1. Answer the following problems about gases.
   (a) The average atomic mass of naturally occurring neon is 20.18 amu. There are two common isotopes of naturally occurring neon as indicated in the table below.

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Mass (amu)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ne-20</td>
<td>19.99</td>
</tr>
<tr>
<td>Ne-22</td>
<td>21.99</td>
</tr>
</tbody>
</table>

(i) Using the information above, calculate the percent abundance of each isotope.

\[
\frac{(19.99x) + (21.99(100-x))}{100} = 20.18 \Rightarrow \frac{-2x = -18.1}{-2} \Rightarrow x = 90.5\% \text{ Ne} \quad 9.5\% \text{ } ^{22}\text{Ne}
\]

(ii) Calculate the number of Ne-22 atoms in a 12.55 g sample of naturally occurring neon.

\[\text{b/c Ne-22 = 9.5}\% \quad (0.95)(12.55g\text{ Ne}) = 1.193(\frac{1\text{ mol}}{20.18g})(\frac{6.02 \times 10^{23}}{1\text{ mol}}) = 3.55 \times 10^{22} \text{ Ne-22}\]

(b) A major line in the emission spectrum of neon corresponds to a frequency of \(4.34 \times 10^{14}\) s\(^{-1}\). Calculate the wavelength, in nanometers, of light that corresponds to this line.

\[\lambda = \frac{3.00 \times 10^{8} \text{ m/s}}{4.34 \times 10^{14} \text{ s}^{-1}} = 6.91 \times 10^{-7} \text{ m} = 691 \text{ nm}\]

(c) In the upper atmosphere, ozone molecules decompose as they absorb ultraviolet (UV) radiation, as shown by the equation below. Ozone serves to block harmful ultraviolet radiation that comes from the Sun.

\[\text{O}_3(g) \rightarrow \text{O}_2(g) + \text{O}(g)\]

A molecule of \(\text{O}_3(g)\) absorbs a photon with a frequency of \(1.00 \times 10^{15}\) s\(^{-1}\).

(i) How much energy, in joules, does the \(\text{O}_3(g)\) molecule absorb per photon?

\[E = hf = (6.626 \times 10^{-34} \text{ J s})(1 \times 10^{15} \text{ Hz}) = 6.63 \times 10^{-19} \text{ J}\]

(ii) The minimum energy needed to break an oxygen-oxygen bond in ozone is \(387 \text{ kJ mol}^{-1}\). Does a photon with a frequency of \(1.00 \times 10^{15}\) s\(^{-1}\) have enough energy to break this bond? Support your answer with a calculation.

\[\frac{6.63 \times 10^{-19} \text{J}}{\text{photon}} \left(\frac{1 \text{ kJ}}{1000 \text{ J}}\right) \left(\frac{6.02 \times 10^{23} \text{ photons}}{1 \text{ mol}}\right) = 399 \text{ kJ mol}^{-1}\]

\text{Yes}
<table>
<thead>
<tr>
<th>Element</th>
<th>First Ionization Energy (kJmol⁻¹)</th>
<th>Second Ionization Energy (kJmol⁻¹)</th>
<th>Third Ionization Energy (kJmol⁻¹)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Element 1</td>
<td>1,251</td>
<td>2,300</td>
<td>3,820</td>
</tr>
<tr>
<td>Element 2</td>
<td>496</td>
<td>4,560</td>
<td>6,910</td>
</tr>
<tr>
<td>Element 3</td>
<td>738</td>
<td>1,450</td>
<td>7,730</td>
</tr>
<tr>
<td>Element 4</td>
<td>1,000</td>
<td>2,250</td>
<td>3,360</td>
</tr>
</tbody>
</table>

2. The table above shows the first three ionization energies for atoms of four elements from the third period of the periodic table. The elements are numbered randomly. Use the information in the table to answer the following questions.

(a) Which element is most metallic in character? Explain your reasoning.

Element 2...loses two...so Element 2 has the lowest 1st IE, therefore it must be a metal.

(b) Identify element 3. Explain your reasoning.

Mg Compared to the other IEs, the 1st + 2nd IE for Element 3 is low, then jumps at 3rd IE (2 valence e⁻)

(c) Write the complete electron configuration for an atom of element 3.

1s² 2s² 2p⁶ 3s²

(d) What is the expected oxidation state for the most common ion of element 2?

(+1) 1 valence e⁻

(e) What is the chemical symbol for element 2?

Na

(f) A neutral atom of which of the four elements has the smallest radius?

Element #1

3. Use the principles of atomic structure and/or bonding to explain each of the following. In each part, your answer must include references to both substances.

a. The atomic radius of Li is larger than that of Be.

Li and Be experience the same shielding, however Be has more nuclear charge causing the radius to become smaller.

b. The second ionization energy of K is greater than the second ionization energy of Ca.

K⁺ 2nd e⁻ is in energy level 3 with less shielding, whereas Ca²⁺ 2nd e⁻ is in energy level 4 with more shielding.

c. The carbon-carbon bond energy in C₂H₄ is greater than it is in C₂H₆.

C₂H₄ contains a double bond whereas C₂H₆ contains all single bonds.

d. The boiling point of Cl₂ is lower than the boiling point of Br₂.

