Which compound contains only covalent bonds?
   - (1) NaOH
   - (2) Ba(OH)₂
   - (3) Ca(OH)₂
   - (4) CH₃OH

31. Which type of chemical bond is formed between two atoms of bromine?
   - (1) metallic
   - (2) hydrogen
   - (3) ionic
   - (4) covalent

32. Which molecule contains a triple covalent bond?
   - (1) H₂
   - (2) N₂
   - (3) O₂
   - (4) Cl₂

**6D-3 Metallic Bond Pick (3 Questions)**

- Is it a compound or a solid solution (homogenesis mixture)?

<table>
<thead>
<tr>
<th>Bond Type</th>
<th>Melting and boiling points</th>
<th>Hardness</th>
<th>Conductivity</th>
</tr>
</thead>
<tbody>
<tr>
<td>Metallic</td>
<td>High</td>
<td>Hard</td>
<td>Yes</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Conductivity</th>
</tr>
</thead>
<tbody>
<tr>
<td>Solid</td>
</tr>
<tr>
<td>liquid</td>
</tr>
<tr>
<td>Aqueous</td>
</tr>
</tbody>
</table>

33. Which type of bond is found between atoms of solid cobalt?
   - (1) nonpolar covalent
   - (2) polar covalent
   - (3) metallic
   - (4) ionic

34. Which substance contains metallic bonds?
   - (1) Hg(ℓ)
   - (2) H₂O(ℓ)
   - (3) NaCl(s)
   - (4) C₂H₁₂O₆(s)
6D-4 Polyatomic Pick  (12 Questions)

- Polyatomic ions consist of elements that are held together by covalent bonds. They react with other ions, forming a compound that has both ionic and covalent bonds. Table E list the polyatomic ions that are used in this course.

35. Which compound contains both ionic and covalent bonds?
(1) CaCO₃  (3) MgF₂  
(2) PCl₃  (4) CH₂O

36. The bonds in the compound MgSO₄ can be described as
(1) ionic, only  (2) covalent, only  
(3) both ionic and covalent  (4) neither ionic nor covalent

37. The chemical bonding in sodium phosphate, Na₃PO₄, is classified as
(1) ionic, only  (2) metallic, only  
(3) both ionic and covalent  (4) both ionic and metallic

38. Which compound contains both ionic and covalent bonds?
(1) ammonia  (3) sodium nitrate  
(2) methane  (4) potassium chloride

6E-1 Polarity and Bonds  (16 Questions)

- Atoms of the same element form nonpolar bonds. That is because the electronegativity difference is 0. Atoms of different elements may form polar bonds. The greater the difference between the atoms of the bond, the greater the polarity.
  - Electronegativities of selected elements are found in Table S.

39. Which bond is least polar?
(1) As-Cl  (3) P-Cl  
(2) Bi-Cl  (4) N-Cl

40. Which of these formulas contains the most polar bond?
(1) H—Br  (3) H—F  
(2) H—Cl  (4) H—I

41. Which molecule has a nonpolar covalent bond?
\[ \text{H—H} \quad \text{H—N} \quad \text{H—O} \quad \text{H—Cl} \]
(1)  (2)  (3)  (4)

42. The bonds between hydrogen and oxygen in a water molecule are classified as
(1) polar covalent  (2) nonpolar covalent  
(3) ionic  (4) metallic

6E-2 Polarity and Molecules  (19 Questions)

- Nonpolar molecules may contain polar bonds because the charge is distributed symmetrically due to its structure.
  - Symmetrical - nonpolar, even if it has polar bonds (different electronegativities).
  - Nonsymmetrical - It is going to be polar (H₂O and NH₃ are often used as examples.)

43. Given the formula representing a molecule:
\[ \text{H—C≡C—H} \]
The molecule is
(1) symmetrical and polar  (2) symmetrical and nonpolar  
(3) asymmetrical and polar  (4) asymmetrical and nonpolar

44. Why is a molecule of CO₂ nonpolar even though the bonds between the carbon atom and the oxygen atoms are polar?
(1) The shape of the CO₂ molecule is symmetrical.  
(2) The shape of the CO₂ molecule is asymmetrical.  
(3) The CO₂ molecule has a deficiency of electrons.  
(4) The CO₂ molecule has an excess of electrons.

45. Which pair of characteristics describes the molecule illustrated below?
\[ \text{H—O—H} \]
(1) symmetrical and polar  (2) symmetrical and nonpolar  
(3) asymmetrical and polar  (4) asymmetrical and nonpolar

46. Which formula represents a polar molecule?
(1) Br₂  (3) CH₄  
(2) CO₂  (4) NH₃
Hydrogen bonding only occurs with NH₃, HF and H₂O. It increases the boiling point of these substances because the molecules are very attracted to one another.

47. Which intermolecular force of attraction accounts for the relatively high boiling point of water?
   (1) hydrogen bonding   (3) metallic bonding
   (2) covalent bonding   (4) ionic bonding

48. Which of the following compounds has the highest boiling point?
   (1) H₂O   (3) H₂Se
   (2) H₂S   (4) H₂Te

6F-2 Intermolecular Forces (10 Questions)

• Stronger forces help hold the substance together, increasing melting and boiling points.

49. Which statement explains why Br₂ is a liquid at STP and I₂ is a solid at STP?
   (1) Molecules of Br₂ are polar, and molecules of I₂ are nonpolar.
   (2) Molecules of I₂ are polar, and molecules of Br₂ are nonpolar.
   (3) Molecules of Br₂ have stronger intermolecular forces than molecules of I₂.
   (4) Molecules of I₂ have stronger intermolecular forces than molecules of Br₂.

50. Which compound has hydrogen bonding between its molecules?
   (1) CH₄   (3) KH
   (2) CaH₂   (4) NH₃

51. Which of these substances has the strongest intermolecular forces?
   (1) H₂O   (3) H₂Se
   (2) H₂S   (4) H₂Te
7 A - 1 Solvents, Solutes and Solutions (9 Questions)

- A solvent dissolves the solute to form a solution.
- Aqueous (aq) Solutions - water is the solvent.
- Liquid solutions characteristics:
  - Solutions do not settle when left standing
  - Solutions are clear but may have color,
  - Light can pass through a solution without being dispersed
  - Solutions are Homogeneous mixtures
  - Solutions can pass through a fine filter

1. When a mixture of water, sand, and salt is filtered, what passes through the filter paper?
   (1) water, only
   (2) water and sand, only
   (3) water and salt, only
   (4) water, sand, and salt

2. A sample is prepared by completely dissolving 10.0 grams of NaCl in 1.0 liter of H₂O. Which classification best describes this sample?
   (1) homogeneous compound
   (2) homogeneous mixture
   (3) heterogeneous compound
   (4) heterogeneous mixture

7A-2 Solubility Factors (19 Questions)

- Temperature
  - Most solids become more soluble in liquids when temperature increases (There are exceptions!)
  - Gases become more soluble in liquids when temperature decreases.
- Pressure
  - Little or no effect on solubility of solids in liquids.
  - The solubility of gases in liquids increase as the pressure increases.
    - That is why bubbles of gas leave soda when the pressure is released by opening the cap.
- Nature of Solvent and solute
  - Like dissolves like
    - Polar solvents will dissolve polar & ionic solutes
      - Example water dissolves salt (NaCl)
    - Nonpolar solvents will dissolve nonpolar molecules
      - Oils will dissolve grease
  - Note: Soap molecules are long molecules that have both polar and nonpolar ends. That is why grease can be dissolved by soppy water

3. At room temperature, the solubility of which solute in water would be most affected by a change in pressure?
   (1) methanol
   (2) sugar
   (3) carbon dioxide
   (4) sodium nitrate

4. Under which conditions of temperature and pressure is a gas most soluble in water?
   (1) high temperature and low pressure
   (2) high temperature and high pressure
   (3) low temperature and low pressure
   (4) low temperature and high pressure

5. The solubility of KClO₃(s) in water increases as the
   (1) temperature of the solution increases
   (2) temperature of the solution decreases
   (3) pressure on the solution increases
   (4) pressure on the solution decreases

6. At standard pressure, a certain compound has a low boiling point and is insoluble in water. At STP, this compound most likely exists as
   (1) ionic crystals
   (2) metallic crystals
   (3) nonpolar molecules
   (4) polar molecules

7B-1 Solubility Graphs (Reference table G) (26 Questions)

- Used to determine the amount of solute will dissolve in 100 grams of water between 0°C and 100°C.
- Besides determining the amount of solute that should be dissolved, three conditions can be inferred if you know exactly how much solute is dissolved in the water.
- Unsaturated - the solvent contains less solute than the maximum it can hold at a given temperature.
  - You can add some more solute, some of it would dissolve.
- Saturated - the solvent contains the maximum amount it can hold at a given temperature.
  - If you add more solute, that amount would remain separate from the solution.
- Supersaturated - the solvent contains more than the maximum amount it should be able hold at a given temperature.
- this situation occasionally occurs when a saturated solution temperature decreases and no expected crystals of solute are formed.
- If you add more solute, it would act as a seed, causing extra solute to come out of solution.

Notes:
• SO₂, NH₃ and HCl are gases. As the temperature of the solution increases, less gas is able to be held by the solvent (water).
• Temperature has a much greater effect on solubility of KNO₃ than NaCl.
• For any given substance on the graph, if the value of the grafted point is:
  - on the line - saturated solution
  - below the line - Unsaturated solution
  - above the line - Supersaturated solution

Base your answers to questions 7 and 8 on the information below.

A student uses 200 grams of water at a temperature of 60°C to prepare a saturated solution of potassium chloride, KCl.

7. According to Reference Table G, how many grams of KCl must be used to create this saturated solution? [1]

Answer: ____________________ grams

8. This solution is cooled to 10°C and the excess KCl precipitates (settles out). The resulting solution is saturated at 10°C. How many grams of KCl precipitated out of the original solution? [1]

Answer: ____________________ grams

Base your answer to question 9 on the information below.

In a laboratory, a student makes a solution by completely dissolving 80.0 grams of KNO₃(s) in 100.0 grams of hot water. The resulting solution has a temperature of 60°C. The room temperature in the laboratory is 22°C.

9. Classify, in terms of saturation, the type of solution made by the student. [1]

10. What is the mass of KNO₃(s) that must dissolve in 100. grams of water to form a saturated solution at 50.°C? [1]

________________________ g

11. What is the mass of KNO₃(s) that must dissolve in 100. grams of water to form a saturated solution at 50.°C? [1]

________________________ g

12. An unsaturated solution is formed when 80. grams of a salt is dissolved in 100. grams of water at 40.°C. This salt could be

(1) KCl  (3) NaCl
(2) KNO₃  (4) NaNO₃

13. Which compound is least soluble in water at 60.°C? [1]

(1) KClO₃  (3) NaCl
(2) KNO₃  (4) NH₄Cl

14. Which compound becomes less soluble in water as the temperature of the solution is increased? [1]

(1) HCl  (3) NaCl
(2) KCl  (4) NH₄Cl

15. At standard pressure, which substance becomes less soluble in water as temperature increases from 10.°C to 80.°C? [1]

(1) HCl  (3) NaCl
(2) KCl  (4) NH₄Cl

16. One hundred grams of water is saturated with NH₄Cl at 50°C. According to Table G, if the temperature is lowered to 10°C, what is the total amount of NH₄Cl that will precipitate? [1]

(1) 5.0 g  (3) 30. g
(2) 17 g  (4) 50. g
7B-2  Solubility Tables (13 Questions)

Table F
Solubility Guidelines

Be Careful:
Left side - Ions are soluble with exceptions
Right Side - Ions are insoluble with exceptions

<table>
<thead>
<tr>
<th>Ions That Form Soluble Compounds</th>
<th>Exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>Group 1 ions (Li⁺, Na⁺, etc.)</td>
<td></td>
</tr>
<tr>
<td>ammonium (NH₄⁺)</td>
<td></td>
</tr>
<tr>
<td>nitrate (NO₃⁻)</td>
<td></td>
</tr>
<tr>
<td>acetate (C₂H₃O₂⁻ or CH₃COO⁻)</td>
<td></td>
</tr>
<tr>
<td>hydrogen carbonate (HCO₃⁻)</td>
<td></td>
</tr>
<tr>
<td>chlorate (ClO₅⁻)</td>
<td></td>
</tr>
<tr>
<td>perchlorate (ClO₄⁻)</td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Ions That Form Insoluble Compounds</th>
<th>Exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>carbonate (CO₃²⁻)</td>
<td>when combined with Group 1 ions or ammonium (NH₄⁺)</td>
</tr>
<tr>
<td>chromate (CrO₄²⁻)</td>
<td>when combined with Group 1 ions or ammonium (NH₄⁺)</td>
</tr>
<tr>
<td>phosphate (PO₄³⁻)</td>
<td>when combined with Group 1 ions or ammonium (NH₄⁺)</td>
</tr>
<tr>
<td>sulfide (S²⁻)</td>
<td>when combined with Group 1 ions or ammonium (NH₄⁺)</td>
</tr>
<tr>
<td>hydroxide (OH⁻)</td>
<td>when combined with Group 1 ions, Ca²⁺, Ba²⁺, or Sr²⁺</td>
</tr>
</tbody>
</table>

17. According to Table F, which of these salts is least soluble in water?
(1) LiCl  (3) FeCl₂
(2) RbCl  (4) PbCl₂

18. Which of the following compounds is least soluble in water?
(1) copper (II) chloride  (3) iron (III) hydroxide
(2) aluminum acetate  (4) potassium sulfate

19. Identify one ion from Table F that can combine with Pb²⁺(aq) to produce an insoluble compound. [1]

20. According to Reference Table F, calcium hydroxide is soluble in water. Identify another hydroxide compound that contains a Group 2 element and is also soluble in water. [1]

Answer: __________________________

7C-1 Making Molar Solutions (7 Questions)

- Molarity is the molar concentration of a solution, expressed as the number of moles of solute per liter of solution.
- Molarity Formula in Reference tables: molarity (M) = moles of solute / liters of solution
- Directions to make a solution with a specific molarity:
  - Determine the volume of the solution to be produced in liters.
  - Determine the desired molarity.
  - Use the above molarity formula to determine the moles of solute.
  - Use the following formula to change the moles into grams of solute
    - mass in grams = moles × gram-formula mass
  - Pour the solvent into the given solute until you have the proper volume.

21. Molarity is defined as the
(1) moles of solute per kilogram of solvent
(2) moles of solute per liter of solution
(3) mass of a solution
(4) volume of a solvent

22. Which unit can be used to express solution concentration?
(1) J/mol  (3) mol/L
(2) L/mol  (4) mol/s

23. A 3.0 M HCl(aq) solution contains a total of
(1) 3.0 grams of HCl per liter of water
(2) 3.0 grams of HCl per mole of solution
(3) 3.0 moles of HCl per liter of solution
(4) 3.0 moles of HCl per mole of water
Using the Molarity Formula (12 Questions)

- molarity (M) = \( \frac{\text{moles of solute}}{\text{liters of solution}} \)
- This formula can be rearranged to:
  - liters of solution = \( \frac{\text{moles of solute}}{\text{molarity (M)}} \)
  - moles of solute = molarity (M) \times \text{liters of solution}

24. How many milliliters of 12.0 M HCl(aq) must be diluted with water to make exactly 500. mL of 3.00 M hydrochloric acid?
   (1) 100. mL  (3) 200. mL
   (2) 125. mL  (4) 250. mL

25. Which sample of HCl(aq) contains the greatest number of moles of solute particles?
   (1) 1.0 L of 2.0 M HCl(aq)
   (2) 2.0 L of 2.0 M HCl(aq)
   (3) 3.0 L of 0.50 M HCl(aq)
   (4) 4.0 L of 0.50 M HCl(aq)

26. What is the total number of moles of NaCl(s) needed to make 3.0 liters of a 2.0 M NaCl solution?
   (1) 1.0 mol  (3) 6.0 mol
   (2) 0.70 mol  (4) 8.0 mol

27. In the space provided below, show a correct numerical setup for determining how many liters of a 1.2 M solution can be prepared with 0.50 mole of C\(_6\)H\(_{12}\)O\(_6\).

7C-3 Using the Gram Formula mass & the Molarity Formula (5 questions)

- Moles = \( \frac{\text{Given mass}}{\text{Formula mass}} \) & molarity (M) = \( \frac{\text{moles of solute}}{\text{liters of solution}} \)
- These two formulas can be combined to: molarity (M) = \( \frac{\text{Given mass}}{\text{Formula mass} \times \text{liters of solution}} \)

28. How many moles of solute are contained in 200 milliliters of a 1 M solution?
   (1) 1  (3) 0.8
   (2) 0.2  (4) 200

29. What is the total number of grams of NaI(s) needed to make 1.0 liter of a 0.010 M solution?
   (1) 0.015  (3) 1.5
   (2) 0.15  (4) 15

30. What is the molarity of 1.5 liters of an aqueous solution that contains 52 grams of lithium fluoride, LiF, (gram-formula mass = 26 grams/mole)?
   (1) 1.3 M  (3) 3.0 M
   (2) 2.0 M  (4) 0.75 M

31. Show a numerical setup for calculating the mass of the solute used to make the solution. [1]

32. Determine the molarity of the solution. [1]
7C-4 Using the Percent Composition Formula (3 Questions)

- % Composition by mass = \( \frac{\text{mass of part}}{\text{mass of whole}} \times 100 \)
- This formula can be rearranged to: \( \text{mass of part} = \frac{\text{% Composition by mass} \times \text{mass of whole}}{100} \)

33. Solubility data for four different salts in water at 60°C are shown in the table below.

<table>
<thead>
<tr>
<th>Salt</th>
<th>Solubility in Water at 60°C</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>10 grams / 50 grams H₂O</td>
</tr>
<tr>
<td>B</td>
<td>20 grams / 60 grams H₂O</td>
</tr>
<tr>
<td>C</td>
<td>30 grams / 120 grams H₂O</td>
</tr>
<tr>
<td>D</td>
<td>40 grams / 80 grams H₂O</td>
</tr>
</tbody>
</table>

Which salt is most soluble at 60°C?
(1) A    (3) C
(2) B    (4) D

Base your answer to question 34 on the information below.