Cl₂ + Br₂ both have LDF₃, however Br₂ has greater LDF due to more e⁻.
4. The structures of a water molecule and a crystal of LiCl(s) are represented above. A student prepares a 1.0 M solution by dissolving 4.2 g of LiCl(s) in enough water to make 100 mL of solution.

   a. In the space provided below, show the interactions of the components of LiCl(aq) by making a drawing that represents the different particles present in the solution. Base the particles in your drawing on the particles shown in the representations above. Include only one formula unit of LiCl and no more than ten molecules of water.

   Your drawing must include the following details.
   I. identity of ions (symbol and charge)
   II. the arrangement and proper orientation of the particles in the solution

\[ \text{LiCl (aq)} \]

5. Answer the following questions about acetylsalicylic acid, the active ingredient in aspirin.

   a. The elements contained in acetylsalicylic acid are hydrogen, carbon, and oxygen. The combustion of 3.000 g of the pure compound yields 1.200 g of water and 3.72 L of dry carbon dioxide, measured at 750. mmHg and 25°C. Calculate the mass, in g, of each element in the 3.000 g sample.

   \[ \begin{align*}
   \text{mass of H}_2\text{O} &= 1.200 \text{ g} \\
   \text{mass of CO}_2 &= 3.72 \text{ L} \\
   \text{mass of C} &= 3.000 - (1.200 + 3.72) = 1.08 \text{ g} \\
   \text{mass of O} &= 2(1.200) + 3(3.72) - 1.08 = 12.938 \text{ g} \\
   \text{mass of H} &= 1.08 \text{ g} \\
   \end{align*} \]

   \[ \text{mass of H} + \text{mass of O} = 13.026 \text{ g} \]

   \[ \text{mass of C} + \text{mass of H} + \text{mass of O} = 13.026 \text{ g} \]

   A student dissolved 1.625 g of pure acetylsalicylic acid in distilled water and titrated the resulting solution to the equivalence point using 88.43 mL of 0.102 M NaOH (aq). Assuming that acetylsalicylic acid has only one ionizable hydrogen, calculate the molar mass of the acid.

   \[ \text{mol} = \text{L} \times \text{M} \]

   \[ \text{mol} = \text{L} \times \text{M} = 0.08843 \text{ L NaOH} \times 0.102 \text{ M NaOH} = 0.009020 \text{ mol OH}^- \times \frac{1.4}{10} = 0.009020 \text{ mol acid} \]

   Molar Mass = \frac{\text{g}}{\text{mol}} \quad \Rightarrow \quad 1.625 \text{ g Acid} = \frac{180.99 \text{ gmol}^{-1}}{0.009020 \text{ mol acid}}
6. Answer the following questions about BeC₂O₄ and its hydrate.

a. Calculate the mass percent of carbon in the hydrated form of the solid that has the formula BeC₂O₄ • 3H₂O

\[ \% C = \frac{24.02 \text{ g C}}{157.09 \text{ g BeC}_2\text{O}_4 \cdot 3\text{H}_2\text{O}} \times 100 = 15.90\% \]

b. When heated to 220 °C, BeC₂O₄ • 3H₂O dehydrates completely as represented below:

\[
\text{BeC}_2\text{O}_4 \cdot 3\text{H}_2\text{O} (s) \rightarrow \text{BeC}_2\text{O}_4 (s) + 3\text{H}_2\text{O} (g)
\]

If 3.21 g of BeC₂O₄ • 3H₂O is heated to 220 °C, calculate:

I. The mass of BeC₂O₄ (s) formed.

\[
3.21 \text{ g BeC}_2\text{O}_4 \cdot 3\text{H}_2\text{O} \left(\frac{97.03 \text{ g BeC}_2\text{O}_4}{157.09 \text{ g BeC}_2\text{O}_4 \cdot 3\text{H}_2\text{O}}\right) \approx 2.06 \text{ g BeC}_2\text{O}_4
\]

II. The volume of the H₂O (g) released measured at 220 °C and 735 mmHg.

\[
\frac{3.21 \text{ g H}_2\text{O}}{11.15 \text{ g H}_2\text{O} \text{ g moles}^{-1}} \cdot \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} = 0.0638 \text{ mol H}_2\text{O}
\]

\[
V = \frac{nRT}{P} = \frac{0.0638 \text{ mol H}_2\text{O} \cdot (22.4 \text{ L mol}^{-1} \cdot 493 \text{ K})}{735 \text{ mmHg}} = 2.67 \text{ L}
\]

c. A 0.345 g sample of anhydrous BeC₂O₄ (s), which contains an inert impurity, was dissolved in sufficient water to produce 100 mL of solution. A 20.0 mL portion of the solution was titrated with KMnO₄ (aq). The balanced equation for the reaction that occurred is as follows:

\[
16 \text{H}^+ (aq) + 2 \text{MnO}_4^- (aq) + 5 \text{C}_2\text{O}_4^{2-} (aq) \rightarrow 2 \text{Mn}^{2+} (aq) + 10 \text{CO}_2 (g) + 8 \text{H}_2\text{O} (l)
\]

The volume of 0.0150 M KMnO₄ (aq) required to reach the equivalence point was 17.80 mL.

I. Identify the substance being oxidized during the titration reaction.

\[ \text{O}_2
\]

II. For the titration at the equivalence point, calculate the number of moles of each of the following that reacted.

1. \( \text{MnO}_4^- \)

\[
0.0150 \text{ M MnO}_4^- \cdot 0.01780 \text{ L} = 0.0267 \text{ mol MnO}_4^-
\]

2. \( \text{C}_2\text{O}_4^{2-} \)

\[
0.0267 \text{ mol MnO}_4^- \cdot \frac{5 \text{ mol C}_2\text{O}_4^{2-}}{2 \text{ mol MnO}_4^-} = 0.068 \times 10^{-4} \text{ mol C}_2\text{O}_4^{2-}
\]

III. Calculate the total number of moles of \( \text{C}_2\text{O}_4^{2-} \) (aq) that were present in the 100 mL of prepared solution.

\[
6.68 \times 10^{-4} \text{ mol C}_2\text{O}_4^{2-} \text{ in 20 mL, so 100 mL would have } 5 \times \text{ as much: }
\]

\[
6.68 \times 10^{-4} \text{ mol C}_2\text{O}_4^{2-} \times 5 = 0.00334 \text{ mL}
\]

IV. Calculate the mass percent of BeC₂O₄ (s) in the impure 0.345 g sample.

\[
0.00334 \text{ mol C}_2\text{O}_4^{2-} \left(\frac{1 \text{ mol BeC}_2\text{O}_4}{1 \text{ mol C}_2\text{O}_4^{2-}} \cdot \frac{97.03 \text{ g BeC}_2\text{O}_4}{1 \text{ mol BeC}_2\text{O}_4}\right) = 0.324 \text{ g BeC}_2\text{O}_4
\]

\[
0.345 \text{ g Impure BeC}_2\text{O}_4 \times 100 = 93.9\% \text{ BeC}_2\text{O}_4
\]
7. Answer the following questions about the element selenium, Se (atomic number 34).

a. Samples of natural selenium contain six stable isotopes. In terms of atomic structure, explain what these isotopes have in common, and how they differ.

b. Write the complete electron configuration (e.g. 1s² 2s² ... etc.) for a selenium atom in the ground state. Indicate the number of unpaired electrons in the ground state atom, and explain your reasoning.

\[
\begin{align*}
1s^2 & \quad 2s^2 \quad 2p^6 \quad 3s^2 \quad 3p^6 \quad 4s^2 \quad 3d^{10} \quad 4p^4 \quad 4f^1 \quad 5s^1 \\
\text{would have 2 unpaired e}^- \quad \text{(Hund's rule)}
\end{align*}
\]

c. In terms of atomic structure, explain why the first ionization energy of selenium is

I. Less than that of bromine (atomic number 35), and

\[
\text{Se} \text{ is less than Br due to the fact that Se has less nuclear charge compared to Br, thus causing Se electrons to be less attracted to the nucleus.}
\]

II. Greater than that of tellurium (atomic number 52).

\[
\text{Se's 1st IE is greater than Te because Se is experiencing less shielding than Te, as a result Se's e}^- \text{are more attracted to the nucleus compared to Te.}
\]

d. Selenium reacts with fluorine to form SeF₄. Draw the complete Lewis electron-dot structure for SeF₄ and sketch the molecular structure. Indicate whether the molecule is polar or nonpolar, and justify your answer.

\[
\text{SeF}_4 \quad \text{Se sees no shape} \quad \text{Polar due to Asymmetry}
\]

8. Answer the following questions about Photoelectron Spectroscopy.

a. Identify the element in the photoelectron spectrum shown below. Briefly explain your reasoning.

[Diagram showing the photoelectron spectrum with peaks at 126, 907, 531, and 0.74, labeled with Rs, 2p, 3s, and Mg, respectively, and a note about the energy levels and ionization energy.]
b. Identify if either of the following statements is correct. Briefly explain your reasoning:

I. The photoelectron spectrum of Mg$^{2+}$ is expected to be identical to the photoelectron spectrum of Ne.

- *No, they would be different.* Mg$^{2+}$ has 12 protons, Ne only 10. Therefore the energy levels would not match.

II. The photoelectron spectrum of $^{35}$Cl is identical to the photoelectron spectrum of $^{37}$Cl.

- *Yes, they would be the same.* The amount of protons in each isotope remains the same, therefore the spectrum would match.

C. Is it possible to deduce the electron configuration for an atom from its photoelectron spectrum? If so, explain how. If not, explain why not.

- *Yes.* Looking at the relative height of the peaks (ie e$^{-}$) and looking at the amount of energy needed to remove electrons (ie energy levels).

9. Naturally occurring chlorine molecules, Cl$_2$, have masses of 70, 72 and 74 amu as seen in the mass spectrum above. They occur in the percentages 56.25%, 37.50% and 6.250% respectively. Use this data to calculate the average atomic mass of chlorine atoms and to find the relative abundance of $^{35}$Cl and $^{37}$Cl isotopes.

\[
\text{Average Atomic Mass} = \frac{(56.25 \times 70) + (37.50 \times 72) + (6.25 \times 74)}{100} \approx 35.48 \text{ amu}
\]

- $75\%$ $^{35}\text{Cl}$
- $25\%$ $^{37}\text{Cl}$

10. A typical mass spectrum for Mg contains three peaks at m/z values of 24, 25 and 26 respectively.

a. What does the existence of three peaks suggest?

- *3 different isotopes: $^{24}\text{Mg}$, $^{25}\text{Mg}$, $^{26}\text{Mg}$*

b. The relative intensities of the three peaks (24, 25 and 26) are found to be 63, 8.1 and 9.1 respectively.

I. What do these data tell us about the isotope with m/z = 26?

- *More abundant than isotope 25, but less abundant than isotope 24."