When a person perspires (sweats), the body loses many sodium ions and potassium ions. The evaporation of sweat cools the skin.

After a strenuous workout, people often quench their thirst with sports drinks that contain NaCl and KCl. A single 250-gram serving of one sports drink contains 0.055 gram of sodium ions.

34. In the space below, show a correct numerical setup for calculating the concentration of sodium ions in this sports drink, expressed as percent by mass. [1]

37C-5 Using the Parts Per Million (ppm) Formula (14 Questions)

- parts per million = \( \frac{\text{mass of solute}}{\text{mass of solution}} \times 1,000,000 \)
- This formula can be rearranged to: \( \text{mass of solute} = \frac{\text{parts per million} \times \text{mass of solution}}{1,000,000} \)

35. If 0.025 gram of Pb(NO₃)₂ is dissolved in 100 grams of H₂O, what is the concentration of the resulting solution, in parts per million?
(1) 2.5 X 10⁻⁴ ppm  (3) 250 ppm
(2) 2.5 ppm       (4) 4.0 X 10³ ppm

36. What is the concentration of a solution, in parts per million, if 0.02 gram of Na₃PO₄ is dissolved in 1000 grams of water?
(1) 20 ppm  (3) 0.2 ppm
(2) 2 ppm   (4) 0.02 ppm

37. An aqueous solution contains 300 parts per million of KOH. Determine the number of grams of KOH present in 1000 grams of this solution. [1]

Answer: _____________________ g
D-1 Colligative Property: Vapor Pressure (19 Questions)

- Some molecules at the surface of a liquid (or solid) are able to escape into the atmosphere as a vapor. How much vapor is able to escape depends on:
  - Temperature
    - The higher the temperature, easier the vapor can escape from the liquid
  - Intermolecular forces - the stronger the intermolecular forces, the harder it is for vapor to escape from a substance
    - Polar molecules are held together by dipole-dipole forces.
    - Some hydrogen compounds, such as water, are held together by hydrogen bonds.
  - Atmospheric pressure
    - The lower the pressure, the easier for molecules to escape from the substance.

40. As the pressure on the surface of a liquid decreases, the temperature at which the liquid will boil
   (1) decreases
   (2) increases
   (3) remains the same

41. The relatively high boiling point of water is due to water having
   (1) hydrogen bonding
   (2) metallic bonding
   (3) nonpolar covalent bonding
   (4) strong ionic bonding

42. As the temperature of a liquid increases, its vapor pressure
   (1) decreases
   (2) increases
   (3) remains the same

43. At standard pressure, CH₄ boils at 112 K and H₂O boils at 373 K. What accounts for the higher boiling point of H₂O at standard pressure?
   (1) covalent bonding
   (2) ionic bonding
   (3) hydrogen bonding
   (4) metallic bonding

D-2 Colligative Property: Using Table H - The Boiling Point (17 Questions)

- As the pressure increases, the boiling point of a liquid increases.
  - When vapor pressure = atmospheric pressure, boiling occurs.
- For Table H:
  - on the curve - the boiling point (gas and/or liquid)
  - above the curve - A gas
  - Below the curve - a liquid

44. A liquid’s boiling point is the temperature at which its vapor pressure is equal to the atmospheric pressure. Using Reference Table H, what is the boiling point of propanone at an atmospheric pressure of 70 kPa? [1]

Answer: __________ °C

45. A liquid boils when the vapor pressure of the liquid equals the atmospheric pressure on the surface of the liquid. Using Reference Table H, determine the boiling point of water when the atmospheric pressure is 90. kPa. [1]

Answer: __________ °C
46. According to Reference Table H, what is the boiling point of ethanoic acid at 80 kPa?
   (1) 28°C (3) 111°C
   (2) 100°C (4) 125°C

47. The vapor pressure of a liquid is 0.92 atm at 60°C. The normal boiling point of the liquid could be
   (1) 35°C (3) 55°C
   (2) 45°C (4) 65°C

48. Which liquid has the highest vapor pressure at 75°C?
   (1) ethanoic acid (3) propanone
   (2) ethanol (4) water

49. Which substance has the lowest vapor pressure at 75°C?
   (1) water (3) propanone
   (2) ethanoic acid (4) ethanol

50. At 65°C, which compound has a vapor pressure of 58 kilopascals?
   (1) ethanoic acid (3) propanone
   (2) ethanol (4) water

Base your answer to question 51 on the information below.
The boiling point of a liquid is the temperature at which the vapor pressure of the liquid is equal to the pressure on the surface of the liquid. The heat of vaporization of ethanol is 838 joules per gram. A sample of ethanol has a mass of 65.0 grams and is boiling at 1.00 atmosphere.

51. Based on Table H, what is the temperature of this sample of ethanol? [1] °C

7D-3 Colligative Property: The Nature of the Solute (18 Questions)
• Adding solute to water will cause an increase in the boiling point and a decrease of the freezing point. The amount of change is solely determined by the number of added particles. The greater the number of particles, the greater the change.
- Different types of substances will add different number of particles.
  - Molecular - glucose (C6H12O6)
    C6H12O6(s) ⇌ C6H12O6(aq)
    1 mol 1 mol Total 1 mol
  - Ionic - Rock salt (NaCl) & Calcium chloride (CaCl2)
    NaCl(s) ⇌ Na+(aq) + Cl−(aq)
    1 mol 1 mol + 1 mol Total 2 mol
    CaCl2(s) ⇌ Ca2+(aq) + 2Cl−(aq)
    1 mol 1 mol + 2 mol Total 3 mol

52. Compared to pure water, an aqueous solution of calcium chloride has a
   (1) higher boiling point and higher freezing point
   (2) higher boiling point and lower freezing point
   (3) lower boiling point and higher freezing point
   (4) lower boiling point and lower freezing point

53. What occurs when NaCl(s) is added to water?
   (1) The boiling point of the solution increases, and the freezing point of the solution decreases.
   (2) The boiling point of the solution increases, and the freezing point of the solution increases.
   (3) The boiling point of the solution decreases, and the freezing point of the solution decreases.
   (4) The boiling point of the solution decreases, and the freezing point of the solution increases.

54. Which aqueous solution of KI freezes at the lowest temperature?
   (1) 1 mol of KI in 500. g of water
   (2) 2 mol of KI in 500. g of water
   (3) 1 mol of KI in 1000. g of water
   (4) 2 mol of KI in 1000. g of water

55. Compared to a 0.1 M aqueous solution of NaCl, a 0.8 M aqueous solution of NaCl has a
   (1) higher boiling point and a higher freezing point
   (2) higher boiling point and a lower freezing point
   (3) lower boiling point and a higher freezing point
   (4) lower boiling point and a lower freezing point

Base your answer to question 56 on the information below.
Ethanol, C2H5OH, is a volatile and flammable liquid with a distinct odor at room temperature. Ethanol is soluble in water. The boiling point of ethanol is 78.2°C at 1 atmosphere. Ethanol can be used as a fuel to produce heat energy, as shown by the balanced equation below.

\[
\text{C}_2\text{H}_5\text{OH} (l) + 3\text{O}_2 (g) \rightarrow 2\text{CO}_2 (g) + 3\text{H}_2\text{O} (l) + 1367 \text{ kJ}
\]

56. At 1 atmosphere, compare the boiling point of pure ethanol to the boiling point of a solution in which a nonvolatile substance is dissolved in ethanol. [1]
8A-1 Collision Theory (2 Questions)

- In order for a chemical reaction to occur, there the particles:
  - must be correctly aligned
  - have sufficient collision (Kinetic) energy.

1. A reaction is most likely to occur when reactant particles collide with
   (1) proper energy, only
   (2) proper orientation, only
   (3) both proper energy and proper orientation
   (4) neither proper energy nor proper orientation

2. A chemical reaction between iron atoms and oxygen molecules can only occur if
   (1) the particles are heated
   (2) the atmospheric pressure decreases
   (3) there is a catalyst present
   (4) there are effective collisions between the particles

8A-2 Collision Theory: Surface Area (6 Questions)

- The greater the surface area, the greater exposed number of particles increasing the chance of collisions and therefore the rate of reaction.

3. At STP, which 4.0-gram zinc sample will react fastest with dilute hydrochloric acid?
   (1) lump
   (2) bar
   (3) powdered
   (4) sheet metal

4. A 5.0-gram sample of zinc and a 50.-milliliter sample of hydrochloric acid are used in a chemical reaction. Which combination of these samples has the fastest reaction rate?
   (1) a zinc strip and 1.0 M HCl(aq)
   (2) a zinc strip and 3.0 M HCl(aq)
   (3) zinc powder and 1.0 M HCl(aq)
   (4) zinc powder and 3.0 M HCl(aq)

5. Given the reaction at 25°C:
   \[ \text{Zn(s)} + 2\text{HCl(aq)} \rightarrow \text{ZnCl}_2(\text{aq}) + \text{H}_2(\text{g}) \]
   The rate of this reaction can be increased by using 5.0 grams of powdered zinc instead of a 5.0-gram strip of zinc because the powdered zinc has
   (1) lower kinetic energy
   (2) lower concentration
   (3) more surface area
   (4) more zinc atoms

8A-3 Collision Theory: Concentration (14 Questions)

- The greater the concentration of one or more reactants increases the rate of reaction by increasing the number of collisions.

6. At 20°C, a 1.2-gram sample of Mg ribbon reacts rapidly with 10.0 milliliters of 1.0 M HCl(aq). Which change in conditions would have caused the reaction to proceed more slowly?
   (1) increasing the initial temperature to 25°C
   (2) decreasing the concentration of HCl(aq) to 0.1 M
   (3) using 1.2 g of powdered Mg
   (4) using 2.4 g of Mg ribbon

7. Each of four test tubes contains a different concentration of HCl(aq) at 25°C. A 1-gram cube of Zn is added to each test tube. In which test tube is the reaction occurring at the fastest rate?
<table>
<thead>
<tr>
<th>HCl(aq)</th>
<th>HCl(aq)</th>
<th>HCl(aq)</th>
<th>HCl(aq)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 M</td>
<td>0.1 M</td>
<td>0.01 M</td>
<td>0.001 M</td>
</tr>
<tr>
<td>10 mL</td>
<td>10 mL</td>
<td>10 mL</td>
<td>10 mL</td>
</tr>
</tbody>
</table>
   (1)    (2)    (3)    (4)    

8A-4 Collision Theory: Pressure (0 Questions)

- Gases only - As pressure increases, the concentration of particles increases, increasing chances of collisions and therefore increasing the rate of reaction.

Not tested yet, but it will be!
**8 A - 5 Collision Theory: Temperature (11 Questions)**

- The greater the temperature, the greater the speed of molecules, increasing the energy and number of collisions, therefore increasing the rate of reaction.

8. Increasing the temperature increases the rate of a reaction by

   (1) lowering the activation energy
   (2) increasing the activation energy
   (3) lowering the frequency of effective collisions between reacting molecules
   (4) increasing the frequency of effective collisions between reacting molecules

9. Explain, in terms of collision theory, why the rate of a chemical reaction increases with an increase in temperature. [1]

**8 A - 6 Collision Theory: Catalyst (13 Questions)**

- A catalysts is a substance that can facilitate (help) a reaction by allowing the reaction to proceed with a lower activation (Collision) energy. Catalyst are unchanged after the reaction, ready to facilitate another reaction

10. Adding a catalyst to a chemical reaction results in

   (1) a decrease in activation energy and a decrease in the reaction rate
   (2) a decrease in activation energy and an increase in the reaction rate
   (3) an increase in activation energy and a decrease in the reaction rate
   (4) an increase in activation energy and an increase in the reaction rate

Base your answer to question 11 on the information below.

   During a bread-making process, glucose is converted to ethanol and carbon dioxide, causing the bread dough to rise. Zymase, an enzyme produced by yeast, is a catalyst needed for this reaction.

11. State the effect of zymase on the activation energy for this reaction. [1]

**8 A - 7 Collision Theory: Nature of Reactants (1 Questions)**

- Ionic compounds react faster than covalent because covalent have more bonds that require greater energy to break during the collision.

12. Based on the nature of the reactants in each of the equations below, which reaction at 25°C will occur at the fastest rate?

   (1) C(s) + O\(_2\)(g) \(\rightarrow\) CO\(_2\)(g)
   (2) NaOH(aq) + HCl(aq) \(\rightarrow\) NaCl(aq) + H\(_2\)O(l)
   (3) CH\(_3\)OH(l) + CH\(_3\)COOH(l) \(\rightarrow\)
       CH\(_3\)COOCH\(_3\)(aq) + H\(_2\)O(l)
   (4) CaCO\(_3\)(s) \(\rightarrow\) CaO(s) + CO\(_2\)(g)

**8 A - 8 Collision Theory: Nonspecific (6 Questions)**

- Any of the above

13. State two methods to increase the rate of a chemical reaction and explain, in terms of particle behavior, how each method increases the reaction rate. [2]

   Method and explanation 1: _____________________________

   Method and explanation 2: _____________________________

   _____________________________

   _____________________________

   _____________________________

   _____________________________
8 B - 1 Potential Energy Diagrams Components (19 Questions)

You must know:
- Activation Energy
- Activated Complex
- Heat of Reaction (∆H)
  - ∆H = PE_{products} - PE_{reactants}
  - Changes in potential energy due to energy released or absorbed during the reaction
- Reactants
- Products
- Time
- Exothermic or Endothermic
- A Catalyzed Diagram

14. The potential energy diagram below represents a reaction.

Which arrow represents the activation energy of the forward reaction?
(1) A (2) B (3) C (4) D

15. Which expression represents the ∆H for a chemical reaction in terms of the potential energy, PE, of its products and reactants?
(1) PE of products + PE of reactants
(2) PE of products − PE of reactants
(3) PE of products × PE of reactants
(4) PE of products ÷ PE of reactants

Based on the potential energy diagram below.

Chemical cold packs are often used to reduce swelling after an athletic injury. The diagram represents the potential energy changes when a cold pack is activated.

16. Which lettered interval on the diagram represents the potential energy of the products? [1]

Answer: __________________

17. Which lettered interval on the diagram represents the heat of reaction? [1]

Answer: __________________

8 B - 2 Potential Energy Diagrams Endothermic (5 Questions)

- Products contain more potential energy than reactants so energy is absorbed.
  - The temperature of the components will decrease as the reaction progresses
  - ∆H is a positive value (+) since there is an increase in potential energy.
18. Given the balanced equation representing a reaction:

\[ N_2(g) + O_2(g) + 182.6 \text{ kJ} \rightarrow 2\text{NO}(g) \]

On the labeled axes below draw a potential energy diagram for this reaction. [1]

19. The potential energy diagram for a chemical reaction is shown below.

Each interval on the axis labeled “Potential Energy (kJ)” represents 40 kilojoules. What is the heat of reaction?

(1) -120 kJ  (3) +40 kJ
(2) -40 kJ  (4) +160 kJ

20. When a spark is applied to a mixture of hydrogen and oxygen, the gases react explosively. Which potential energy diagram best represents the reaction?

If the word "energy" was added to the equation to correctly indicate the energy change in this heat pack reaction, would the word "energy" be placed on the "reactant side" or on the "product side" of the equation? [1]

Ans.________________________

22. Given the balanced equation representing a reaction at 101.3 kPa and 298 K:

\[ N_2(g) + 3H_2(g) \rightarrow 2\text{NH}_3(g) + 91.8 \text{ kJ} \]

Which statement is true about this reaction?

(1) It is exothermic and \( \Delta H \) equals –91.8 kJ.
(2) It is exothermic and \( \Delta H \) equals +91.8 kJ.
(3) It is endothermic and \( \Delta H \) equals –91.8 kJ.
(4) It is endothermic and \( \Delta H \) equals +91.8 kJ.

23. Given the reaction:

\[ S(s) + O_2(g) \rightarrow \text{SO}_2(g) + \text{energy} \]

Which diagram best represents the potential energy changes for this reaction?

The balanced equation: 2 Fe(s) + 3 O\(_2\)(g) \rightarrow Fe\(_2\)O\(_3\)(s)
8B - 4 Potential Energy Diagrams Catalyzed (8 Questions)

- A catalyst speeds up the rate of reaction by providing a different pathway requiring a lower activation energy.
- **VERY IMPORTANT** - $\Delta H$ remains the same whether the reaction is catalyzed or not!

Base your answers to questions 24 through 26 on the potential energy diagram below.

24. What is the heat of reaction for the forward reaction? [1]

_____________________________ kJ

25. What is the activation energy for the forward reaction with the catalyst? [1]

_____________________________ kJ

26. Explain, in terms of the function of a catalyst, why the curves on the potential energy diagram for the catalyzed and uncatalyzed reactions are different. [1]

Base your answer to question 27 on the information and diagram below, which represent the changes in potential energy that occur during the given reaction.

Given the reaction: $A + B \rightarrow C$

27. On the diagram above, draw a dashed line to indicate a potential energy curve for the reaction if a catalyst is added. [1]

8B - 5 Using Table I (12 Questions)

- Table I: The Heats of Reaction at 101.3 kPa (Standard Pressure) and 298 K (Room Temperature 77°F)
  - A Positive (+) $\Delta H$ indicates the products have more potential energy than the reactants and the reaction is endothermic (It absorbed energy!).
  - A negative (-) $\Delta H$ indicates the reactants have more potential energy than the products and the reaction is exothermic (It gave-off energy!).

28. According to Table I, which salt releases energy as it dissolves?

  (1) KNO$_3$
  (2) LiBr
  (3) NH$_4$NO$_3$
  (4) NaCl

29. According to Table I which potential energy diagram best represents the reaction that forms H$_2$O(ℓ) from its elements?

30. Which reaction releases the greatest amount of energy per 2 moles of product?

  (1) 2CO(g) + O$_2$(g) $\rightarrow$ 2CO$_2$(g)
  (2) Al(s) + 3O$_2$(g) $\rightarrow$ 2Al$_2$O$_3$(s)
  (3) 2H$_2$(g) + O$_2$(g) $\rightarrow$ 2H$_2$O(g)
  (4) N$_2$(g) + 3H$_2$(g) $\rightarrow$ 2NH$_3$(g)

31. Which balanced equation represents an endothermic reaction?

  (1) C(s) + O$_2$(g) $\rightarrow$ CO$_2$(g)
  (2) CH$_4$(g) + 2O$_2$(g) $\rightarrow$ CO$_2$(g) + 2H$_2$O(ℓ)
  (3) N$_2$(g) + 3H$_2$(g) $\rightarrow$ 2NH$_3$(g)
  (4) N$_2$(g) + O$_2$(g) $\rightarrow$ 2NO(g)
8 C - 1 Phase Equilibrium (6 Questions)

- Phase equilibrium can occur at the melting point (freezing point) and the vaporization point (condensation point).
- When a substance is in phase equilibrium, it simultaneously exists in two phases which both are changing into the other phase at the same rate.
  - For example, at 0°C, ice and water can coexist in equilibrium. Ice changes into water and water changes into ice at the same rate. There may be different amounts, but since the rate of change is the same!

32. Given the diagram representing a closed system at constant temperature:

![Stoppered Flask Diagram]

Which statement describes this system at equilibrium?
(1) The mass of H₂O(ℓ) equals the mass of H₂O(g).
(2) The volume of H₂O(ℓ) equals the volume of H₂O(g).
(3) The number of moles of H₂O(ℓ) equals the number of moles of H₂O(g).
(4) The rate of evaporation of H₂O(ℓ) equals the rate of condensation of H₂O(g).

33. Given the equation representing a phase change at equilibrium:

\[ \text{H}_2\text{O(s)} \rightleftharpoons \text{H}_2\text{O(ℓ)} \]

Which statement describes this equilibrium?
(1) The H₂O(s) melts faster than the H₂O(ℓ) freezes.
(2) The H₂O(ℓ) freezes faster than the H₂O(s) melts.
(3) The mass of H₂O(s) must equal the mass of H₂O(ℓ).
(4) The mass of H₂O(ℓ) and the mass of H₂O(s) each remain constant.

34. Given the equation: \( \text{H}_2\text{O(s)} \rightleftharpoons \text{H}_2\text{O(ℓ)} \)
At which temperature will equilibrium exist when the atmospheric pressure is 1 atm?
(1) 0 K  (3) 273 K
(2) 100 K  (4) 373 K

8 C - 2 Solution Equilibrium (8 Questions)

- Solution equilibrium occurs when there is undissolved solute in a saturated solution.
- When a substance is in solution equilibrium, at any given time some of the extra solute is being dissolved and entering the solution as some of the dissolved solute is leaving the solution keeping the concentrations of the solution constant.
  - For example, excess sugar at the bottom of a hot cup of tea, and the tea can coexist in equilibrium. Some of the sugar will be dissolved by the hot tea while simultaneously the same amount of sugar will leave the saturated solution (and fall to the bottom of the cup). There may be a different amounts in and out of the solution, but the rate of change is the same!

35. A solution that is at equilibrium must be
(1) concentrated  (3) saturated
(2) dilute  (4) unsaturated

36. Which statement must be true when solution equilibrium occurs?
(1) The solution is at STP.
(2) The solution is supersaturated.
(3) The concentration of the solution remains constant.
(4) The masses of the dissolved solute and the undissolved solute are equal.

37. Some solid KNO₃ remains at the bottom of a stoppered flask containing a saturated KNO₃(aq) solution at 22°C.
Which statement explains why the contents of the flask are at equilibrium?
(1) The rate of dissolving is equal to the rate of crystallization.
(2) The rate of dissolving is greater than the rate of crystallization.
(3) The concentration of the solid is equal to the concentration of the solution.
(4) The concentration of the solid is greater than the concentration of the solution.
8 C - 3 Chemical Equilibrium (17 Questions)

- At some point during a chemical reaction, the concentrations of product and reactants will remain the same (State of Equilibrium) if the products and reactants remain in a closed system and no gas or precipitate forms.

- Note when in equilibrium, the RATE OF CHANGE is the SAME but the amounts may be different!

38. Given the equilibrium reaction at STP:

\[ \text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g}) \]

Which statement correctly describes this system?
(1) The forward and reverse reaction rates are equal.
(2) The forward and reverse reaction rates are both increasing.
(3) The concentrations of N\(_2\)O\(_4\) and NO\(_2\) are equal.
(4) The concentrations of N\(_2\)O\(_4\) and NO\(_2\) are both increasing.

39. Given the reaction system in a closed container at equilibrium and at a temperature of 298 K:

\[ \text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g}) \]

The measurable quantities of the gases at equilibrium must be
(1) decreasing
(2) increasing
(3) equal
(4) constant

40. Which factors must be equal in a reversible chemical reaction at equilibrium?
(1) the activation energies of the forward and reverse reactions
(2) the rates of the forward and reverse reactions
(3) the concentrations of the reactants and products
(4) the potential energies of the reactants and products

41. Which statement correctly describes a chemical reaction at equilibrium?
(1) The concentrations of the products and reactants are equal.
(2) The concentrations of the products and reactants are constant.
(3) The rate of the forward reaction is less than the rate of the reverse reaction.
(4) The rate of the forward reaction is greater than the rate of the reverse reaction.

8 D - 1 Le Châtelier's Principle: Concentration Change (23 Questions)

- If the concentration of one substance is increased, initially, the action that reduces that substance is favored. As the product produced increases, the system will eventually establish a new equilibrium.

42. Given the reaction:

\[ \text{N}_2(\text{g}) + \text{O}_2(\text{g}) + 182.6 \text{ kJ} \rightleftharpoons 2 \text{NO}(\text{g}) \]

Which change would cause an immediate increase in the rate of the forward reaction?
(1) increasing the concentration of NO(\text{g})
(2) increasing the concentration of N\(_2\)(\text{g})
(3) decreasing the reaction temperature
(4) decreasing the reaction pressure

43. Given the equation representing a reaction at equilibrium:

\[ \text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g}) \]

What occurs when the concentration of H\(_2\)(\text{g}) is increased?
(1) The equilibrium shifts to the left, and the concentration of N\(_2\)(\text{g}) decreases.
(2) The equilibrium shifts to the left, and the concentration of N\(_2\)(\text{g}) increases.
(3) The equilibrium shifts to the right, and the concentration of N\(_2\)(\text{g}) decreases.
(4) The equilibrium shifts to the right, and the concentration of N\(_2\)(\text{g}) increases.
44. Given the reaction at equilibrium:

\[ A(g) + B(g) \rightleftharpoons AB(g) + \text{heat} \]

The concentration of \(A(g)\) can be increased by

1. lowering the temperature
2. adding a catalyst
3. increasing the concentration of \(AB(g)\)
4. increasing the concentration of \(B(g)\)

45. Given the equation representing a reaction at equilibrium:

\[ N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) + \text{energy} \]

Which change causes the equilibrium to shift to the right?

1. decreasing the concentration of \(H_2(g)\)
2. decreasing the pressure
3. increasing the concentration of \(N_2(g)\)
4. increasing the temperature

8 D - 2 Le Châtelier's Principle: Temperature Change (8 Questions)

- Heat can be considered as a product or reactant.
- As the temperature is raised, both the forward and backward reactions increase, but not equally. The system will undergo changes to reduce stress.

Base your answer to question 46 on the information below. Given the reaction at equilibrium:

\[ 2NO_2(g) + 7H_2(g) \rightleftharpoons 2NH_3(g) + 4H_2O(g) + 1127 \text{ kJ} \]

46. Explain, in terms of Le Chatelier’s principle, why the concentration of \(NH_3(g)\) decreases when the temperature of the equilibrium system increases. [1]

47. Given the reaction at equilibrium:

\[ C_2(g) + D_2(g) \rightleftharpoons 2CD(g) + \text{energy} \]

Which change will cause the equilibrium to shift?

1. increase in pressure
2. increase in volume
3. addition of heat
4. addition of a catalyst

48. Given the reaction at equilibrium:

\[ N_2(g) + 3 H_2(g) \rightleftharpoons 2NH_3(g) + 92.05 \text{ kJ} \]

\(a\) State the effect on the number of moles of \(N_2(g)\) if the temperature of the system is increased. [1]

49. Given the equilibrium reaction in a closed system:

\[ H_2(g) + I_2(g) + \text{heat} \rightleftharpoons 2HI(g) \]

What will be the result of an increase in temperature?

1. The equilibrium will shift to the left and [\(H_2\)] will increase.
2. The equilibrium will shift to the left and [\(H_2\)] will decrease.
3. The equilibrium will shift to the right and [\(HI\)] will increase.
4. The equilibrium will shift to the right and [\(HI\)] will decrease.

8 D - 3 Le Châtelier's Principle: Pressure Change (3 Questions)

- Only affects gases
- As pressure increases, the reaction shifts towards the side with the fewer number of gas molecules.
- As the pressure decreases, the reaction shifts towards the side with the greater number of gas molecules.
- Has no effect on systems with no gas and when the same number of gas molecules are on both sides.

50. Given the reaction at equilibrium:

\[ N(g) + 3 H_2(g) \rightleftharpoons 2 NH_3(g) + 92.05 \text{ kJ} \]

\(b\) State the effect on the number of moles of \(H_2(g)\) if the pressure on the system is increased. [1]

Base your answer to question 51 on the information below. Given the reaction at equilibrium:

\[ 2NO_2(g) \rightleftharpoons N_2O_4(g) + 55.3 \text{ kJ} \]

51. Explain, in terms of Le Chatelier’s principle, why the equilibrium shifts to the right to relieve the stress when the pressure on the system is increased at constant temperature. [1]
8 D - 4 Le Châtelier's Principle: Catalyzed Change (1 Questions)

- A catalyst will increase the forward and backward reactions but will have no effect on the equilibrium concentrations. (The reaction will reach equilibrium quicker.)

52. Given the reaction at equilibrium:

\[ \text{N}_2(g) + 3 \text{H}_2(g) \rightleftharpoons 2 \text{NH}_3(g) + 92.05 \text{ kJ} \]

**c** State the effect on the number of moles of \( \text{NH}_3(g) \) if a catalyst is introduced into the reaction system. Explain why this occurs. [2]

8E- 1 More Entropy (25 Questions)

- Nature tend to move towards a state of lower energy and greater randomness
  - Exothermic rather than endothermic because less activation energy is required.
  - Randomness - Greater randomness (entropy) occur with more particles and less organization.
  - Trends towards entropy include:
    - Solids to Liquids to gases
    - compounds to elements
    - less particles to more particles

53. Which of these changes produces the greatest increase in entropy?
   (1) \( \text{CaCO}_3(s) \rightarrow \text{CaO}(s) + \text{CO}_2(g) \)
   (2) \( 2\text{Mg}(s) + \text{O}_2(g) \rightarrow 2 \text{MgO}(s) \)
   (3) \( \text{H}_2\text{O}(g) \rightarrow \text{H}_2\text{O}(\ell) \)
   (4) \( \text{CO}_2(g) \rightarrow \text{CO}_2(s) \)

54. At STP, a sample of which element has the highest entropy?
   (1) \( \text{Na}(s) \)
   (2) \( \text{Hg}(\ell) \)
   (3) \( \text{Br}_2(\ell) \)
   (4) \( \text{F}_2(g) \)

55. Which list of the phases of \( \text{H}_2\text{O} \) is arranged in order of increasing entropy?
   (1) ice, steam, and liquid water
   (2) ice, liquid water, and steam
   (3) steam, liquid water, and ice
   (4) steam, ice, and liquid water

8 E - 2 Less Entropy (5 Questions)

- Under certain conditions, substances can move towards greater organization (Decrease in entropy)
  - Opposite of above

59. Which process is accompanied by a decrease in entropy?
   (1) boiling of water
   (2) condensing of water vapor
   (3) subliming of iodine
   (4) melting of ice

60. Which sample has the lowest entropy?
   (1) 1 mole of \( \text{KNO}_3(\ell) \)
   (2) 1 mole of \( \text{KNO}_3(s) \)
   (3) 1 mole of \( \text{H}_2\text{O}(\ell) \)
   (4) 1 mole of \( \text{H}_2\text{O}(g) \)

61. Which phase change represents a decrease in entropy?
   (1) solid to liquid
   (2) gas to liquid
   (3) liquid to gas
   (4) solid to gas

56. As carbon dioxide sublimes, its entropy
   (1) decreases
   (2) increases

57. Which 10-milliliter sample of water has the greatest degree of disorder?
   (1) \( \text{H}_2\text{O}(g) \) at 120°C
   (2) \( \text{H}_2\text{O}(\ell) \) at 80°C
   (3) \( \text{H}_2\text{O}(\ell) \) at 20°C
   (4) \( \text{H}_2\text{O}(s) \) at 0°C

58. Even though the process is endothermic, snow can sublime. Which tendency in nature accounts for this phase change?
   (1) a tendency toward greater entropy
   (2) a tendency toward greater energy
   (3) a tendency toward less entropy
   (4) a tendency toward less energy

62. Explain why the entropy of the system decreases as the reaction proceeds. [1]
9 A - 1 Redox (14 Questions)
• Redox - The transfer of electrons from one atom to another. The electrons lost = the electrons gained.
  - The atom oxidized loses electrons causing a gain in the oxidation number (More Positive)
  - The atom reduced gains electrons causing a reduction in the oxidation number (More Negative)

1. Which particles are gained and lost during a redox reaction?
   (1) electrons   (3) neutrons
   (2) protons    (4) positrons

2. Which change in oxidation number indicates oxidation?
   (1) –1 to +2   (3) +2 to –3
   (2) –1 to –2   (4) +3 to +2

3. In any redox reaction, the substance that undergoes reduction will
   (1) lose electrons and have a decrease in oxidation number
   (2) lose electrons and have an increase in oxidation number
   (3) gain electrons and have a decrease in oxidation number
   (4) gain electrons and have an increase in oxidation number

9A-2 Determining the Oxidation Number (26 Questions)
• Rules for determining oxidation number (Remember: Sum of oxidation numbers = 0 in compounds)
  - 1. Uncombined elements = 0
  - 2. Ions, including polyatomic, = its charge
  - 3. Group 1 = +1 and Group 2 = +2
  - 4. Group 7 = -1 if they are most electronegative (Table S)
  - 5. Hydrogen = +1 except combined with a metal = -1
  - 6. Oxygen = -2 except when combined with fluorine = +2 or in the peroxide ion (O$_2^{2-}$) = -1

Base your answer to question 4 on the following redox reaction, which occurs spontaneously in an electrochemical cell.

Zn + Cr$^{3+}$ $\rightarrow$ Zn$^{2+}$ + Cr

4. Which species loses electrons and which species gains electrons?[1]

____________________ loses electrons.
____________________ gains electrons.

5. What is the oxidation number of chromium in K$_2$C$_2$O$_7$?
   (1) +6
   (2) +2
   (3) +7
   (4) +12

6. Given the balanced equation representing a reaction:
   Fe$_2$O$_3$ + 2Al $\rightarrow$ Al$_2$O$_3$ + 2Fe
   During this reaction, the oxidation number of Fe changes from
   (1) +2 to 0 as electrons are transferred
   (2) +2 to 0 as protons are transferred
   (3) +3 to 0 as electrons are transferred
   (4) +3 to 0 as protons are transferred

7. What is the oxidation state of nitrogen in the compound NH$_4$Br?
   (1) –1
   (2) +2
   (3) –3
   (4) +4

9 B - 1 Recognizing a Redox Reaction (10 Questions)
• Look for Single replacement reactions or equations that have an uncombined element(s) on one side and the same element(s) is(are) part of a compound on the other side.

8. In which reaction are electrons transferred from one reactant to another reactant?
   (1) 2Ca(s) + O$_2$ (g) $\rightarrow$ 2CaO(s)
   (2) AgNO$_3$(aq) + KCl(aq) $\rightarrow$ AgCl(s) + KNO$_3$(aq)
   (3) HCl(aq) + NaOH(aq) $\rightarrow$ NaCl(aq) + H$_2$O($\ell$)
   (4) H$_2$O$^+(aq)$ + OH$^-$ (aq) $\rightarrow$ 2H$_2$O($\ell$)

9. Which balanced equation represents a redox reaction?
   (1) AgNO$_3$ + NaCl $\rightarrow$ AgCl + NaNO$_3$
   (2) BaCl$_2$ + K$_2$CO$_3$ $\rightarrow$ BaCO$_3$ + 2KCl
   (3) CuO + CO $\rightarrow$ Cu + CO$_2$
   (4) HCl + KOH $\rightarrow$ KCl + H$_2$O

10. Which equation represents an oxidation-reduction reaction?
    (1) H$^+$ + OH$^-$ $\rightarrow$ H$_2$O
    (2) $^{235}_{92}$U $\rightarrow$ $^{235}_{67}$Th + $^{2}$He
    (3) Zn + Sn$^{4+}$ $\rightarrow$ Zn$^{2+}$ + Sn$^{2+}$
    (4) 3AgNO$_3$ + Li$_3$PO$_4$ $\rightarrow$ Ag$_3$PO$_4$ + 3LiNO$_3$
9 B - 2 Determining Which Electrons Are Transferred (4 Questions)

- Use the rules for determining the oxidation number for each uncombined element and each compound in the equation to determine the changes produced by the chemical reaction.