II. Calculate the relative atomic mass of magnesium.

\[
\frac{(24 \times 63) + (25 \times 8.1) + (26 \times 9.1)}{80.2} \approx 24.32 \text{ amu}
\]
AP Chemistry MC Review Questions: Semester I

1. **E** How many grams of calcium nitrate, Ca(NO₃)₂, contains 24 grams of oxygen atoms?
   (A) 164 grams  (B) 96 grams  (C) 62 grams  (D) 50. grams  (E) 41 grams

2. **D** The simplest formula for an oxide of nitrogen that is 36.8 percent nitrogen by weight is
   (A) N₂O     (B) NO     (C) NO₂   (D) N₂O₃   (E) N₂O₅

   \[ CH₃CH₂COOH(l) + 1 \text{ O}_2(g) \rightarrow 6 \text{ CO}_2(g) + 6 \text{ H}_2\text{O}(l) \]

3. **D** How many moles of O₂ are required to oxidize 1 mole of CH₃CH₂COOH according to the reaction represented above?
   (A) 2 moles  (B) 5/2 moles  (C) 3 moles  (D) 7/2 moles  (E) 9/2 moles

4. **B** When a hydrate of Na₂CO₃ is heated until all the water is removed, it loses 54.3 percent of its mass. The formula of the hydrate is
   (A) Na₂CO₃ · 10 H₂O  (B) Na₂CO₃ · 7 H₂O  (C) Na₂CO₃ · 5 H₂O  (D) Na₂CO₃ · 3 H₂O  (E) Na₂CO₃ · H₂O

5. **B** In which of the following compounds is the mass ratio of chromium to oxygen closest to 1.6 to 1.0?
   (A) CrO₃     (B) Cr₂O₃ (C) CrO    (D) Cr₃O₆  (E) Cr₂O₃

6. **C** The data below were gathered in order to determine the density of an unknown solid.
   
   Mass of an empty container = 3.0 grams
   Mass of the container plus the solid sample = 25.0 grams
   Volume of the solid sample = 11.0 cubic centimeters

   The density of the sample should be reported as
   (A) 0.5 g/cm³ (B) 0.50 g/cm³  (C) 2.0 g/cm³  (D) 2.00 g/cm³  (E) 2.27 g/cm³

**Question 7 refers to the following elements.**
(A) Lithium  (B) Nickel  (C) Bromine  (D) Uranium  (E) Fluorine

7. **E** Is a gas in its standard state at 298 K.

8. **D** The safest and most effective emergency procedure to treat an acid splash on skin is to do which of the following immediately?
   (A) Dry the affected area with paper towels.
   (B) Sprinkle the affected area with powdered Na₂SO₄(s).
   (C) Flush the affected area with water and then with a dilute NaOH solution.
   (D) Flush the affected area with water and then with a dilute NaHCO₃ solution.

9. **B** The melting point of MgO is higher than that of NaF. Explanations for this observation include which of the following?
   I. Mg²⁺ is more positively charged than Na⁺.
   II. O²⁻ is more negatively charged than F⁻.
   III. The O²⁻ ion is smaller than the F⁻ ion.
   (A) II only  (B) I and II only  (C) I and III only  (D) II and III only  (E) I, II, and III
10. **C** When hafnium metal is heated in an atmosphere of chlorine gas, the product of the reaction is found to contain 62.2 percent Hf by mass and 37.4 percent Cl by mass. What is the empirical formula for this compound?
(A) HfCl   (B) HfCl₂  (C) HfCl₃  (D) HfCl₄  (E) Hf₂Cl₃

11. **B** After completing an experiment to determine gravimetrically the percentage of water in a hydrate, a student reported a value of 38%. The correct value for the percentage of water in the hydrate is 51%.

Which of the following is the most likely explanation for this difference?
(A) Strong initial heating caused some of the hydrate sample to spatter out of the crucible.
(B) The dehydrated sample absorbed moisture after heating.
(C) The amount of the hydrate sample used was too small.
(D) The crucible was not heated to constant mass before use.
(E) Excess heating caused the dehydrated sample to decompose.

12. **C** A compound contains 1.10 mol of K, 0.55 mol of Te, and 1.65 mol of O. What is the simplest formula for this compound?
(A) KTeO   (B) KTe₂O   (C) K₂TeO₃   (D) K₂TeO₆   (E) K₄TeO₆

13. **E** Of the following compounds, which is the most ionic?
(A) SiCl₄   (B) BrCl   (C) PCl₅   (D) Cl₂O   (E) CaCl₂

14. **D** The atomic mass of copper is 63.55. Given that there are only two naturally occurring isotopes of copper, ⁶⁵Cu and ⁶⁷Cu, the natural abundance of the ⁶⁵Cu isotope must be approximately
(A) 90%   (B) 70%   (C) 50%   (D) 25%   (E) 10%

15. **B** All of the halogens in their elemental form at 25°C and 1 atm are
(A) conductors of electricity   (B) diatomic molecules   (C) odorless   (D) colorless   (E) gases

16. **B** Which of the following describes the changes in forces of attraction that occur as H₂O changes phase from a liquid to a vapor?
(A) H – O bonds break as H – H and O – O bonds form.
(B) Hydrogen bonds between H₂O molecules are broken.
(C) Covalent bonds between H₂O molecules are broken.
(D) Ionic bonds between H⁺ ions and OH⁻ ions are broken.
(E) Covalent bonds between H⁺ ions and H₂O molecules become more effective.