  - Mg(s) + HCl(aq) \rightarrow H_2(g) + MgCl_2(aq)

  - Magnesium is oxidized - loses electrons - oxidation number increased from 0 to +2
  - Hydrogen is reduced - gains electrons - oxidation number decreases from +1 to 0

11. Given the redox reaction:

   Cr^{3+} + Al \rightarrow Cr + Al^{3+}

   As the reaction takes place, there is a transfer of
   (1) electrons from Al to Cr^{3+}
   (2) electrons from Cr^{3+} to Al
   (3) protons from Al to Cr^{3+}
   (4) protons from Cr^{3+} to Al

12. Given the balanced ionic equation representing a reaction:

   2Al^{3+}(aq) + 3Mg(s) \rightarrow 3Mg^{2+}(aq) + 2Al(s)

   In this reaction, electrons are transferred from
   (1) Al to Mg^{2+}
   (2) Al^{3+} to Mg
   (3) Mg to Al^{3+}
   (4) Mg^{2+} to Al

9B-3a Recognizing Oxidation in Equations (2 Questions)

- Oxidation occurs when an atom loses electrons, increasing its oxidation number

13. Given the reaction:

   Mg(s) + 2H^{+}(aq) + 2Cl^{-}(aq) \rightarrow Mg^{2+}(aq) + 2Cl^{-}(aq) + H_2(g)

   Which species undergoes oxidation?
   (1) Mg(s)  (3) Cl^{-}(aq)
   (2) H^{+}(aq)  (4) H_2(g)

14. Given the reaction:

   Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)

   Which statement correctly describes what occurs when
   this reaction takes place in a closed system?
   (1) Atoms of Zn(s) lose electrons and are oxidized.
   (2) Atoms of Zn(s) gain electrons and are reduced.
   (3) There is a net loss of mass.
   (4) There is a net gain of mass.

9B-3b Recognizing Oxidation in Equations (2 Questions)

- Reduction occurs when an atom gains electrons, lowering its oxidation number

15. Given the equation:

   C(s) + H_2O(g) \rightarrow CO(g) + H_2(g)

   Which species undergoes reduction?
   (1) C(s)  (3) C^{2+}
   (2) H^{+}  (4) H_2(g)

16. Given the reaction:

   2Al(s) + Fe_2O_3(s) \rightarrow Al_2O_3(s) + 2Fe(s)

   Which species undergoes reduction?
   (1) Al  (3) Al^{3+}
   (2) Fe  (4) Fe^{3+}

9 B - 4 Balancing Redox Equations (4 Questions)

17. Balance the following redox equation below using the smallest whole-number coefficients. [1]

   ____ Zn + ____ Cr^{3+} \rightarrow ____ Zn^{2+} + ____ Cr

18. Balance the redox equation below, using the smallest whole-number coefficients. [1]

   ____Cu(s) + ____AgNO_3(aq) \rightarrow ____Cu(NO_3)_2(aq) + ____Ag(s)
**9 C - 1 Half-Reactions Defined (1 Question)**

- Half reactions show the exchange of electrons of the oxidation or the reduction portion of a Redox reaction.
- Usually only show one type of atom
- Follow the Laws of Conservation of Mass and Charge

19. Half-reactions can be written to represent all
   - (1) double-replacement reactions
   - (2) neutralization reactions
   - (3) fission and fusion reactions
   - (4) oxidation and reduction reactions

**9 C - 2 Oxidation Half-Reaction (8 Questions)**

- When writing oxidation equations, electrons (e⁻) are being removed and therefore they are on the product side.
  - First determine the oxidation numbers for all elements so you can determine which element loses electrons (oxidation number increases) and therefore is being oxidized.

\[
\begin{align*}
\text{Fe}^{2+} (aq) + \text{Zn(s)} &\rightarrow \text{Fe(s)} + \text{Zn}^{2+} (aq) \\
\end{align*}
\]

- Zinc oxidation number increases as it becomes more positive since electrons are lost.
  - \(0 \rightarrow +2\)
- Separate zinc from the rest of the substances and write the half-reaction.
  - \(\text{Zn} \rightarrow \text{Zn}^{2+} + 2e^-\)

Base your answer to question 20 on the following redox reaction, which occurs spontaneously in an electrochemical cell.

\[\text{Zn} + \text{Cr}^{3+} \rightarrow \text{Zn}^{2+} + \text{Cr}\]

20. Write the half-reaction for the oxidation that occurs.

\[\text{(1) Zn(s) + 2e}^- \rightarrow \text{Zn}^{2+} (aq)\]

21. Given the balanced ionic equation:

\[\text{Zn(s) + Cu}^{2+} (aq) \rightarrow \text{Zn}^{2+} (aq) + \text{Cu(s)}\]

Which equation represents the oxidation half-reaction?

\[\begin{align*}
(1) &\text{Zn(s) + 2e}^- \rightarrow \text{Zn}^{2+} (aq) \\
(2) &\text{Zn(s)} \rightarrow \text{Zn}^{2+} (aq) + 2e^- \\
(3) &\text{Cu}^{2+} (aq) \rightarrow \text{Cu(s)} + 2e^- \\
(4) &\text{Cu}^{2+} (aq) + 2e^- \rightarrow \text{Cu(s)} \\
\end{align*}\]

**9 C - 3 Reduction Half-Reaction (14 Questions)**

- When writing reduction equations, electrons (e⁻) are being added and therefore they are on the reactant side.
  - First determine the reduction numbers for all elements so you can determine which element gains electrons (oxidation number decreases) and therefore is being reduced.

\[
\begin{align*}
\text{Fe}^{2+} (aq) + \text{Zn(s)} &\rightarrow \text{Fe(s)} + \text{Zn}^{2+} (aq) \\
\end{align*}
\]

- Iron oxidation number decreases as it becomes more negative since electrons are gained.
  - \(+2 \rightarrow 0\)
- Separate iron from the rest of the substances and write the half-reaction.
  - \(\text{Fe}^{2+} + 2e^- \rightarrow \text{Fe}\)

22. Given the balanced ionic equation representing a reaction:

\[\text{2Al(s) + 3Cu}^{2+} (aq) \rightarrow 2\text{Al}^{3+} (aq) + 3\text{Cu(s)}\]

Which half-reaction represents the reduction that occurs?

\[\begin{align*}
(1) &\text{Al} \rightarrow \text{Al}^{3+} + 3e^- \\
(2) &\text{Al}^{3+} + 3e^- \rightarrow \text{Al} \\
(3) &\text{Cu} \rightarrow \text{Cu}^{2+} + 2e^- \\
(4) &\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu} \\
\end{align*}\]

23. Which half-reaction equation represents the reduction of an iron(II) ion?

\[\begin{align*}
(1) &\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + e^- \\
(2) &\text{Fe}^{2+} + 2e^- \rightarrow \text{Fe} \\
(3) &\text{Fe}^{3+} + e^- \rightarrow \text{Fe}^{2+} \\
(4) &\text{Fe} \rightarrow \text{Fe}^{2+} + 2e^- \\
\end{align*}\]

24. Which half-reaction correctly represents reduction?

\[\begin{align*}
(1) &\text{Mn}^{4+} \rightarrow \text{Mn}^{3+} + e^- \\
(2) &\text{Mn}^{4+} + e^- \rightarrow \text{Mn}^{3+} \\
(3) &\text{Mn}^{4+} + 3e^- \rightarrow \text{Mn}^{7+} \\
(4) &\text{Mn}^{4+} + 3e^- \rightarrow \text{Mn}^{7+} \\
\end{align*}\]
### 9 D - 1 Voltaic Cells - Energy Conversion (8 Questions)

Voltaic cells convert chemical energy to electrical energy.
- Electrons flow spontaneously when there is an electrical connection between the electrodes.

25. In a voltaic cell, chemical energy is converted to
   (1) electrical energy, spontaneously
   (2) electrical energy, nonspontaneously
   (3) nuclear energy, spontaneously
   (4) nuclear energy, nonspontaneously

26. Which energy conversion occurs in a voltaic cell?
   (1) chemical energy to electrical energy
   (2) chemical energy to nuclear energy
   (3) electrical energy to chemical energy
   (4) nuclear energy to electrical energy

### 9 D - 2 Voltaic Cells - Electrodes (8 Questions)

- A voltaic cell contains two containers, each with an electrode
  - Oxidation occurs at the electrode called the anode (An Ox).
    - Remember during oxidation, electrons are lost increasing the oxidation number.
  - Reduction occurs at the electrode called the cathode (Red Cat).
    - Remember, during reduction, electrons are gained, decreasing the oxidation number.

27. Which half-reaction can occur at the anode in a voltaic cell?
   (1) Ni$^{2+} + 2e^- \rightarrow Ni$
   (2) Sn + 2e^- \rightarrow Sn$^{2+}$
   (3) Zn$^- \rightarrow Zn^{2+} + 2e^-$
   (4) Fe$^{3+} \rightarrow Fe^{2+} + e^-$

28. Which statement is true about oxidation and reduction in an electrochemical cell?
   (1) Both occur at the anode.
   (2) Both occur at the cathode.
   (3) Oxidation occurs at the anode and reduction occurs at the cathode.
   (4) Oxidation occurs at the cathode and reduction occurs at the anode.

Base your answer to question 29 on the following redox reaction, which occurs spontaneously in an electrochemical cell.

\[ Zn + Cr^{3+} \rightarrow Zn^{2+} + Cr \]

29. Which half-reaction occurs at the cathode? [1]

Base your answer to question 30 on the diagram of the voltaic cell below.

![Voltaic Cell Diagram]

\[ 2Ag^{+}(aq) + Pb(s) \rightarrow Pb^{2+}(aq) + 2Ag(s) \]

30. When the switch is closed, in which half-cell does oxidation occur? [1]

Answer: __________________________

### 9 D - 3 Using Table J (9 Questions)

- Table J, Activity Series, is very important for quickly determining the anode and cathode of a voltaic cell.
  - Anode - The more reactive metal, as it is more likely it to be oxidized (lose electrons).
  - Cathode - The less reactive metal is the site of reduction (gain electrons).

31. Which ion is most easily reduced?
   (1) Zn$^{2+}$
   (2) Mg$^{2+}$
   (3) Co$^{2+}$
   (4) Ca$^{2+}$

32. Which metal is more active than Ni and less active than Zn?
   (1) Cu
   (2) Cr
   (3) Mg
   (4) Pb

Base your answer to question 33 on the information below.

A flashlight can be powered by a rechargeable nickel-cadmium battery. In the battery, the anode is Cd(s) and the cathode is NiO$_2$(s). The unbalanced equation below represents the reaction that occurs as the battery produces electricity. When a nickel-cadmium battery is recharged, the reverse reaction occurs.

\[ Cd(s) + NiO_2(s) + H_2O(\ell) \rightarrow Cd(OH)_2(s) + Ni(OH)_2(s) \]

33. Explain why Cd would be above Ni if placed on Table J. [1]
34. A diagram of a chemical cell and an equation are shown below.

When the switch is closed, electrons will flow from
(1) the Pb(s) to the Cu(s)
(2) the Cu(s) to the Pb(s)
(3) the Pb\(^{2+}\)(aq) to the Pb(s)
(4) the Cu\(^{2+}\)(aq) to the Cu(s)

35. Given the balanced equation representing the reaction occurring in a voltaic cell:
\[
\text{Zn(s) + Pb}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Pb(s)}
\]
In the completed external circuit, the electrons flow from
(1) Pb(s) to Zn(s)
(2) Pb\(^{2+}\)(aq) to Zn\(^{2+}\)(aq)
(3) Zn(s) to Pb(s)
(4) Zn\(^{2+}\)(aq) to Pb\(^{2+}\)(aq)

36. Identify the particles transferred between Fe\(^{2+}\) and Zn during the reaction in the cell with zinc and iron electrodes. [1]

37. State the purpose of the salt bridge in this voltaic cell. [1]

38. When a voltaic cell operates, ions move through the
(1) anode
(2) cathode
(3) salt bridge
(4) external circuit

39. What is the purpose of the salt bridge in a voltaic cell?
(1) It blocks the flow of electrons.
(2) It blocks the flow of positive and negative ions.
(3) It is a path for the flow of electrons.
(4) It is a path for the flow of positive and negative ions.
9D-6 Voltaic Cells - Chemical Changes (2 Questions)

- At the anode - Electrons are lost causing some of the solid metal electrode to become positive ions that are entering the solution.
- At the cathode - Electrons are gained, causing some positive ions in solution to convert into the solid metal electrode, increasing its mass.

Base your answer to question 40 on the diagram below.

The diagram shows a voltaic cell with copper and aluminum electrodes immediately after the external circuit is completed.

40. As this voltaic cell operates, the mass of the Al(s) electrode decreases. Explain, in terms of particles, why this decrease in mass occurs. [1]

Base your answer to question 41 on the information below.

A voltaic cell with magnesium and copper electrodes is shown in the diagram below. The copper electrode has a mass of 15.0 grams.

41. Explain, in terms of copper ions and copper atoms, why the mass of the copper electrode increases as the cell operates. Your response must include information about both copper ions and copper atoms. [1]

9E-1 Electrolytic Cells - Energy Conversion (4 Questions)

- Electrolytic cells use electrical energy to produce chemical change.
- This process is called electrolysis.
- Electrolysis means to split by electricity.
- A power source such as a batteries must be used.

42. Which statement describes electrolysis?
(1) Chemical energy is used to produce an electrical change.
(2) Chemical energy is used to produce a thermal change.
(3) Electrical energy is used to produce a chemical change.
(4) Thermal energy is used to produce a chemical change.

43. Which energy conversion occurs during the operation of an electrolytic cell?
(1) chemical energy to electrical energy
(2) electrical energy to chemical energy
(3) nuclear energy to electrical energy
(4) electrical energy to nuclear energy
Electrolytic cells - Electrodes (5 Questions)

- An electrolytic cell is contained in one container, with two electrodes.
  - Oxidation occurs at the electrode called the anode (An Ox).
  - Remember during oxidation, electrons are lost increasing the oxidation number.
  - Reduction occurs at the electrode called the cathode (Red Cat).
  - Remember, during reduction, electrons are gained, decreasing the oxidation number.

44. Given the balanced equation representing a reaction occurring in an electrolytic cell:

\[ 2\text{NaCl}(l) \rightarrow 2\text{Na}(l) + \text{Cl}_2(g) \]

Where is Na\((l)\) produced in the cell?
(1) at the anode, where oxidation occurs
(2) at the anode, where reduction occurs
(3) at the cathode, where oxidation occurs
(4) at the cathode, where reduction occurs

45. Reduction occurs at the cathode in
(1) electrolytic cells, only
(2) voltaic cells, only
(3) both electrolytic cells and voltaic cells
(4) neither electrolytic cells nor voltaic cells

Base your answer to question 46 on the diagram and balanced equation below, which represent the electrolysis of molten NaCl.

46. When the switch is closed, which electrode will attract the sodium ions? [1]

The Power Source (6 Questions)

- All electrolytic cells require a power source usually depicted by a battery.
- The electric current forces the non spontaneous reaction to occur.

Base your answer to question 47 on the information below.

The diagram below represents an operating electrolytic cell used to plate silver onto a nickel key. As the cell operates, oxidation occurs at the silver electrode and the mass of the silver electrode decreases.

47. State the purpose of the power source in the cell. [1]

48. Which statement describes one characteristic of an operating electrolytic cell?
(1) It produces electrical energy.
(2) It requires an external energy source.
(3) It uses radioactive nuclides.
(4) It undergoes a spontaneous redox reaction.
### Comparison between Voltaic and Electrolytic cells (4 Questions)

<table>
<thead>
<tr>
<th>Voltaic</th>
<th>Electrolytic</th>
</tr>
</thead>
<tbody>
<tr>
<td>Chemical reactions are spontaneous</td>
<td>Chemical reactions are nonspontaneous</td>
</tr>
<tr>
<td>Reduction occurs at the cathode electrode</td>
<td>Oxidation occurs at the anode electrode</td>
</tr>
<tr>
<td>Each electrode is in their own container.</td>
<td>Both electrodes are in one container.</td>
</tr>
<tr>
<td>Electrons move from the anode (-) to the cathode (+) through a connecting wire.</td>
<td>Electrons are forced from the anode (+) to the cathode (-) by the battery,</td>
</tr>
<tr>
<td>The anode is made of a more reactive metal than the cathode</td>
<td>The anode is made of a less reactive metal than the cathode. When the battery is removed, it could become a voltaic cell, producing a spontaneous current. This is how rechargeable batteries operate.</td>
</tr>
<tr>
<td>Ions move through the salt bridge</td>
<td>Ions can move throughout the electrolyte solution</td>
</tr>
<tr>
<td>This is the basis of the electroplating industry where various thin layers of metal are added to existing metal for decorative and/or protective finishes.</td>
<td></td>
</tr>
</tbody>
</table>

49. Which process requires an external power source?
(1) neutralization  (3) fermentation
(2) synthesis       (4) electrolysis

Base your answer to question 50 on the information and diagram below.

The apparatus shown in the diagram consists of two inert platinum electrodes immersed in water. A small amount of an electrolyte, $\text{H}_2\text{SO}_4$, must be added to the water for the reaction to take place. The electrodes are connected to a source that supplies electricity.

50. What particles are provided by the electrolyte that allow an electric current to flow? [1]

Ans: __________________________

51. What type of electrochemical cell is shown? [1]

Ans: __________________________

52. State one difference between voltaic cells and electrolytic cells. Include information about both types of cells in your answer. [1]

______________________________

______________________________
9E-6 Electrolytic Cells - Chemical Changes (2 Questions)

- At the anode - Electrons are lost causing some of the solid metal electrode to become positive ions that are entering the solution. These positive ions will migrate to and coat the negative cathode.
- At the cathode - Electrons are gained, causing some positive ions in solution to convert to its solid which will coat the metal electrode, increasing its mass.

Base your answer to question 53 on the information below.

The diagram below represents an operating electrolytic cell used to plate silver onto a nickel key. As the cell operates, oxidation occurs at the silver electrode and the mass of the silver electrode decreases.

53. Explain, in terms of Ag atoms and Ag\(^{+}\)(aq) ions, why the mass of the silver electrode decreases as the cell operates. [1]

54. A compound is broken down by chemical means during
(1) chromatography   (3) electrolysis
(2) distillation    (4) filtration
10A-1 Arrhenius Acids and Bases (22 Questions)

- Arrhenius acid is a substance that produces hydrogen ions (H⁺) in an aqueous solution.
  - The hydrogen ion is the only positive ion produced.
  - Example: HF → H⁺ + F⁻ (Hydrofluoric Acid)
    - The hydrogen ion can react with the water to produce the hydronium ion (H₃O⁺)
    - Hydrogen ion and hydronium ion are interchangeable.
  - The strength of the acid depends on the number of hydrogen ions produced.
- Each Arrhenius base produces a hydroxide ion (OH⁻) in an aqueous solution.
  - As with acids, strength depends on the number of hydroxide ions produced.

1. An Arrhenius base yields which ion as the only negative ion in an aqueous solution?
   (1) hydride ion (3) hydronium ion
   (2) hydrogen ion (4) hydroxide ion

2. Which ion is the only negative ion produced by an Arrhenius base in water?
   (1) NO₃⁻ (3) OH⁻
   (2) Cl⁻ (4) H⁺

3. Which formula represents a hydronium ion?
   (1) H₃O⁺ (3) OH⁻
   (2) NH₄⁺ (4) HCO₃⁻

4. An Arrhenius acid has
   (1) only hydroxide ions in solution
   (2) only hydrogen ions in solution
   (3) hydrogen ions as the only positive ions in solution
   (4) hydrogen ions as the only negative ions in solution

5. The only positive ion found in an aqueous solution of sulfuric acid is the
   (1) hydroxide ion (3) sulfite ion
   (2) hydronium ion (4) sulfate ion

10A-2 Acid Examples (9 Questions)

- Acids produce the hydrogen (hydronium) ions (H⁺ of H₃O⁺) when in an aqueous solution.
  - The names of acids can be found in Table K
  - The strength of the acid depends on the number of hydrogen ions produced.
    - Hydrochloric acid (HCl) and Sulfuric Acid (H₂SO₄) - Strong acids as ionization in water approaches 100%
    - Organic acids (-COOH or -C=OH) are weak acids and do not ionized nearly as much.
      - Examples include: acetic acid (vinegar) (CH₃COOH) and citric acid (H₃C₆H₅O₇)

6. Which substance is an Arrhenius acid?
   (1) LiF(aq) (3) Mg(OH)₂(aq)
   (2) HBr(aq) (4) CH₃CHO

7. Which compound is an Arrhenius acid?
   (1) H₂SO₄ (3) NaOH
   (2) KCl (4) NH₃

10A-3 Base Examples (8 Questions)

- Bases produce hydroxide ions (OH⁻) when in an aqueous solution.
  - The names of acids can be found in Table L. Examples include:
    - Sodium Hydroxide (NaOH) (Lye) - Strong Base
    - Ammonia (NH₃) - weaker base (NH₃ + H₂O → NH₄⁺ +OH⁻)
  - Be careful alcohols look like bases, but they do not ionize and are not bases!
    - Examples include: Ethyl alcohol (CH₃CH₂OH) and methyl alcohol (CH₃OH)

8. Which compound is an Arrhenius base?
   (1) CH₃OH (3) LiOH
   (2) CO₂ (4) NO₂

9. Which substance yields hydroxide ion as the only negative ion in aqueous solution?
   (1) Mg(OH)₂ (3) MgCl₂
   (2) C₂H₄(OH)₂ (4) CH₃Cl

10. Which compound releases hydroxide ions in an aqueous solution?
    (1) CH₃COOH (3) HCl
    (2) CH₃OH (4) KOH

11. An aqueous solution of lithium hydroxide contains hydroxide ions as the only negative ion in the solution.
    Lithium hydroxide is classified as an
    (1) aldehyde (3) Arrhenius acid
    (2) alcohol (4) Arrhenius base
10B-1 Electrolytes (21 Questions)

- Electrolytes are substances whose water solutions conduct an electric current due to the presence of ions in solution. The greater the concentration of ions, the stronger the electrolyte and the better it can conduct electricity. Electrolytes include:
  - Acids and bases
  - Organic acids (-COOH) are weak electrolytes
  - Alcohols, such as Ethyl alcohol (CH₂CH₂OH) look like bases, but they do not ionize and are not electrolytes.
  - Salts

12. Which substance, when dissolved in water, forms a solution that conducts an electric current?
   (1) C₂H₅OH  (3) C₁₂H₂₂O₁₁
   (2) C₆H₁₂O₆(s)  (4) CH₃COOH

13. A substance that conducts an electrical current when dissolved in water is called
   (1) a catalyst  (3) a nonelectrolyte
   (2) a metalloid  (4) an electrolyte

14. Which compound is an electrolyte?
   (1) C₆H₁₂O₆  (3) CaCl₂
   (2) CH₃OH  (4) CCl₄

15. Which aqueous solution is the best conductor of an electrical current?
   (1) 0.01 M CH₃OH  (3) 0.1 M CH₃OH
   (2) 0.01 M KOH  (4) 0.1 M KOH

16. A substance is classified as an electrolyte because
   (1) it has a high melting point
   (2) it contains covalent bonds
   (3) its aqueous solution conducts an electric current
   (4) its aqueous solution has a pH value of 7

17. Which compound is an electrolyte?
   (1) butene  (3) dimethyl ether
   (2) propane  (4) methanoic acid

10C-1 Neutralization reactions (22 Questions)

- A Reaction between an acid and a base.
  - Water and a salt are always produced.
  - Example: HCl(aq) + NaOH → H₂O(ℓ) + NaCl(aq)
  - Salts are ionic substances having a metallic or polyatomic positive ion and a negative ion other than hydroxide (OH⁻).
  - In all neutralization reactions there must be a 1:1 ratio of moles of Hydrogen and Hydroxide ions.
  - H⁺ (Acid) + OH⁻ (Base) → H₂O(ℓ)

18. Given the reaction:
   HCl(aq) + LiOH(aq) → HOH(ℓ) + LiCl(aq)
   The reaction is best described as
   (1) neutralization  (3) decomposition
   (2) synthesis  (4) oxidation-reduction

19. Which equation represents a neutralization reaction?
   (1) Na₂CO₃ + CaCl₂ → 2 NaCl + CaCO₃
   (2) Ni(NO₃)₂ + H₂S → NiS + 2 HNO₃
   (3) NaCl + AgNO₃ → AgCl + NaNO₃
   (4) H₂SO₄ + Mg(OH)₂ → MgSO₄ + 2 H₂O

20. Which equation represents a neutralization reaction?
   (1) 4Fe(s) + 3O₂(g) → 2Fe₂O₃(s)
   (2) 2H₂(g) + O₂(g) → 2H₂O(ℓ)
   (3) HNO₃(aq) + KOH(aq) → KNO₃(aq) + H₂O(ℓ)
   (4) AgNO₃(aq) + KCl(aq) → KNO₃(aq) + AgCl(s)

21. Which word equation represents a neutralization reaction?
   (1) base + acid → salt + water
   (2) base + salt → water + acid
   (3) salt + acid → base + water
   (4) salt + water → acid + base

22. Which substance is always a product when an Arrhenius acid in an aqueous solution reacts with an Arrhenius base in an aqueous solution?
   (1) HBr  (3) KBr
   (2) H₂O  (4) KOH

23. Which solution reacts with LiOH(aq) to produce a salt and water?
   (1) KCl(aq)  (3) NaOH(aq)
   (2) Ca(Oaq)  (4) H₂SO₄(aq)

10C-2 Reactions Between Metals and Acids (2 Questions)

- Any metal above hydrogen in Table J, (Activity Series), will react with an acid and produce hydrogen gas (H₂) and a salt.
  - Example: Mg(s) + 2HCl(aq) → MgCl₂(aq) + H₂(g)
Base your answer to question 24 on the information below.

In a laboratory investigation, magnesium reacts with hydrochloric acid to produce hydrogen gas and magnesium chloride. This reaction is represented by the unbalanced equation below.

\[ \text{Mg(s)} + \text{HCl(aq)} \rightarrow \text{H}_2(\text{g}) + \text{MgCl}_2 \text{(aq)} \]

24. State, in terms of the relative activity of elements, why this reaction is spontaneous. [1]

25. Explain, in terms of activity, why HCl(aq) reacts with Zn(s), but HCl(aq) does not react with Cu(s). [1]

26. According to Reference Table J, which of these metals will react most readily with 1.0 M HCl to produce \( \text{H}_2(\text{g}) \)?

(1) Ca  (2) K  (3) Mg  (4) Zn

10D-1 Titration defined (5 Questions)

• Titration is a process in which a known concentration of an acid or base is used in a neutralization reaction to determine the concentration of an unknown base or acid.

27. Which process uses a volume of solution of known concentration to determine the concentration of another solution?

(1) distillation  (2) substitution  (3) transmutation  (4) titration

28. In which laboratory process could a student use 0.10 M NaOH(aq) to determine the concentration of an aqueous solution of HBr?

(1) chromatography  (2) decomposition of the solute  (3) evaporation of the solvent  (4) titration

10D-2 Titration Problems (49 Questions)

• The titration formula is \( M_A \times V_A = M_B \times V_B \) where:

- \( M_A \) = Molarity of H\(^+\) and \( V_A \) = Volume of acid in milliliters
- \( M_B \) = Molarity of OH\(^-\) and \( V_B \) = Volume of base in milliliters

- Molarity = \( \frac{\text{moles of solute}}{\text{liters of solution}} \)

- Very important! The molarity must be expressed in terms of the Hydrogen ion (H\(^+\)) or Hydroxide ion (OH\(^-\)).

Examples include:

- \( 1.0 \text{ M HF} = 1.0 \text{ M H}^+ \)
- \( 1.0 \text{ M H}_2\text{SO}_4 = 2.0 \text{ M H}^+ \)
- \( 1.0 \text{ M H}_3\text{PO}_4 = 3.0 \text{ M H}^+ \)
- \( 1.0 \text{ M KOH} = 1.0 \text{ M OH}^- \)
- \( 1.0 \text{ M Mg(OH)}_2 = 2.0 \text{ M OH}^- \)

Base your answer to question 29 on the information below.

A student titrates 60.0 mL of HNO\(_3\)(aq) with 0.30 M NaOH(aq). Phenolphthalein is used as the indicator. After adding 42.2 mL of NaOH(aq), a color change remains for 25 seconds, and the student stops the titration.

29. In the space below, show a correct numerical setup for calculating the molarity of the HNO\(_3\)(aq). [1]

30. If 5.0 milliliters of a 0.20 M HCl solution is required to neutralize exactly 10. milliliters of NaOH, what is the concentration of the base?

(1) 0.10 M  (2) 0.20 M  (3) 0.30 M  (4) 0.40 M

31. A student neutralized 16.4 milliliters of HCl by adding 12.7 milliliters of 0.620 M KOH. What was the molarity of the HCl acid?

(1) 0.168 M  (2) 0.480 M  (3) 0.620 M  (4) 0.801 M

32. When 50. milliliters of an HNO\(_3\) solution is exactly neutralized by 150 milliliters of a 0.50 M solution of KOH, what is the concentration of HNO\(_3\)?

(1) 1.0 M  (2) 1.5 M  (3) 3.0 M  (4) 0.5 M
In a titration experiment, a student uses a 1.4 M HBr(aq) solution and the indicator phenolphthalein to determine the concentration of a KOH(aq) solution. The data for trial 1 is recorded in the table below.

### Trial 1

<table>
<thead>
<tr>
<th>Buret Readings</th>
<th>HBr(aq)</th>
<th>KOH(aq)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial volume (mL)</td>
<td>7.50</td>
<td>11.00</td>
</tr>
<tr>
<td>Final volume (mL)</td>
<td>22.90</td>
<td>33.10</td>
</tr>
<tr>
<td>Volume used (mL)</td>
<td>15.40</td>
<td>22.10</td>
</tr>
</tbody>
</table>

33. In the space provided below, show a correct numerical setup for calculating the molarity of the KOH(aq) solution for trial 1. [1]

34. What volume of 0.120 M HNO₃(aq) is needed to completely neutralize 150.0 milliliters of 0.100 M NaOH(aq)?
   (1) 62.5 mL  (2) 125 mL  (3) 180. mL  (4) 360. mL

35. A 25.0-milliliter sample of HNO₃(aq) is neutralized by 32.1 milliliters of 0.150 M KOH(aq). What is the molarity of the HNO₃(aq)?
   (1) 0.117 M  (2) 0.150 M  (3) 0.193 M  (4) 0.300 M

36. How many milliliters of 0.100 M NaOH(aq) would be needed to completely neutralize 50.0 milliliters of 0.300 M HCl(aq)?
   (1) 16.7 mL  (2) 50.0 mL  (3) 150. mL  (4) 300. mL

37. A student completes a titration by adding 12.0 milliliters of NaOH(aq) of unknown concentration to 16.0 milliliters of 0.15 M HCl(aq). What is the molar concentration of the NaOH(aq)?
   (1) 0.11 M  (2) 0.20 M  (3) 1.1 M  (4) 5.0 M

38. If a person produces 0.050 mole of excess HCl in the stomach, how many moles of Mg(OH)₂ are needed to neutralize this excess hydrochloric acid? [1]

   ____________ mol

39. Complete the equation below for the titration reaction by writing the formula of each product. [1]

   HCl(aq) + KOH(aq) → ____________ + ____________

40. In the space below, show a correct numerical setup for calculating the molarity of the KOH(aq) solution. [1]

   Base your answers to questions 39 and 40 on the information below.

   In a titration, 15.65 milliliters of a KOH(aq) solution exactly neutralized 10.00 milliliters of a 1.22 M HCl(aq) solution.

41. Identify one additional safety precaution the student should have taken before performing the titration. [1]

   ________________________________

   ________________________________

42. Write a chemical name for the acid used in the titration. [1]

   Answer: ________________
10E-1 pH (36 Questions)

- A scale called the pH scale is used to express the acidity or alkalinity of an acid or base. The scale is a logarithmic scale and each unit indicates a tenfold change in the presence of the hydrogen (hydronium) ion.
  - Acid values are from 0 to 7 while base values are from 7 to 14. (7 is neutral)
  - A substance with a pH of 1.0 contains 10 times more hydrogen ions that a substance with a pH of 2.0 and 100 times more hydrogen ions with a pH of 3.0.
  - Note: The presence of Hydroxide ions in solution are inversely proportioned to the hydrogen ions. As the hydrogen ions decrease, the presence of hydroxide ions increases. A substance with a pH of 10.0 contains ten times more hydroxide ions than a substance with a pH of 9.0. However, it is important to realize that the pH scale indicates the hydrogen ion concentration.

![The pH Scale]

43. Which of the following pH values indicates the highest concentration of hydronium ions in a solution?
   (1) pH = 1  (2) pH = 2  (3) pH = 3  (4) pH = 4

44. Which of these 1 M solutions will have the highest pH?
   (1) NaOH  (2) CH₃OH  (3) HCl  (4) NaCl

45. Given the following solutions:
   Solution A: pH of 10
   Solution B: pH of 7
   Solution C: pH of 5

Which list has the solutions placed in order increasing H⁺ concentration?
   (1) A, B, C  (2) B, A, C  (3) C, A, B  (4) C, B, A

46. The pH of an aqueous solution changes from 4 to 3 when the hydrogen ion concentration in the solution is decreased by a factor of 10. (1) decreased by a factor of $\frac{3}{4}$  (2) decreased by a factor of 10  (3) increased by a factor of $\frac{4}{3}$  (4) increased by a factor of 10

47. Which change in pH represents a hundredfold increase in the concentration of hydronium ions in a solution?
   (1) pH 1 to pH 2  (2) pH 1 to pH 3  (3) pH 2 to pH 1  (4) pH 3 to pH 1

48. When the pH value of a solution is changed from 2 to 1, the concentration of hydronium ions
   (1) decreases by a factor of 2  (2) increases by a factor of 2  (3) decreases by a factor of 10  (4) increases by a factor of 10

10E-2 Acid Base Indicators (39 Questions)

- An indicator is a substance that will change color when there is a change in the pH.
- Each indicator has a range of pH in which the color will change.
  - This is the indicator working range.
  - See table M

Base your answer to question 49 on the information below.

A student titrates 60.0 mL of HNO₃(aq) with 0.30 M NaOH(aq). Phenolphthalein is used as the indicator. After adding 42.2 mL of NaOH(aq), a color change remains for 25 seconds, and the student stops the titration.

49. What color change does phenolphthalein undergo during this titration? [1]

___________________________ to _________________________

50. Which statement correctly describes a solution with a pH of 9?
   (1) It has a higher concentration of H₃O⁺ than OH⁻ and causes litmus to turn blue.
   (2) It has a higher concentration of OH⁻ than H₃O⁺ and causes litmus to turn blue.
   (3) It has a higher concentration of H₃O⁺ than OH⁻ and causes methyl orange to turn yellow.
   (4) It has a higher concentration of OH⁻ than H₃O⁺ and causes methyl orange to turn red.

51. Which solution when mixed with a drop of bromthymol blue will cause the indicator to change from blue to yellow?
   (1) 0.1 M HCl  (2) 0.1 M NH₃  (3) 0.1 M CH₃OH  (4) 0.1 M NaOH
52. Which indicator is yellow in a solution with a pH of 9.8?
   (1) methyl orange  (3) brom cresol green  
   (2) brom thymol blue  (4) thymol blue

53. In which 0.01 M solution is phenolphthalein pink?
   (1) CH₃OH(aq)  (3) CH₃COOH(aq)  
   (2) Ca(OH)₂(aq)  (4) HNO₃(aq)

54. Based on the results of testing colorless solutions
   with indicators, which solution is most acidic?
   (1) a solution in which brom thymol blue is blue
   (2) a solution in which brom cresol green is blue
   (3) a solution in which phenolphthalein is pink
   (4) a solution in which methyl orange is red

10F-1 Bronsted-Lowry Acids and Bases (8 Questions)

- An acid is any substance that donates a hydrogen atom (H⁺) including those not in an aqueous solution.
  - All Arrhenius acids are Bronsted-Lowry acids, but not all Bronsted-Lowry acids are arrhenius acids
- A base is any substance that accepts a hydrogen ion (proton).
  - All Arrhenius bases are Bronsted-Lowry bases, but not all Bronsted-Lowry bases are arrhenius bases.
- Bronsted-Lowry Acids and bases exist as conjugate Acid-Base pairs.
  - HF → H⁺ + F⁻ (H⁺ acts as the acid while F⁻ acts as the base).
- Note: most of these questions include the words "One acid-base theory."