17. **A** Which of the following conclusions can be drawn from J. J. Thomson’s cathode ray experiments?
(A) Atoms contain electrons.
(B) Practically all the mass of an atom is contained in its nucleus.
(C) Atoms contain protons, neutrons, and electrons.
(D) Atoms have a positively charged nucleus surrounded by an electron cloud.
(E) No two electrons in one atom can have the same four quantum numbers.

**Questions 18-21**

(A) Heisenberg uncertainty principle
(B) Pauli exclusion principle
(C) Hund’s rule (principle of maximum multiplicity)
(D) Shielding effect
(E) Wave nature of matter

18. **C** Can be used to predict that a gaseous carbon atom in its ground state is paramagnetic
19. **E** Explains the experimental phenomenon of electron diffraction
20. **B** Indicates that an atomic orbital can hold no more than two electrons
21. **A** Predicts that it is impossible to determine simultaneously the exact position and the exact velocity of an electron
22. Which of the following is a correct interpretation of the results of Rutherford's experiments in which gold atoms were bombarded with alpha particles?
   (A) Atoms have equal numbers of positive and negative charges.
   (B) Electrons in atoms are arranged in shells.
   (C) Neutrons are at the center of an atom.
   (D) Neutrons and protons in atoms have nearly equal mass.
   (E) The positive charge of an atom is concentrated in a small region.

23. The emission spectrum of hydrogen consists of several series of sharp emission lines in the ultraviolet (Lyman series) in the visible (Balmer series) and in the infrared (Paschen series, Brackett series, etc.) regions of the spectrum.
   (A) What feature of the electronic energies of the hydrogen atom explains why the emission spectrum consists of discrete wavelength rather than a continuum wavelength?
   (B) Account for the existence of several series of lines in the spectrum. What quantity distinguishes one series of lines from another?
   (C) Draw an electronic energy level diagram for the hydrogen atom and indicate on it the transition corresponding to the line of lowest frequency in the Balmer series.
   (D) What is the difference between an emission spectrum and an absorption spectrum? Explain why the absorption spectrum of atomic hydrogen at room temperature has only the lines of the Lyman series.

24. Atoms of an element, X, have the electronic configuration shown above. The compound most likely formed with magnesium, Mg, is:
   (A) MgX (B) Mg2X (C) MgX2 (D) MgX3 (E) Mg3X2

25. The elements in which of the following have most nearly the same atomic radius?
   (A) Be, B, C, N
   (B) Ne, Ar, Kr, Xe
   (C) Mg, Ca, Sr, Ba
   (D) C, P, Se, I
   (E) Cr, Mn, Fe, Co

26. Which of the following represents the ground state electron configuration for the Mn3+ ion? (Atomic number Mn = 25)
   (A) 1s2 2s2 2p6 3s2 3p6 3d4
   (B) 1s2 2s2 2p6 3s2 3p6 3d5 4s2
   (C) 1s2 2s2 2p6 3s2 3p6 3d5 4s2
   (D) 1s2 2s2 2p6 3s2 3p6 3d4 4s2
   (E) 1s2 2s2 2p6 3s2 3p6 3d3 4s1

27. Gaseous atoms of which of the elements above are paramagnetic?
   (A) Ca and As only
   (B) Zn and As only
   (C) Ca, V, and Co only
   (D) V, Co, and As only
   (E) V, Co, and Zn only

Use these answers for questions 28-30.
   (A) O (B) La (C) Rb (D) Mg (E) N

28. What is the most electronegative element of the above?
29. Which element exhibits the greatest number of different oxidation states?
30. Which of the elements above has the smallest ionic radius for its most commonly found ion?

Use these answers for questions 31-34.
   (A) 1s2 2s2 2p6 3s2 3p3
   (B) 1s2 2s2 2p6 3s2 3p2
   (C) 1s2 2s2 2p6 2d10 3s2 3p6
   (D) 1s2 2s2 2p6 3s2 3p6 3d5
(E) 1s 2s 2p 3s 3p 3d 4s
31. An impossible electronic configuration
32. The ground-state configuration for the atoms of a transition element
33. The ground-state configuration of a negative ion of a halogen
34. The ground-state configuration of a common ion of an alkaline earth element

Questions 35-38 refer to atoms for which the occupied atomic orbital's are shown below.
35. Represents an atom that is chemically unreactive
36. Represents an atom in an excited state
37. Represents an atom that has four valence electrons
38. Represents an atom of a transition metal

(A) 1s 2s
(B) 1s 2s
(C) 1s 2s 2p
(D) 1s 2s 2p
(E) [Ar] 4s

<p>| Ionization Energies for element X (kJ/mol) |
|-------------------------------|-------------------------------|-----------------|----------------|----------------|</p>
<table>
<thead>
<tr>
<th>First</th>
<th>Second</th>
<th>Third</th>
<th>Fourth</th>
<th>Fifth</th>
</tr>
</thead>
<tbody>
<tr>
<td>580</td>
<td>1815</td>
<td>2740</td>
<td>11600</td>
<td>14800</td>
</tr>
</tbody>
</table>

39. The ionization energies for element X are listed in the table above. On the basis of the data, element X is most likely to be
(A) Na  (B) Mg  (C) Al  (D) Si  (E) P

40. In the periodic table, as the atomic number increases from 11 to 17, what happens to the atomic radius?
(A) It remains constant.  (B) It increases only.  (C) It increases, and then decreases.
(D) It decreases only.  (E) It decreases, then increases.

Consider atoms of the following elements. Assume that the atoms are in the ground state.
(A) S  (B) Ca  (C) Ga  (D) Sb  (E) Br

41. The atom that contains exactly two unpaired electrons
42. The atom that contains only one electron in the highest occupied energy sublevel

43. In which of the following groups are the three species isoelectronic; i.e., have the same number of electrons?
(A) S²⁺, K⁺, Ca²⁺  (B) Sc, Ti, V²⁺  (C) O²⁻, S²⁻, Cl⁻  (D) Mg²⁺, Ca²⁺, Sr²⁺
(E) Cs, Ba²⁺, La³⁺

44. Which of the following properties generally decreases across the periodic table from sodium to chlorine?
(A) First ionization energy  (B) Atomic mass  (C) Electronegativity
(D) Maximum value of oxidation number  (E) Atomic radius
45. **B** The effective nuclear charge experienced by the outermost electron of Na is different than the effective nuclear charge experienced by the outermost electron of Ne. This difference best accounts for which of the following?
(A) Na has a greater density at standard conditions than Ne.
(B) Na has a lower first ionization energy than Ne.
(C) Na has a higher melting point than Ne.
(D) Na has a higher neutron-to-proton ratio than Ne.
(E) Na has fewer naturally occurring isotopes than Ne.

46. **B** When 70. mL of 3.0 M Na₂CO₃ is added to 30. mL of 1.0 M NaHCO₃ the resulting concentration of Na⁺ is
(A) 2.0 M  (B) 2.4 M  (C) 4.0 M  (D) 4.5 M  (E) 7.0 M

47. **B** A 20.0 mL sample of 0.200 M K₂CO₃ solution is added to 30.0 mL of 0.400 M Ba(NO₃)₂ solution. Barium carbonate precipitates. The concentration of barium ion, Ba²⁺, in solution after reaction is:
(A) 0.150 M  (B) 0.160 M  (C) 0.200 M  (D) 0.240 M  (E) 0.267 M

48. **E** 5 Fe²⁺ + MnO₄⁻ + 8H⁺ → 5 Fe³⁺ + Mn²⁺ + 4 H₂O

49. **B** In a titration experiment based on the equation above, 25.0 mL of an acidified Fe²⁺ solution requires 14.0 mL of standard 0.50 M MnO₄⁻ solution to reach the equivalence point. The concentration of Fe²⁺ in the original solution is:
(A) 0.0010 M  (B) 0.0056 M  (C) 0.028 M  (D) 0.090 M  (E) 0.14 M

50. **B** Commercial vinegar was titrated with NaOH solution to determine the content of acetic acid, H₃C₂H₅O₂. For 20.0 mL of the vinegar, 32.0 mL of 0.500 M NaOH solution was required. What was the concentration of acetic acid in the vinegar if no other acid was present?
(A) 1.60 M  (B) 0.800 M  (C) 0.540 M  (D) 0.600 M  (E) 0.400 M

51. **C** If 87 grams of K₂SO₄ (molar mass 174 grams) is dissolved in enough water to make 250 mL of solution, what are the concentrations of the potassium and the sulfate ions?

<table>
<thead>
<tr>
<th>[K⁺]</th>
<th>[SO₄²⁻]</th>
</tr>
</thead>
<tbody>
<tr>
<td>(A) 0.020M</td>
<td>0.020 M</td>
</tr>
<tr>
<td>(B) 1.0 M</td>
<td>2.0 M</td>
</tr>
<tr>
<td>(C) 2.0 M</td>
<td>1.0 M</td>
</tr>
<tr>
<td>(D) 2.0 M</td>
<td>2.0 M</td>
</tr>
<tr>
<td>(E) 4.0 M</td>
<td>2.0 M</td>
</tr>
</tbody>
</table>

52. **E** What volume of 0.150 M HCl is required to neutralize 25.0 mL of 0.120 M Ba(OH)₂?

53. **C** What mass of Au is produced when 0.0500 mol of Au₂S₃ is reduced completely with excess H₂?

54. **D** According to the balanced equation above, how many moles of HI would be necessary to produce 2.5 mol of I₂ starting with 4.0 mol of K₂MnO₄ and 3.0 mol of H₂SO₄?
55. Approximately what mass of \( \text{CuSO}_4 \cdot 5\text{H}_2\text{O} \) (250 g/ mol) is required to prepare 250 mL of 0.10 M copper(I) sulfate solution?
(A) 4.0 g (B) 6.2 g (C) 34 g (D) 85 g (E) 140 g

56. When 8.0 g of \( \text{N}_2\text{H}_4 \) (32 g mol\(^{-1}\)) and 92 g of \( \text{Na}_2\text{O}_2 \) (92 g mol\(^{-1}\)) are mixed together and react according to the equation above, what is the maximum mass of \( \text{H}_2\text{O} \) that can be produced?
(A) 9.0 g (B) 18 g (C) 36 g (D) 72 g (E) 144 g

57. According to the balanced equation above, how many moles of \( \text{ClO}_2^- \) (aq) are needed to react completely with 20. mL of 0.20 M \( \text{KMnO}_4 \) solution?
(A) 0.0030 mol (B) 0.0053 mol (C) 0.0075 mol (D) 0.013 mol (E) 0.030 mol

58. If 200. mL of 0.60 M \( \text{MgCl}_2 \) (aq) is added to 400. mL of distilled water, what is the concentration of \( \text{Mg}^{2+} \) (aq) in the resulting solution? (Assume volumes are additive.)
(A) 0.20 M (B) 0.30 M (C) 0.40 M (D) 0.60 M (E) 1.2 M