55. One acid-base theory states that an acid is
   (1) an electron donor  (3) an H⁺ donor
   (2) a neutron donor  (4) an OH⁻ donor

56. One acid-base theory defines a base as an
   (1) H⁺ donor  (3) H donor
   (2) H⁺ acceptor  (4) H acceptor

57. According to one acid-base theory, a water molecule
   acts as an acid when the water molecule
   (1) accepts an H⁺  (3) donates an H⁺
   (2) accepts an OH⁻  (4) donates an OH⁻

58. One alternate acid-base theory states that an acid is
   (1) H⁺ donor  (3) OH⁻ donor
   (2) H⁺ acceptor  (4) OH⁻ acceptor

59. Given the balanced equation representing a reaction:
   \[ \text{HSO}_4^-(aq) \rightarrow \text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{O}^+(aq) \text{ SO}_4^{2-}(aq) \]
   According to one acid-base theory, the \( \text{H}_2\text{O}(l) \) molecules act as
   (1) a base because they accept H ions
   (2) a base because they donate H ions
   (3) an acid because they accept H ions
   (4) an acid because they donate H ions
11A-1 Organic Compounds  (12 Questions)

- Organic Chemistry is the study of compounds containing carbon.
- Carbon is unique in that it readily forms chains, rings and network molecules.
- Each carbon atom shares 4 covalent bonds.
- Some organic compounds are polar due to attached functional groups and will dissolve in water.

1. Which element is present in all organic compounds?
   (1) carbon  (2) hydrogen  (3) nitrogen  (4) oxygen

2. Which element has atoms that can form single, double, and triple covalent bonds with other atoms of the same element?
   (1) hydrogen  (2) oxygen  (3) carbon  (4) fluorine

3. Hexane (C₆H₁₄) and water do not form a solution. Which statement explains this phenomenon?
   (1) Hexane is polar and water is nonpolar.
   (2) Hexane is ionic and water is polar.
   (3) Hexane is nonpolar and water is polar.
   (4) Hexane is nonpolar and water is ionic.

11A-2 Hydrocarbons (28 Questions)

- Hydrocarbons only contain carbon and hydrogen atoms.
- Each carbon shares 4 covalent bonds while hydrogen shares only 1 covalent bond (-1 bond - 2 shared electrons).
- Hydrocarbons are classified into 3 different series or major groupings depending on the number of bonds shared by any two of their carbon atoms (See Tables P and Q):
  - Alkanes
    - General formula CₙH₂ₙ₊₂
    - Always saturated - only single bonds between carbons
    - Examples:
      Methane  Ethane  Propane  Butane  Pentane
  - Alkenes
    - General formula CₙH₂ₙ
    - Always unsaturated - must have a double bond
    - Examples:
      Ethene  Propene  2-Butene  1-Pentene
  - Alkynes
    - General formula CₙH₂ₙ₋₂
    - Always unsaturated - must have a triple bond
    - Examples:
      Ethyne  Propyne  1-Butyne  2-Pentyne

4. Given the formula of a substance:
   \[
   \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\
   \text{C} = \text{C} - \text{C} = \text{C} \\
   \text{H} \quad \text{H}
   \]
   What is the total number of shared electrons in a molecule of this substance?
   (1) 22  (2) 11  (3) 9  (4) 6

5. Which formula represents a hydrocarbon?
   (1) CH₃CH₂CH₂CHO  (2) CH₃CH₂CH₂CH₃  (3) CH₃CH₂CH₂COOH  (4) CH₃CH₂COOCH₃

6. Which structural formula is incorrect?
   \[
   \text{H} - \text{C} = \text{C} - \text{Cl} \quad \text{H} - \text{C} = \text{C} - \text{OH} \quad \text{H} = \text{C} = \text{C} - \text{H} \\
   \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H}
   \]
   (1)  (2)  (3)  (4)
7. Which formula represents propyne?
   (1) C_3H_4  
   (2) C_3H_6  
   (3) C_5H_8  
   (4) C_5H_{10}

8. What is the empirical formula for the compound C_5H_{12}O_6?
   (1) CH_2O  
   (2) C_2H_4O_2  
   (3) C_3H_6O_3  
   (4) C_6H_{12}O_6

9. Which general formula represents the homologous series of hydrocarbons that includes the compound 1-heptyne?
   (1) C_{nH_{2n-6}}  
   (2) C_{nH_{2n-2}}  
   (3) C_{nH_{2n}}  
   (4) C_{nH_{2n+2}}

11A-3 Saturated vs Unsaturated (27 Questions)
   - Alkanes are saturated hydrocarbons, having all single bonds between the carbons.
   - Alkenes and alkynes are unsaturated, having one or more double or triple bonds between the carbons.
     - Very important as often they are the starting molecules for more complex molecules.

13. How is the bonding between carbon atoms different in unsaturated hydrocarbons and saturated hydrocarbons? [1]

14. Which structural formula represents an unsaturated hydrocarbon?
   \[ \begin{array}{c}
   \text{(1)} \\
   \text{H} \\
   \text{H} \\
   \text{H} \\
   \text{H} \\
   \text{H} \\
   \text{C} \equiv \text{C} \\
   \text{H} \\
   \text{H} \\
   \text{(3)} \\
   \text{H} \\
   \text{H} \\
   \text{H} \\
   \text{H} \\
   \text{H} \\
   \text{C} \equiv \text{C} \\
   \text{H} \\
   \text{H} \\
   \end{array} \]

15. Which formula represents an unsaturated hydrocarbon?
   (1) C_2H_6  
   (2) C_3H_8  
   (3) C_6H_8  
   (4) C_6H_{14}

16. Which organic compound is a saturated hydrocarbon?
   (1) ethyne  
   (2) ethene  
   (3) ethanol  
   (4) ethane

17. Which formula represents an unsaturated hydrocarbon?
   (1) CH_2CHCl  
   (2) CH_3CH_2Cl  
   (3) CH_3CH_2CH_3  
   (4) CH_3CHCH_2

18. Which compound is a saturated hydrocarbon?
   (1) CH_2CH_2  
   (2) CH_3CH_3  
   (3) CH_3CHO  
   (4) CH_3CH_2OH

19. Which formula represents an unsaturated hydrocarbon?
   (1) C_5H_{12}  
   (2) C_6H_{14}  
   (3) C_7H_{16}  
   (4) C_8H_{14}

11A-4 Drawing Structural Formulas from Condensed Structural Formulas (1 Question)
   - 1st carbon and what is attached
   - 3rd carbon and what is attached
   - 4th carbon and what is attached
   - 2nd carbon and what is attached

Base your answer to question 35 on the condensed structural formula below.

\[ \text{CH}_3\text{CH}_2\text{CHCH}_2 \]

20. In the space provided to the right, draw the structural formula for this compound. [1]
11A-5 Isomers (22 Questions)
• Isomers are substances with the same molecular and empirical formulas but a different structural formula.
• Isomers have different physical and chemical properties.
• The number of isomers increase as the number of carbons increase.
  - 1-3 carbon atoms - none
  - 4 carbon atoms - 2 isomers
  - 8 carbon atoms - 18 isomers
  - 10 carbon atoms - 75 isomers
• Examples:

21. Which pair of compounds are isomers?
(1) NO₂ and N₂O₄
(2) P₂O₅ and P₄O₁₀
(3) HCOOH and CH₃COOH
(4) CH₃OCH₃ and C₂H₅OH

22. The compounds CH₃OCH₃ and CH₃CH₂OH are isomers of each other. These two compounds must have the same
(1) density
(2) reactivity
(3) melting point
(4) molecular formula

Base your answer to question 6 on the information below.
The formula below represents a hydrocarbon.

23. In the space below, draw a structural formula for one isomer of this hydrocarbon. [1]

24. Given a formula representing a compound:
Which formula represents an isomer of this compound?

25. The three isomers of pentane have different
(1) formula masses
(2) molecular formulas
(3) empirical formulas
(4) structural formulas

26. Which two compounds are isomers of each other?
(1) CH₃CH₂COOH and CH₃COOCH₂CH₃
(2) CH₃CH₂CHO and CH₃COCH₃
(3) CH₃CHBrCH₂ and CH₂BrCHBrCH₃
(4) CH₃CHOHCH₃ and CH₃CHOHCH₂OH

27. Which two compounds have the same molecular formula but different chemical and physical properties?
(1) CH₂CH₂Cl and CH₂CH₂Br
(2) CH₂CHCH₂ and CH₂CH₂CH₃
(3) CH₃CHO and CH₃COCH₃
(4) CH₂CH₂OH and CH₃OCH₃

11B-1 Naming Organic Compounds (16 Questions)
• Step 1: Determine the longest continuous carbon chain and if the compound has a branched carbon
  - a. No branched carbons - Normal form - n- before the name.
    - C—C
  - C—C—C—C or C—C
    - Name: Both are n-butane
  - b. Contains branched carbon - longest carbon chain is 3, so it is a propane
    - C—C—C
  - C
    - Contains Methyl group (CH₃) on middle carbon
      - Only place where it could be attached - Name: Methyl propane
  - c. Contains branched carbon - longest carbon chain is 4, so it is a Butane
- Contains methyl group on number 2 carbon - Name: 2-methyl butane
- Number from end that will give you the lowest number!

- d. Contains branched carbon - longest carbon chain is 5, so it is a Pentane
- Contains a methyl group on the 2nd and 3rd carbons so its 2,3 methyl pentane
- 2-di-, 3-tri-, 4-tetra-, etc so (d) whole name: 2,3 dimethyl pentane

- e. Contains branched carbon - longest carbon chain is 4, so it is a Butane
- Contains fluorine atoms on number 2 and number 3 carbons - Name: 2,3 fluorobutane
- 2-di-, 3-tri-, 4-tetra-, etc so (d) whole name: 2,3 difluorobutane

For some questions, you need to draw the molecules to determine if they are isomers to another molecule.
See Table P for the organic prefixes used in naming.

28. Molecules of 2-methyl butane and 2,2-dimethyl propane have different
   (1) structural formulas
   (2) molecular formulas
   (3) numbers of carbon atoms
   (4) numbers of covalent bonds

29. The formula below represents a product formed when HCl reacts with CH3CH2CHCH2.

```
H H H H
H-C-C-C-C-H
H Cl H H
```

What is an IUPAC name for this product? [1]

30. A molecule of butane and a molecule of 2-butene both have the same total number of
   (1) carbon atoms
   (2) hydrogen atoms
   (3) single bonds
   (4) double bonds

31. Given the structural formula:

```
H     C=C     H
      H C     H
      H     H
```

What is the IUPAC name of this compound?
   (1) propane
   (2) propene
   (3) propanone
   (4) propanal

32. Given the formula:

```
H     C=C     H
      H C     H
      H     H
```

What is the IUPAC name of this compound?
   (1) 2-pentene
   (2) 2-pentyne
   (3) 2-butene
   (4) 2-butyne

33. Which compound is an isomer of pentane?
   (1) butane
   (2) propane
   (3) methyl butane
   (4) methyl propane

34. In the space below, draw a correct structural formula for dichloromethane. [1]

35. In the space below, draw a structural formula for 2,2,4-trimethylpentane. [1]
11C-1 Functional Groups (6 Questions)
- Functional groups are atoms or groups of atoms that replace one or more hydrogens that were attached to a carbon.
  - They contribute to the chemical properties of the compound.
  - See Table R

36. Ethanol and dimethyl ether have different chemical and physical properties because they have different
   (1) functional groups
   (2) molecular masses
   (3) numbers of covalent bonds
   (4) percent compositions by mass

37. Functional groups are used to classify
   (1) organic compounds
   (2) inorganic compounds
   (3) heterogeneous mixtures
   (4) homogeneous mixtures

11C-2 Halides (3 Questions)
- The functional group consists of any Halogen (F, Cl Br or I)
  - Compound called an organic halide
  - See 11B-1 to see how to name them

Base your answer to question 38 on the information below.

The hydrocarbon 2-methylpropane reacts with iodine as represented by the balanced equation below. At standard pressure, the boiling point of 2-methylpropane is lower than the boiling point of 2-iodo-2-methylpropane.

38. To which class of organic compounds does this organic product belong? [1]

11C-3 Alcohols (15 Questions)
- The compound contains one or more hydroxyl groups (-OH) looking like \( \text{OH} \) or \( \text{R-OH} \)
  - The -OH group does not form an ion in water like inorganic bases.
  - They are nonelectrolytes
  - They are polar and miscible with water
  - Naming - Drop the "-e" at the end of the corresponding alkane and add "-ol."
    - Example: Ethanol (Ethyl alcohol)

Base your answer to question 39 on the information below. The incomplete equation below represents an esterification reaction. The alcohol reactant is represented by \( \text{X} \).

39. In the space below, draw the structural formula for the alcohol represented by \( \text{X} \). [1]

40. Which structural formula represents an alcohol?
   (1) \( \text{HOH} \) or \( \text{R-OH} \)
   (2) \( \text{R-C-R} \)
   (3) \( \text{H-C-C-O-C-C-H + H-O} \)
   (4) \( \text{H-C-C-O-H + H-O} \)

41. Which compound is an alcohol?
   (1) propanal
   (2) ethyne
   (3) butane
   (4) methanol
11C-4 Ethers (4 Questions)
- Compounds containing one oxygen atom connecting two carbons with single bonds looking like \( \text{R—O—R'} \)
  - They are very volatile
  - Naming - based on the size of the carbon chains
    - Example: diethyl ether

Base your answer to question 42 on the information below. Diethyl ether is widely used as a solvent.

42. In the space below, draw the structural formula for diethyl ether. [1]

43. Given the structural formula:

\[
\begin{align*}
\text{H} & \quad \text{H} & \quad \text{H} \\
\text{H—C—C—O—C—C—H} & \\
\text{H} & \quad \text{H} & \quad \text{H}
\end{align*}
\]

The compound represented by this formula can be classified as an
- (1) organic acid
- (2) ether
- (3) ester
- (4) aldehyde

11C-5 Aldehydes (4 Questions)
- The compound contains a carboxyl groups (=O) on an end carbon atom looking like \( \text{R—C—H} \)
  - Naming - Drop the "-e" at the end of the corresponding alkane and add "-al."
    - Example: methanal (formaldehyde)

44. What is the IUPAC name for the compound that has the condensed structural formula \( \text{CH}_3\text{CH}_2\text{CH}_2\text{CHO} \)?
- (1) butanal
- (2) butanol
- (3) propanal
- (4) propanol

45. In a propanal molecule, an oxygen atom is bonded with a carbon atom. What is the total number of pairs of electrons shared between these atoms?
- (1) 1
- (2) 2
- (3) 3
- (4) 4

11C-6 Ketones (2 Questions)
- The compound contains a carbonyl groups (=O) on an interior carbon atom looking like \( \text{R—C—R} \)
  - Polar molecule that is miscible in water that is often used as a solvent.
  - Naming - Drop the "-e" at the end of the corresponding alkane and add "-one."
    - Example: propanone (Acetone)

46. What is the IUPAC name of the compound with the following structural formula?
- (1) propanone
- (2) propanal
- (3) butanone
- (4) butanal

47. In the space below, draw the structural formula for propanone. [1]

11C-7 Organic Acids (18 Questions)
- Organic acids are compounds contains a carboxyl groups (-COOH) looking like \( \text{R—CO—OH} \)
  - They are weak electrolytes
  - Naming - Drop the "-e" at the end of the corresponding alkane and add "-oic acid."
    - Example: ethanoic acid (Acetic acid - vinegar)
48. In the space below draw a structural formula for ethanoic acid. [1]

49. What is the total number of carbon atoms in a molecule of ethanoic acid?
   (1) 1  (2) 2  (3) 3  (4) 4

50. The functional group —COOH is found in
   (1) esters  (2) aldehydes  (3) alcohols  (4) organic acids

51. Which of these compounds has chemical properties most similar to the chemical properties of ethanoic acid?
   (1) C₃H₇COOH  (2) C₂H₅OH  (3) C₂H₅COOC₂H₅  (4) C₂H₅OC₂H₅

52. Which compound dissolves in water to form an aqueous solution that can conduct an electric current?
   (1) CCl₄  (2) C₂H₅OH  (3) CH₃COOH  (4) CH₄

53. Given the formula for an organic compound:

   H H H
   H C-C-C-C-C-C-O
   H H H

   This compound is classified as an
   (1) aldehyde  (2) amine  (3) ester  (4) organic acid

54. Given a formula of a functional group:

   -C=O

   An organic compound that has this functional group is classified as
   (1) an acid  (2) an aldehyde  (3) an ester  (4) a ketone

11C-9 Esters (10 Questions)

- The compound contains two oxygen atoms attached to an interior carbon, one double bonded, the other single bonded

   - Esters are manufactured by reacting alcohols with organic acids.
   - They have strong fragrant aromas
   - Naming - Drop the "-e" at the end of the corresponding alkane and add "-oate."
      - Example: methyl propanoate

Base your answer to question 55 on the information below.

One type of soap is produced when ethyl stearate and sodium hydroxide react. The soap produced by this reaction is called sodium stearate. The other product of the reaction is ethanol. This reaction is represented by the balanced equation below.

   C₁₇H₃₅-C-O-C₂H₅ + NaOH → C₁₇H₃₅-C-O⁻Na⁺ + C₂H₅OH

   Ethyl stearate   Sodium hydroxide   Sodium stearate   Ethanol

55. To which class of organic compounds does ethyl stearate belong? [1]

Base your answer to question 56 on the information below.

Many artificial flavorings are prepared using the type of organic reaction shown below.

   H-C-C-OH + HO-C-C-C-H → H-C-C-O-C-C-C-H + HOH

   Reactant 1   Reactant 2

56. In the space below, draw a structural formula for the ester that has an odor like pineapple. [1]

Base your answer to question 57 on the information below.

Many artificial flavorings are prepared using the type of organic reaction shown below.