59. When the skeleton equation above is balanced and all coefficients reduced to their lowest whole-number terms, what is the coefficient for \( \text{H}^+ \)?
(A) 4 (B) 6 (C) 8 (D) 9 (E) 10

60. When the equation for the half-reaction above is balanced, what is the ratio of the coefficients \( \text{OH}^- / \text{CrO}_4^{2-} \)?
(A) 1: 1 (B) 2: 1 (C) 3: 1 (D) 4: 1 (E) 5: 1

61. Which, if any, of the following species is in the greatest concentration in a 0.100-molar solution of \( \text{H}_2\text{SO}_4 \) in water?
(A) \( \text{H}_2\text{SO}_4 \) molecules (B) \( \text{H}^+ \) ions (C) \( \text{HSO}_4^- \) ions (D) \( \text{SO}_4^{2-} \) ions

62. Which of the following species CANNOT function as an oxidizing agent?
(A) \( \text{Cr}_2\text{O}_7^{2-} \) (B) \( \text{MnO}_4^- \) (C) \( \text{NO}_3^- \) (D) \( \text{S} \) (E) \( \text{I} \)

63. Which of the following statements regarding the reaction represented by the equation above is incorrect?
(A) \( \text{MnO}_4^- \) is oxidized by iodide ion.
(C) The oxidation number of manganese changes from +7 to +2.
(D) The oxidation number of manganese remains the same.
(E) The oxidation number of iodine changes from -1 to 0.

64. When the equation for the half reaction above is balanced with the lowest whole-number coefficients, the coefficient for \( \text{H}_2\text{O} \) is:
(A) 2 (B) 4 (C) 6 (D) 7 (E) 14

65. Which of the following does NOT behave as an electrolyte when it is dissolved in water?
(A) \( \text{CH}_3\text{OH} \) (B) \( \text{K}_2\text{CO}_3 \) (C) \( \text{NH}_4\text{Br} \) (D) \( \text{HI} \) (E) Sodium acetate, \( \text{CH}_3\text{COONa} \)
66. Which species acts as an oxidizing agent in the reaction represented above?
   (A) H₂O  (B) ClO₄⁻  (C) ClO₂⁻  (D) MnO₂  (E) MnO₄⁻

67. When dilute nitric acid was added to a solution of one of the following chemicals, a gas was evolved. This gas turned a drop of limewater, Ca(OH)₂, cloudy, due to the formation of a white precipitate. The chemical was:
   (A) household ammonia, NH₃  (B) baking soda, NaHCO₃  (C) table salt, NaCl  
   (D) epsom salts, MgSO₄·7H₂O  (E) bleach, 5% NaOCl

Questions 68-70 refer to the reactions represented below.
   (A) H₂SeO₄(aq) + 2 Cl⁻(aq) + 2 H⁺(aq) → H₂SeO₃(aq) + Cl₂(g) + H₂O(l)
   (B) S(s) + 8 O₂(g) → 8 SO₂(g)
   (C) 3 Br₂(aq) + 6 OH⁻(aq) → 5 Br⁻(aq) + BrO₃⁻(aq) + 3 H₂O(l)
   (D) Ca³⁺(aq) + SO₄²⁻(aq) → CaSO₄(s)
   (E) PtCl₄(s) + 2 Cl⁻(aq) → PtCl₆³⁻(aq)

68. A precipitation reaction
69. A reaction in which the same reactant undergoes both oxidation and reduction
70. A combustion reaction

71. In which of the following species does sulfur have the same oxidation number as it does in H₂SO₄?
   (A) H₂SO₃  (B) S₂O₃⁻  (C) S⁵⁻  (D) S₈  (E) SO₂Cl₂

72. In the laboratory, H₂(g) can be produced by adding which of the following to 1 M HCl(aq)
   I. 1 M NH₃(aq)
   II. Zn(s)
   III. NaHCO₃(s)
   (A) I only  (B) II only  (C) III only  (D) I and II only  (E) I, II, and III

3 Cu⁰(s) + 8 H⁺(aq) + 2 NO₃⁻(aq) → 3 Cu²⁺(aq) + 2 NO(g) + 4 H₂O(l)

73. True statements about the reaction represented above include which of the following?
   I. Cu⁰(s) acts as an oxidizing agent.
   II. The oxidation state of nitrogen changes from +5 to +2
   III. Hydrogen ions are oxidized to form H₂O(l)
   (A) I only  (B) II only  (C) III only  (D) I and II  (E) II and III

74. Which of the following substances is LEAST soluble in water?
   (A) (NH₄)₂SO₄  (B) KMnO₄  (C) BaCO₃  (D) Zn(NO₃)₂  (E) Na₃PO₄

75. A pure, white crystalline solid dissolves in water to yield a basic solution that liberates a gas when excess acid is added to it. On the basis of this information, the solid could be:
   (A) KNO₃  (B) K₂CO₃  (C) KOH  (D) KHSO₄  (E) KCl
76. The graph above shows the speed distribution of molecules in a sample of a gas at a certain temperature. Which of the following graphs shows the speed distribution of the same molecules at a lower temperature (as a dashed curve)?

a.  

b.  

c.  

d.  

77. Which of the following particulate diagrams best shows the formation of water vapor from hydrogen gas and oxygen gas in a rigid container at 125°C?

a.  

b.  

c.  

d.  