   Answer: ___________________________
11C-9 Amines (1 Questions)

- Compounds containing a nitrogen atom connecting three R groups looking like \( R' \equiv N \equiv R'' \)
  - Two of the R groups could be Hydrogen (NH\(_2\))
  - They are important biological chemicals present in living organisms.
    - They make up the amino portion of an amino acid, the building blocks of Proteins.
  - Naming - Drop the "-e" at the end of the corresponding alkane and add "-amine."
    - Example: 1-propanamine

58. Which class of organic compounds has molecules that contain nitrogen atoms?

1 (alcohol)  3 (ether)  2 (amine)  4 (ketone)

11C-10 Amino Acids (2 Questions)

- Compounds containing both an amine (amine) (NH\(_2\)) and organic acid group (COOH).
  - There are 20 different amino acids present in humans.
  - Naming - It won't be tested

59. The molecule below belongs to which class of compounds?

\[
\begin{align*}
\text{HO} & \quad \text{C} & \quad \text{C} & \quad \text{N} & \quad \text{H} \\
\text{H} & \quad \text{C} & \quad \text{H} & \quad \text{H} & \quad \text{H}
\end{align*}
\]

(1) alcohol  (2) ester  (3) aldehyde  (4) amino acid

60. Given the structural formula:

\[
\begin{align*}
\text{H} & \quad \text{H} & \quad \text{O} \\
\text{H} & \quad \text{C} & \quad \text{C} & \quad \text{N} & \quad \text{H} \\
\text{H} & \quad \text{H} & \quad \text{H}
\end{align*}
\]

This structural formula represents a molecule of

(1) an aldehyde  (2) an ester  (3) a ketone  (4) an amino acid

11C-11 Amides (3 Questions)

- Forms when two amino acids combine at the carbon of the carboxyl group and nitrogen of the amino group they form bond looking like \( R' \equiv C \equiv \text{NH} \).
  - This bond (C-N) is called a peptide bond.
    - Water is given off
    - This reaction is important because it is used to combined amino acid molecules into proteins.

Base your answer to question 61 on the information below.

Glycine, \( \text{NH}_2\text{CH}_2\text{COOH} \), is an organic compound found in proteins. Acetamide, \( \text{CH}_3\text{CONH}_2 \), is an organic compound that is an excellent solvent. Both glycine and acetamide consist of the same four elements, but the compounds have different functional groups.

61. In the space below, draw a structural formula for acetamide. [1]

11D-1 Combustion (3 Questions)

- Hydrogen combines with sufficient oxygen to produce carbon dioxide and water
  - \( 2\text{C}_2\text{H}_6(g) + 7\text{O}_2(g) \rightarrow 4\text{CO}_2(g) + 6\text{H}_2\text{O}(l) \)
- Hydrogen combines with deficient oxygen to produce carbon monoxide and water
  - \( 2\text{C}_2\text{H}_6(g) + 5\text{O}_2(g) \rightarrow 4\text{CO}(g) + 6\text{H}_2\text{O}(l) \)

62. Which form of energy is converted to thermal energy when propane burns in air?

(1) electromagnetic  (3) electrical  (2) nuclear  (4) chemical
11D-2 Substitution (8 Questions)

- One element or group is substituted by another element or group.
  - Example: Ethane combines with chlorine to produce chloroethane and hydrogen chloride
    - \( \text{C}_2\text{H}_6 + \text{Cl}_2 \rightarrow \text{C}_2\text{H}_5\text{Cl} + \text{HCl} \)
  - A double replacement reaction.

63. Given the balanced equation representing a reaction:

\[
\text{CH}_3\text{CH}_2\text{CH}_3 + \text{Br}_2 \rightarrow \text{CH}_3\text{CH}_2\text{CH}_2\text{Br} + \text{HBr}
\]

This organic reaction is best classified as:

1) an addition reaction
2) an esterification reaction
3) a polymerization reaction
4) a substitution reaction

64. Given the equation:

\[
x + \text{Cl}_2 \rightarrow \text{C}_2\text{H}_5\text{Cl} + \text{HCl}
\]

Which molecule is represented by \( x \)?

(1) \( \text{C}_2\text{H}_4 \)  (2) \( \text{C}_2\text{H}_6 \)  (3) \( \text{C}_3\text{H}_6 \)  (4) \( \text{C}_3\text{H}_8 \)

11D-3 Addition (8 Questions)

- One or more elements or groups are added to another substance.
  - Ethene (Double bond) combines with hydrogen to produce a saturated hydrocarbon
    - \( \text{C}_2\text{H}_4 + \text{H}_2 \rightarrow \text{C}_2\text{H}_6 \)

65. Which formula correctly represents the product of an addition reaction between ethene and chlorine?

1) \( \text{CH}_2\text{Cl}_2 \)  (2) \( \text{CH}_3\text{Cl} \)  (3) \( \text{C}_2\text{H}_4\text{Cl}_2 \)  (4) \( \text{C}_2\text{H}_3\text{Cl} \)

66. Which formula represents the product of the addition reaction between ethene and chlorine, \( \text{Cl}_2 \)?

\[
\begin{array}{cccc}
\text{Cl} & \text{Cl} & \text{Cl} & \text{Cl} \\
\text{Cl} & \text{C} & \text{C} & \text{Cl} \\
\text{Cl} & \text{C} & \text{C} & \text{Cl} \\
\text{Cl} & \text{C} & \text{C} & \text{Cl} \\
\end{array}
\]

(1)  (2)  (3)  (4)

67. What type of organic reaction is represented by this equation? [1]

68. Identify the type of organic reaction represented by the equation. [1]

11D-4 Esterification (9 Questions)

- Organic acids and alcohol combines to produce an ester and water.

Base your answers to questions 68 through 70 on the information below.

The equation below represents the reaction between butanoic acid and an unidentified reactant, \( X \).

\[
\text{H} - \text{C} - \text{C} - \text{C} - \text{C} - \text{OH} + X \rightarrow \text{H} - \text{C} - \text{C} - \text{C} - \text{C} - \text{OS}-\text{C} - \text{C} - \text{H} + \text{H-O-H}
\]

69. Write the molecular formula of the organic product in the equation. [1]

70. In the space below, draw a structural formula for the unidentified reactant, \( X \), in the equation. [1]
11D-5 Saponification (3 Questions)

- An ester combines with an inorganic base to produce soap and an alcohol
  - Fat and sodium hydroxide can be combined to produce soap and an alcohol

Base your answer to question 71 on the information below.

One type of soap is produced when ethyl stearate and sodium hydroxide react. The soap produced by this reaction is called sodium stearate. The other product of the reaction is ethanol. This reaction is represented by the balanced equation below.

\[
\begin{align*}
C_{17}H_{35}COOCH_{2}CH_{3} + NaOH & \rightarrow C_{17}H_{35}COO^-Na^+ + CH_{3}CH_{2}OH \\
\text{Ethyl stearate} & \quad \text{Sodium hydroxide} & \quad \text{Sodium stearate} & \quad \text{Ethanol}
\end{align*}
\]

72. Identify the type of organic reaction used to make soap. [1]

73. Which reaction results in the production of soap?
   (1) esterification
   (2) fermentation
   (3) polymerization
   (4) saponification

11D-6 Fermentation (1 Questions)

- Energy can be liberated by breaking sugar bonds by the process of anaerobic respiration of fermentation.
  - Alcohol and carbon dioxide are products
  - \( C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2 \)
  - Occurs in yeast in oxygen poor environments.

74. Given the balanced equation with an unknown compound represented by \( X \):

\[
C_6H_{12}O_6(aq) \xrightarrow{\text{enzyme}} 2X + 2CO_2(g)
\]

Which compound is represented by \( X \)?
   (1) \( CH_3OH(aq) \)
   (2) \( CH_2(OH)_4(aq) \)
   (3) \( CH_3CH_2OH(aq) \)
   (4) \( CH_2OHCH_2OH(aq) \)

11D-6 Polymerization (6 Questions)

- Smaller sub-units are combined producing large polymer molecules.
  - Polymers are organic compounds that are made up of smaller units bonded together.
  - Polymers includes starches, proteins, nylon, rayon and polyethylene

75. The process of joining many small molecules into larger molecules is called
   (1) neutralization
   (2) polymerization
   (3) saponification
   (4) substitution

76. The reaction that joins thousands of small, identical molecules to form one very long molecule is called
   (1) esterification
   (2) fermentation
   (3) polymerization
   (4) substitution

77. Which type of reaction is represented by the equation below?
   Note: \( n \) and \( n \) are very large numbers equal to about 2000.

\[
\begin{align*}
n \left( \begin{array}{c}
H \\
\text{C} \quad \text{C} \\
H \\
\end{array} \right) & \rightarrow \left( \begin{array}{c}
\text{H} \\
\text{C} \\
\text{C} \\
\text{H} \\
\end{array} \right)_n \\
\end{align*}
\]

Which type of reaction is represented by the equation?
   (1) esterification
   (2) fermentation
   (3) saponification
   (4) polymerization
### Table O
Symbols Used in Nuclear Chemistry

<table>
<thead>
<tr>
<th>Name (originates in nucleus)</th>
<th>Notation</th>
<th>Symbol</th>
<th>Mass</th>
<th>Charge</th>
<th>Penetrating power</th>
<th>Notes</th>
</tr>
</thead>
<tbody>
<tr>
<td>alpha particle</td>
<td>²⁰⁶He or ²α</td>
<td>α</td>
<td>4</td>
<td>+2</td>
<td>Low</td>
<td>Helium Nucleus</td>
</tr>
<tr>
<td>beta particle (electron)</td>
<td>e⁻ or β⁻</td>
<td>β⁻</td>
<td>0</td>
<td>-1</td>
<td>Moderate</td>
<td>Originates when a neutron is converted to a proton</td>
</tr>
<tr>
<td>gamma radiation</td>
<td>γ⁻</td>
<td>γ</td>
<td>0</td>
<td>0</td>
<td>Great</td>
<td>High energy electromagnetic radiation</td>
</tr>
<tr>
<td>neutron</td>
<td>n⁻ or n</td>
<td>n</td>
<td>1</td>
<td>0</td>
<td></td>
<td></td>
</tr>
<tr>
<td>proton</td>
<td>¹⁰¹H or ¹p</td>
<td>p</td>
<td>1</td>
<td>+1</td>
<td></td>
<td></td>
</tr>
<tr>
<td>positron</td>
<td>e⁺ or β⁺</td>
<td>β⁺</td>
<td>0</td>
<td>+1</td>
<td></td>
<td>Originates when a proton is converted to a neutron</td>
</tr>
</tbody>
</table>

1. Which nuclear emission has the greatest mass?
   (1) alpha particle
   (2) beta particle
   (3) gamma ray
   (4) positron

2. Compared to the mass and the penetrating power of an alpha particle, a beta particle has
   (1) less mass and greater penetrating power
   (2) less mass and less penetrating power
   (3) more mass and greater penetrating power
   (4) more mass and less penetrating power

3. Which list of nuclear emissions is arranged in order from the least penetrating power to the greatest penetrating power?
   (1) alpha particle, beta particle, gamma ray
   (2) alpha particle, gamma ray, beta particle
   (3) gamma ray, beta particle, alpha particle
   (4) beta particle, alpha particle, gamma ray

4. Alpha particles and beta particles differ in
   (1) mass, only
   (2) charge, only
   (3) both mass and charge
   (4) neither mass nor charge

5. Which statement best describes gamma radiation?
   (1) It has a mass of 1 and a charge of 1.
   (2) It has a mass of 0 and a charge of –1.
   (3) It has a mass of 0 and a charge of 0.
   (4) It has a mass of 4 and a charge of +2.

6. Which nuclear decay emission consists of energy, only?
   (1) alpha particle
   (2) beta particle
   (3) gamma ray
   (4) positron

7. What is the mass number of an alpha particle?
   (1) 1
   (2) 2
   (3) 3
   (4) 4

### Table N
Selected Radioisotopes (Abridged)

<table>
<thead>
<tr>
<th>Nuclide</th>
<th>Half-Life</th>
<th>Decay Mode</th>
<th>Nuclide Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>¹⁹⁸Au</td>
<td>2.69 d</td>
<td>β⁻</td>
<td>gold-198</td>
</tr>
<tr>
<td>¹⁹⁹Au</td>
<td>17.2 s</td>
<td>β⁺</td>
<td>neon-199</td>
</tr>
<tr>
<td>³²³P</td>
<td>14.3 d</td>
<td>β⁻</td>
<td>phosphorus-32</td>
</tr>
<tr>
<td>²³⁵U</td>
<td>7.1 × 10⁸ y</td>
<td>α⁺</td>
<td>uranium-235</td>
</tr>
<tr>
<td>²³⁸U</td>
<td>4.51 × 10⁹ y</td>
<td>α⁺</td>
<td>uranium-238</td>
</tr>
</tbody>
</table>

*Note: Some questions require you to use both Table O and Table N*
8. Positrons are spontaneously emitted from the nuclei of
   (1) potassium-37  (3) nitrogen-16
   (2) radium-226   (4) thorium-232

9. Which notation of a radioisotope is correctly paired with the notation of its emission particle?
   (1) $^{37}$Ca and $^1_2$He
   (2) $^{235}$U and $^3_1$H
   (3) $^{16}$N and $^1_0$p
   (4) $^3_0$H and $^3_0$e

10. Which list of radioisotopes contains an alpha emitter, a beta emitter, and a positron emitter?
    (1) C-14, N-16, P-32
    (2) Cs-137, Fr-220, Tc-99
    (3) Kr-85, Ne-19, Rn-222
    (4) Pu-239, Th-232, U-238

11. Which radioisotope is a beta emitter?
    (1) $^{90}$Sr
    (3) $^{37}$K
    (2) $^{220}$Fr
    (4) $^{238}$U

12. Which isotope will spontaneously decay and emit particles with a charge of +2?
    (1) $^{53}$Fe
    (3) $^{198}$Au
    (2) $^{137}$Cs
    (4) $^{220}$Fr

13. According to Reference Table N, which pair of isotopes spontaneously decays?
    (1) C-12 and N-14
    (2) C-12 and N-16
    (3) C-14 and N-14
    (4) C-14 and N-16

12A-3 Unstable Nuclei (7 Questions)
Radioisotopes have unstable nuclei that spontaneously decay by emitting radiation. Unstable nuclei have a greater ratio of neutrons to protons.

14. The stability of an isotope is based on its
    (1) number of neutrons, only
    (2) number of protons, only
    (3) ratio of neutrons to protons
    (4) ratio of electrons to protons

15. A beta particle may be spontaneously emitted from
    (1) a ground-state electron
    (2) a stable nucleus
    (3) an excited electron
    (4) an unstable nucleus

12B-1 Transmutation (22 Questions)
Transmutation - when one element changes into another element (Atomic number changes).
Natural - One reactant nuclei decays by emitting an alpha particle, beta particle of a positron. (See table N)
- General formula: Reactant $\rightarrow$ New element + Nuclide
- Example: $^{12}$C $\rightarrow$ $^{10}$N + $^0_0$e

Artificial - Two reactants - one is bombarded by another, causing change
- General formula: Reactant 1 + Reactant 2 (Nuclide) $\rightarrow$ One or more new elements + Nuclide(s)
- Example: $^{13}_5$Al + $^4_2$He $\rightarrow$ $^{19}_0$P + $^0_1$n

16. Radioactive cobalt-60 is used in radiation therapy treatment. Cobalt-60 undergoes beta decay. This type of nuclear reaction is called
    (1) natural transmutation
    (2) artificial transmutation
    (3) nuclear fusion
    (4) nuclear fission

17. The change that is undergone by an atom of an element made radioactive by bombardment with high-energy protons is called
    (1) natural transmutation
    (2) artificial transmutation
    (3) natural decay
    (4) radioactive decay

18. Which equation represents a spontaneous nuclear decay?
    (1) C+O $\rightarrow$ O$_2$
    (2) H$_2$CO$_3$ $\rightarrow$ CO$_2$ + H$_2$O
    (3) $^{11}_5$Al + $^2_0$He $\rightarrow$ $^{13}_0$P + $^0_1$n
    (4) $^{90}_3$Sr $\rightarrow$ $^{90}_0$ + $^0_0$Y

19. Given the nuclear reaction:
    $^{60}_2$Co $\rightarrow$ $^0_0$e + $^{60}_2$Ni
    This reaction is an example of
    (1) fission
    (2) fusion
    (3) artificial transmutation
    (4) natural transmutation

20. What is the name of the process in which the nucleus of an atom of one element is changed into the nucleus of an atom of a different element?
    (1) decomposition
    (3) substitution
    (2) transmutation
    (4) reduction

21. Which reaction is an example of natural transmutation?
    (1) $^{239}_9$Pu $\rightarrow$ $^{235}_9$U + $^4_2$He
    (2) $^{17}_7$Al + $^2_0$He $\rightarrow$ $^{19}_0$P + $^0_1$n
    (3) $^{238}_9$U + $^0_1$n $\rightarrow$ $^{239}_9$Pu + $^2_0$He
    (4) $^{238}_9$Pu + $^0_1$n $\rightarrow$ $^{146}_5$Ba + $^{90}_3$Sr + $^{3_1}$N
22. The chart below shows the spontaneous nuclear decay of U-238 to Th-234 to Pa-234 to U-234.

What is the correct order of nuclear decay modes for the change from U-238 to U-234?

(1) β⁻ decay, γ decay, β⁻ decay
(2) β⁻ decay, β⁻ decay, α decay
(3) α decay, α decay, β⁻ decay
(4) α decay, β⁻ decay, β⁻ decay

12B-3 (27 Questions)

The sum of the masses and the sum of the charges (atomic number) must be equal on both sides of the equation.

To solve for a nuclear equation

1. If you have more than 1 unknown, check table N to determine the missing form of radiation.
2. Determine what you have by adding the mass numbers and charges (atomic number) of the reactants and then of the products.
3. The difference is the mass number and the atomic number of the missing substance.

Example 1:

\( ^{226}_{88}\text{Ra} \rightarrow \quad + \quad \)  

Step 1: Check Table N to determine form of radiation.