Questions 78-80 refer to three gases in identical rigid containers under the conditions given in the table below.

<table>
<thead>
<tr>
<th>Container</th>
<th>A</th>
<th>B</th>
<th>C</th>
</tr>
</thead>
<tbody>
<tr>
<td>Gas</td>
<td>Methane</td>
<td>Ethane</td>
<td>Butane</td>
</tr>
<tr>
<td>Formula</td>
<td>CH₄</td>
<td>C₂H₆</td>
<td>C₄H₁₀</td>
</tr>
<tr>
<td>Molar Mass g/mol</td>
<td>16</td>
<td>30</td>
<td>58</td>
</tr>
<tr>
<td>Temperature (°C)</td>
<td>27</td>
<td>27</td>
<td>27</td>
</tr>
<tr>
<td>Pressure (atm)</td>
<td>2.0</td>
<td>4.0</td>
<td>2.0</td>
</tr>
</tbody>
</table>

78. The average kinetic energy of the gas molecules is
   a. greatest in container A
   b. greatest in container B
   c. greatest in container C
   d. the same in all three containers
79. The density of the gas, in g/L, is
   a. greatest in container A
   b. greatest in container B
   c. greatest in container C
   d. the same in all three containers

80. If the pressure of each gas is increased at constant temperature until condensation occurs, which gas will condense at the lowest pressure?
   a. Methane
   b. Ethane
   c. Butane
   d. All the gases will condense at the same pressure.

81. The figure above represents three sealed 1.0 L vessels, each containing a different inert gas at 298 K. The pressure of Ar in the first vessel is 2.0 atm. The ratio of the numbers of Ar, Ne, and He atoms in the vessels is 2:1:6, respectively. After all the gases are combined in a previously evacuated 2.0 L vessel, what is the total pressure of the gases at 298 K?
   a. 3.0 atm  b. 4.5 atm  c. 9.0 atm  d. 18 atm

Questions 82-83 refer to the following gas molecules at the conditions indicated.

a. H₂(g) molecules at 10⁻³ atm and 200°C
b. O₂(g) molecules at 20 atm and 200°C
c. SO₂(g) molecules at 20 atm and 200°C
d. NH₃(g) molecules at 20 atm and 200°C
e. NH₃(g) molecules at 20 atm and 300°C

82. Behave most like an ideal gas

83. Have lowest root-mean-square speed

\[ \text{PCl}_3(g) \leftrightarrow \text{PCl}_3(g) + \text{Cl}_2(g) \]

PCl₃(g) decomposes into PCl₃(g) and Cl₂(g) according to the equation above. A pure sample of PCl₃(g) is placed in a rigid, evacuated 1.00 L container. The initial pressure of the PCl₃(g) is 1.00 atm. The temperature is held constant until the PCl₃(g) reaches equilibrium with its decomposition products. The figures below show the initial and equilibrium conditions of the system.

84. If the decomposition reaction were to go to completion, the total pressure in the container would be:
   a. 1.4 atm  b. 2.0 atm  c. 2.8 atm  d. 3.0 atm
85. When 6.0 L of He(g) and 10. L of N₂(g), both at 0 °C and 1.0 atm, are pumped into an evacuated 4.0 L rigid container, the final pressure in the container at 0 °C is
   a. 2.0 atm  b. 4.0 atm  c. 6.4 atm  d. 16 atm

86. The vapor pressure of pure water at 25 °C is 24.0 mm Hg. What is the expected vapor pressure at 25 °C of an ideal solution of a nonvolatile nonelectrolyte in which the mole fraction of water is 0.900?
   a. 1.48 mm Hg  b. 2.40 mm Hg  c. 21.6 mm Hg  d. 24.0 mm Hg

87. Of the following, the best explanation for the fact that most gases are easily compressed is that the molecules in a gas:
   a. are in constant motion  
   b. are relatively far apart  
   c. have relatively small masses  
   d. move slower as temperature decreases

88. A sample of an unknown gas from a cylinder is collected over water in the apparatus shown above. After all the gas sample has been collected, the water levels inside and outside the gas collection tube are made the same. Measurements that must be made to calculate the molar mass of the gas include all of the following EXCEPT:
   a. temperature of the water  
   b. volume of gas in the gas-collection tube  
   c. initial and final mass of the gas cylinder  
   d. mass of the water in the apparatus

<table>
<thead>
<tr>
<th>Gas</th>
<th>Amount</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ar</td>
<td>0.35 mol</td>
</tr>
<tr>
<td>CH₄</td>
<td>0.90 mol</td>
</tr>
<tr>
<td>N₂</td>
<td>0.25 mol</td>
</tr>
</tbody>
</table>

89. Three gases in the amounts shown in the table above are added to a previously evacuated rigid tank. If the total pressure in the tank is 3.0 atm at 25°C, the partial pressure of N₂ (g) in the tank is closest to:
   a. 0.75 atm  
   b. 0.50 atm  
   c. 0.33 atm  
   d. 0.25 atm  
   e. 0.17 atm

90. At approximately what temperature will 40. g of Argon gas at 2.0 atm occupy a volume of 22.4 L.
   a. 600 K  
   b. 550 K  
   c. 270 K  
   d. 140 K

91. Of the following gases, which has the greatest average molecular speed at 298 K?
   a. Cl₂  
   b. NO  
   c. H₂S  
   d. HCN  
   e. PH₃