\( ^{226}_{88}\text{Ra} \rightarrow \overset{4}{2}_{2}\text{He} + \quad \)  

Step 2: Add the masses and charges (Atomic Numbers) and determine the difference

<table>
<thead>
<tr>
<th></th>
<th>Reactants</th>
<th>Products</th>
<th>Difference</th>
</tr>
</thead>
<tbody>
<tr>
<td>Masses</td>
<td>226</td>
<td>4</td>
<td>222</td>
</tr>
<tr>
<td>Atomic numbers or charges</td>
<td>88</td>
<td>2</td>
<td>86</td>
</tr>
</tbody>
</table>

\( ^{226}_{88}\text{Ra} \rightarrow \overset{4}{2}_{2}\text{He} + ^{222}_{86}\text{Rn} \)  

Step 3: use the difference to determine the unknown
Example 2:

$^{235}_{92}U + ^1_n \rightarrow ^{96}_{40}Sr + _____ +2^1_n + \text{energy}$

Step 2: Add the masses and charges (Atomic Numbers) and determine the difference.

<table>
<thead>
<tr>
<th>Reactants</th>
<th>Products</th>
<th>Difference</th>
</tr>
</thead>
<tbody>
<tr>
<td>Masses</td>
<td>236</td>
<td>94</td>
</tr>
<tr>
<td>Atomic numbers or charges</td>
<td>92</td>
<td>36</td>
</tr>
</tbody>
</table>

$^{235}_{92}U + ^1_n \rightarrow ^{96}_{40}Sr + ^{139}_{50}Ba + 2^1_n + \text{energy}$

Step 3: use the difference to determine the unknown.

Example 3:

$^{256}_{99}Es + X \rightarrow ^1_n + ^{268}_{101}Md$

Step 2: Add the masses and charges (Atomic Numbers) and determine the difference.

<table>
<thead>
<tr>
<th>Reactants</th>
<th>Products</th>
<th>Difference</th>
</tr>
</thead>
<tbody>
<tr>
<td>Masses</td>
<td>253</td>
<td>257</td>
</tr>
<tr>
<td>Atomic numbers or charges</td>
<td>99</td>
<td>101</td>
</tr>
</tbody>
</table>

$^{256}_{99}Es + ^{2}{_2}\text{He} \rightarrow ^1_n + ^{268}_{101}Md$

Step 3: use the difference to determine the unknown.

23. Using Reference Table N, complete the below equation for the nuclear decay of $^{226}_{88}$Ra. Include both atomic number and mass number for each particle. [1]

$^{226}_{88}$Ra → _______ + _______

24. In the reaction $^{239}_{93}$Np → $^{239}_{94}$Pu + X, what does X represent?
   (1) a neutron
   (2) a proton
   (3) an alpha particle
   (4) a beta particle

25. Given the fusion reaction:

$^1H + ^1H \rightarrow X + \text{energy}$

Which particle is represented by X?
   (1) $^1H$
   (2) $^1H$
   (3) $^3He$
   (4) $^3He$

26. Given the nuclear equation:

$^{56}_{28}$Cu → $^{58}_{28}$Ni + X

What nuclear particle is represented by X? [1]

27. Which equation represents positron decay?
   (1) $^{86}_{39}$Rb → $^0_1\text{e} + ^{86}_{38}\text{Sr}$
   (2) $^{237}_{92}\text{U} \rightarrow ^{233}_{90}\text{Th} + ^4_{3}\text{He}$
   (3) $^{40}_{17}\text{Al} + ^2_1\text{He} \rightarrow ^{36}_{19}\text{P} + ^1_0\text{n}$
   (4) $^{11}_{5}\text{C} \rightarrow ^0_1\text{e} + ^8_{11}\text{B}$

28. Which equation represents the radioactive decay of $^{226}_{88}$Ra?
   (1) $^{226}_{88}$Ra → $^{222}_{86}$Rn + $^4_{2}\text{He}$
   (2) $^{226}_{88}$Ra → $^{222}_{86}$Ac + $^6_{2}\text{He}$
   (3) $^{226}_{88}$Ra → $^{222}_{86}$Fr + $^6_{2}\text{He}$
   (4) $^{226}_{88}$Ra → $^{222}_{86}$Ra + $^1_0\text{n}$

29. Complete the equation below for the nuclear decay of the radioisotope $^{56}_{28}$Fe used to study red blood cells. $^{56}_{28}$Fe gives off beta particles and has a half life of 44.5 days. Include both the atomic number and the mass number for each missing particle. [1]

$^{56}_{28}$Fe → _______ + _______

30. Based on Reference Table N, complete the decay equation for N-16 below. [1]

$^{16}_{8}\text{N} \rightarrow _______ + _______$

12C-1 (13 Questions)

Nuclear fission and Nuclear fusion reactions are similar in that both reactions convert mass into large amounts of energy ($E=mc^2$). Therefore the products have less mass than the reactants. (There goes the law of conservation of mass!)

31. During a nuclear reaction, mass is converted into
   (1) charge
   (2) energy
   (3) isomers
   (4) volume

32. Which type of reaction releases the greatest amount of energy per mole of reactant?
   (1) combustion
   (2) decomposition
   (3) nuclear fusion
   (4) oxidation-reduction
33. A nuclear fission reaction and a nuclear fusion reaction are similar because both reactions
(1) form heavy nuclides from light nuclides
(2) form light nuclides from heavy nuclides
(3) release a large amount of energy
(4) absorb a large amount of energy

34. The amount of energy released from a fission reaction is much greater than the energy released from a chemical reaction because in a fission reaction
(1) mass is converted into energy
(2) energy is converted into mass
(3) ionic bonds are broken
(4) covalent bonds are broken

12C-2 Pros and Cons of Nuclear Energy (8 Questions)

<table>
<thead>
<tr>
<th>Pros</th>
<th>Cons</th>
</tr>
</thead>
<tbody>
<tr>
<td>• A large amount of energy is produced from a small amount of material</td>
<td>• THE RADIOACTIVITY!!!!!</td>
</tr>
<tr>
<td>• No global warming (greenhouse) gases are produced</td>
<td>• A radioactive leak can render an area uninhabitable.</td>
</tr>
<tr>
<td>• No acid rain gases are produced</td>
<td>• A radioactive leak can cause harm to organisms from direct tissue damage or by increasing cancer rates.</td>
</tr>
<tr>
<td>• Fusion produces less radioactive material than fusion</td>
<td>• It is hard to store fission wastes since it remains highly radioactive for a long period of time.</td>
</tr>
<tr>
<td>(Note - Its still radioactive!)</td>
<td>• Thermal pollution can damage the surrounding aquatic environment</td>
</tr>
<tr>
<td></td>
<td>• Expensive due to litigation.</td>
</tr>
</tbody>
</table>

35. What is one benefit associated with a nuclear fission reaction?
(1) The products are not radioactive.
(2) Stable isotopes are used as reactants.
(3) There is no chance of biological exposure.
(4) A large amount of energy is produced.

36. Which risk is associated with using nuclear fission to produce energy in a power plant?
(1) depletion of hydrocarbons
(2) depletion of atmospheric oxygen
(3) exposure of workers to radiation
(4) exposure of workers to sulfur dioxide

12C-3 Identifying Fusion and Fission Reactions (13 Questions)

<table>
<thead>
<tr>
<th>Fusion</th>
<th>Fission</th>
</tr>
</thead>
<tbody>
<tr>
<td>• Fusion means to &quot;put together&quot; or to combine</td>
<td>• Fission means &quot;to split&quot; or to divide</td>
</tr>
<tr>
<td>• Involves two lighter reactants combining to form a heavier product</td>
<td>• Usually involves a heavy reactant being smashed by a neutron which produces 2 medium size products along with more free neutrons.</td>
</tr>
<tr>
<td>• Hydrogen is involved</td>
<td>• Uranium is often but not always involved</td>
</tr>
<tr>
<td>• Common fusion reaction is $^1_1$H + $^1_1$H $\rightarrow$ $^2_2$He + $^1_0$N</td>
<td>• Common fusion reaction: $^{235}<em>{92}$U + $^1_0$n $\rightarrow$ $^{84}</em>{36}$Kr + $^{139}_{54}$Ba + $2_0$n + energy</td>
</tr>
</tbody>
</table>

37. A nuclear reaction in which two light nuclei combine to form a more massive nucleus is called
(1) addition
(2) fission
(3) fusion
(4) substitution

38. Given the nuclear equation:

$^{235}_{92}$U + $^1_0$n $\rightarrow$ $^{144}_{56}$Ba + $^{91}_{41}$Kr + $3_0$n + energy

a State the type of nuclear reaction represented by the equation. [1]

39. Which equation represents nuclear fusion?
(1) $^{12}_{6}$C $\rightarrow$ $^{12}_{6}$C + $^1_0$e
(2) $^{27}_{13}$Al + $^1_0$He $\rightarrow$ $^{13}_{10}$P + $^1_0$n
(3) $^{235}_{92}$U + $^1_0$n $\rightarrow$ $^{144}_{56}$Ba + $^{91}_{41}$Kr + $3_0$n
(4) $^1_1$H $\rightarrow$ $^2_2$He + $^0_0$n

40. Nuclear fusion differs from nuclear fission because nuclear fusion reactions
(1) form heavier isotopes from lighter isotopes
(2) form lighter isotopes from heavier isotopes
(3) convert mass to energy
(4) convert energy to mass
12D-1 Half-life & Table N (9 Questions)

- Half-life is the time necessary for 1/2 of the radioisotope to transmute to a different element. The shorter the half-life the more unstable the isotope and the quicker it will change into a different element. A radioisotope with a long half life remains for a long time and could affect people thousands of years in the future.
- The half-life must be considered when choosing a radioisotope for a practical application. If it is going to be used inside a body, you want a short half-life to limit the amount of radiation exposure to the person. If the radioisotope is used in a tool, a longer half-life will allow the tool to operate over a greater period of time.

41. Suppose a 40-gram sample of iodine-131 and a 40-gram sample of cesium-137 were both abandoned in the clinic in 1987. Explain why the sample of iodine-131 would not pose as great a radiation risk to people today as the sample of cesium-137 would. [1]

42. What is the half-life and decay mode of Rn-222?
   (1) 1.91 days and alpha decay
   (2) 1.91 days and beta decay
   (3) 3.82 days and alpha decay
   (4) 3.82 days and beta decay

43. Which nuclide is listed with its half-life and decay mode?
   (1) K-37, 1.24 h, α
   (2) N-16, 7.2 s, β⁻
   (3) Rn-222, 1.6 x 10³ y, α
   (4) U-235, 7.1 x 10⁸ y, β⁻

Base your answer to question 44 on the information below.

The radioisotopes carbon-14 and nitrogen-16 are present in a living organism. Carbon-14 is commonly used to date a once-living organism.

44. Explain why N-16 is a poor choice for radioactive dating of a bone. [1]

12D-2 Half-Life Problems(25 Questions)

- There is no magic formula for solving half-life problems. You must rely on logic. The below table was constructed using the definition of half-life. You should write your own as soon as you get an exam.

<table>
<thead>
<tr>
<th>Number of half-lives</th>
<th>0</th>
<th>1</th>
<th>2</th>
<th>3</th>
<th>4</th>
<th>5</th>
<th>6</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fraction of Radioisotope Remaining</td>
<td>1</td>
<td>1/2</td>
<td>1/4</td>
<td>1/8</td>
<td>1/16</td>
<td>1/32</td>
<td>1/64</td>
</tr>
</tbody>
</table>

- Example 1: How many days are required for 160. grams of radon-222 to decay to 40.0 grams?
  - Step 1: Determine how many half-lives have passed: 1/4 of the original amount remains, therefore two half-lives has passed.
  - Step 2: Find the half-life of radon in Table N: ²²²Rn = 3.82 d
  - Step 3: Determine how many days does 2 half lives of Radon-222 equal.: 7.64 days

- Example 2: What is the half-life of sodium-25 if .500 gram of a 8.00-gram sample of sodium-25 remains unchanged after 237 seconds?
  - Step 1: Determine what fraction of the original amount remain: \( \frac{500}{8000} = \frac{1}{16} \)
  - Step 2: Find the number of half-lives based on the remaining fraction: 4
  - Step 3: Divide the total time by the number of half lives to determine the time of each half-life: 237 seconds ÷ 4 = 59.3 seconds

- Example 3: Polonium-218 has a half life of 3.04 minutes. Determine the original mass of a sample of Po-218, if 0.50 milligram of the sample remains unchanged after 12.16 minutes. [1]
  - Step 1: Determine how many half-lives have passed: 12.16 min ÷ 3.04 min = 4 Half-lives.
  - Step 2: Determine what fraction of the original amount should remain: 4 Half-lives \( \frac{1}{16} \)
  - Step 3: Determine the mass of the original sample by multiplying the remaining sample by the reciprocal of the fraction remaining: 0.50 mg x 16 = 8.0 mg.
45. Based on Reference Table N, what fraction of a radioactive $^{90}$Sr sample would remain unchanged after 56.2 years?

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>(1)</td>
<td>$\frac{1}{2}$</td>
</tr>
<tr>
<td>(2)</td>
<td>$\frac{1}{4}$</td>
</tr>
<tr>
<td>(3)</td>
<td>$\frac{1}{8}$</td>
</tr>
<tr>
<td>(4)</td>
<td>$\frac{1}{16}$</td>
</tr>
</tbody>
</table>

46. If $\frac{1}{8}$ of an original sample of krypton-74 remains unchanged after 34.5 minutes, what is the half-life of krypton-74?

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>(1)</td>
<td>11.5 min</td>
</tr>
<tr>
<td>(2)</td>
<td>23.0 min</td>
</tr>
<tr>
<td>(3)</td>
<td>34.5 min</td>
</tr>
<tr>
<td>(4)</td>
<td>46.0 min</td>
</tr>
</tbody>
</table>

47. What is the half-life of sodium-25 if 1.00 gram of a 16.00-gram sample of sodium-25 remains unchanged after 237 seconds?

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>(1)</td>
<td>47.4 s</td>
</tr>
<tr>
<td>(2)</td>
<td>59.3 s</td>
</tr>
<tr>
<td>(3)</td>
<td>79.0 s</td>
</tr>
<tr>
<td>(4)</td>
<td>118 s</td>
</tr>
</tbody>
</table>

48. Based on Reference Table N, what is the fraction of a sample of potassium-42 that will remain unchanged after 62.0 hours? [1]

Ans: ____________________ mol

49. An original sample of the radioisotope fluorine-21 had a mass of 80.0 milligrams. Only 20.0 milligrams of this original sample remain unchanged after 8.32 seconds. What is the half-life of fluorine-21?

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>(1)</td>
<td>1.04 s</td>
</tr>
<tr>
<td>(2)</td>
<td>2.08 s</td>
</tr>
<tr>
<td>(3)</td>
<td>4.16 s</td>
</tr>
<tr>
<td>(4)</td>
<td>8.32 s</td>
</tr>
</tbody>
</table>

50. Determine the total time that must elapse until only 1/4 of an original sample of the radioisotope Rn-222 remains unchanged. [1]

_____________________________

51. Which statement explains why, nuclear waste materials may pose a problem?

(1) They frequency have short half-lives and remain radioactive for brief periods of time.
(2) They frequently have short half-lives and remain radioactive for extended periods of time.
(3) They frequently have long half-lives and remain radioactive for brief periods of time.
(4) They frequently have long half-lives and remain radioactive for extended periods of time.

52. If animals feed on plants that have taken up Sr-90, the Sr-90 can find its way into their bone structure. Explain one danger to the animals. [1]

53. Which nuclide is used to investigate human thyroid gland disorders?

(1) carbon-14 (3) cobalt-60
(2) potassium-37 (4) iodine-131

54. Which nuclides are used to date the remains of a once-living organism?

(1) C-14 and C-12 (3) I-131 and Xe-131
(2) Co-60 and Co-59 (4) U-238 and Pb-206

55. Which isotope is used to treat cancer?

(1) C-14 (3) Co-60
(2) U-238 (4) Pb-206

12E-1 & 2 Pros, Cons and uses of Radioactivity (26 Questions)

<table>
<thead>
<tr>
<th>Pros</th>
<th>Cons</th>
</tr>
</thead>
<tbody>
<tr>
<td>Can be used in practical applications</td>
<td></td>
</tr>
<tr>
<td>- Iodine-131 - Treat and diagnose thyroid disorders</td>
<td></td>
</tr>
<tr>
<td>- Carbon-14 - Dating remains of a once living organism</td>
<td></td>
</tr>
<tr>
<td>- Cobalt-60 - Treatment of cancer</td>
<td></td>
</tr>
<tr>
<td>- Cobalt-60 - Food irradiation (Killing disease causing bacteria, extending the shelf life of the food)</td>
<td></td>
</tr>
<tr>
<td>- Americium-241 - Used in smoke detectors</td>
<td></td>
</tr>
<tr>
<td>- Technetium-99 - bone scanning - for diagnosing bone diseases</td>
<td></td>
</tr>
<tr>
<td>Radioactivity can render an area uninhabitable.</td>
<td></td>
</tr>
<tr>
<td>- If radioactive particles becomes waterborne or airborne, the radioactive particles can spread over a large area and move far from their source.</td>
<td></td>
</tr>
<tr>
<td>- Radioactivity can cause harm to organisms by directly damaging tissues or by increasing mutation and cancer rates.</td>
<td></td>
</tr>
<tr>
<td>- This is especially true for radioactive isotopes that can replace nonradioactive elements in the organism</td>
<td></td>
</tr>
<tr>
<td>- Gamma radiation is most dangerous due to their great penetrating power.</td>
<td></td>
</tr>
</tbody>
</table>

51. Which statement explains why, nuclear waste materials may pose a problem?

(1) They frequency have short half-lives and remain radioactive for brief periods of time.
(2) They frequently have short half-lives and remain radioactive for extended periods of time.
(3) They frequently have long half-lives and remain radioactive for brief periods of time.
(4) They frequently have long half-lives and remain radioactive for extended periods of time.

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(1) carbon-14 (3) cobalt-60
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