Instructions

1. Do not open the exam until you are told to start.

2. This exam is closed note and closed book. You are not allowed to use any outside material while taking this exam.

3. Use the spaces provided to write down your answers. To receive full credit, you must show all work. Do not write answers on any other pieces of paper. If you need more room, write on the back of the exam and be sure to include a note describing where the work is located.

4. When solving numerical problems, make sure you include the proper units in your final answer.

5. If a question asks for a response in sentence or paragraph form, make sure you respond in that format.

6. Useful data for the exam and a periodic table are provided on the last page of the exam. Carefully tear out these sheets if you wish.

<table>
<thead>
<tr>
<th>Page #</th>
<th>Points possible</th>
<th>Points awarded</th>
</tr>
</thead>
<tbody>
<tr>
<td>2-4</td>
<td>35</td>
<td></td>
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<tr>
<td>5</td>
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<td>6</td>
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<td>Total</td>
<td>100</td>
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</tbody>
</table>
Multiple Choice

Unless otherwise directed, choose the single best answer for each question. When balancing chemical equations use smallest whole number coefficients. (3 points each, 5 points for #3)

1. When atoms react, they do so by gaining electrons, losing electrons, or sharing electrons. The ease with which they can do these things affects their reactivity. Based on your knowledge of effective nuclear charge, the trends in first ionization energy, the trends in atomic size, and the ways in which metals and non-metals react, which of the atoms below would you expect to be the most reactive?

a.) Au  
b.) Ba  
c.) W  
d.) Cs  
e.) They would all be equally reactive.

2. How many moles of lithium are needed to produce 0.60 moles of Li₃N? Use the equation below to help you solve the problem.

\[ 6 \text{ Li(s)} + \text{N}_2(g) \rightarrow 2\text{Li}_3\text{N(s)} \]

- a.) 0.30
- b.) 1.8
- c.) 0.20
- d.) 0.40
- e.) None of the above.

3. A 13.25 mL sample of HCl is titrated with 0.2345 M Ca(OH)₂. If the titration requires 25.36 mL of the Ca(OH)₂ solution to reach the equivalence point, what was the molarity of the original HCl solution? (5 points)

\[ 2\text{HCl(aq)} + \text{Ca(OH)}_2(aq) \rightarrow 2\text{H}_2\text{O(aq)} + \text{CaCl}_2(aq) \]

- a.) 0.4488 M HCl  
- b.) 0.01189 M HCl  
- c.) 0.3081 M HCl  
- d.) 0.8976 M HCl  
- e.) None of the above.

\[ \frac{25.36 \text{ mL}}{1000 \text{ mL}} \times \frac{0.2345 \text{ mol Ca(OH)}_2}{2 \text{ mol HCl}} = 0.01189 \text{ M HCl} \]

\[ [\text{HCl}] = \frac{0.01189 \text{ M HCl}}{0.01325} = 0.8976 \text{ M HCl} \]
4. Certain reactions will give off or release energy when they occur. This energy can be used in a number of ways, including powering vehicles, heating homes, and cooking. The reaction shown below releases 241.8 kJ of energy. Since it releases energy, energy can be listed as a product in the balanced chemical equation and can be treated like any other product in stoichiometric calculations.

\[
2\text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(g) + 241.8 \text{ kJ}
\]

Molar Mass (g/mole)

\[
\begin{array}{ccc}
2.01588 & 31.9988 & 18.0153 \\
\end{array}
\]

How much energy will be released when 0.453 g of H\textsubscript{2} reacts with excess O\textsubscript{2}? (A kJ is just a unit of energy)

\[
\frac{0.453 \text{ g H}_2}{2.01588 \text{ g H}_2} = \frac{241.8 \text{ kJ}}{2 \text{ mol H}_2}
\]

When ethyne is combusted, it reacts with O\textsubscript{2} and produces CO\textsubscript{2} and H\textsubscript{2}O. The balanced chemical equation for this reaction is shown below. **Use this equation and the diagrams below to answer the following three questions.**

\[
2\text{C}_2\text{H}_2(g) + 5\text{O}_2(g) \rightarrow 4\text{CO}_2(g) + 2\text{H}_2\text{O}(g)
\]

5. Which compound is the limiting reagent?

a.) C\textsubscript{2}H\textsubscript{2}

b.) O\textsubscript{2}

c.) CO\textsubscript{2}

d.) H\textsubscript{2}O

e.) None of the above.

6. What was the **total** number of molecules you started with before the reaction?

a.) 8

b.) 7

c.) 6

d.) 5

e.) None of the above.
7. What will happen to the pressure after the reaction if the temperature is increased by 100°C?

a.) The pressure will increase.
b.) The pressure will decrease.
c.) The pressure will not change.
d.) None of the above.

8. Calcium chloride is soluble in water. A 5.55 g sample of CaCl₂ is dissolved in enough water to make 110.0 mL of solution. What is the concentration of Cl⁻ in the solution?

\[
\begin{align*}
\text{CaCl}_2 & \xrightarrow{1 \text{ mol CaCl}_2} 1\text{ mol CaCl}_2 \\
& \xrightarrow{2 \text{ mol Cl}^-} 0.100 \text{ mol Cl}^- \\
= & \frac{0.100 \text{ mol Cl}^-}{0.110 \text{ L}} = 0.909 \text{ M Cl}^- \\
\end{align*}
\]

a.) 0.00550 M
b.) 0.455 M
c.) 0.668 M
d.) 0.909 M
e.) None of the above

9. What is the mass of AgNO₃ in 145 mL of a 4.31 M AgNO₃ solution?

\[
\begin{align*}
\text{AgNO}_3 & \xrightarrow{1 \text{ mol AgNO}_3} 1\text{ mol AgNO}_3 \\
& \xrightarrow{169.873 \text{ g}} 1\text{ mol AgNO}_3 \\
= & \frac{145\text{ mL} \times 1\text{ L}}{1000\text{ mL}} \times 4.31\text{ M AgNO}_3 = 0.6169163\text{ g AgNO}_3 \\
\end{align*}
\]

a.) 5.05 g
b.) 106 g
c.) 24.6 g
d.) 170 g
e.) None of the above

10. When 31.52 mL of H₂O is added to 56.10 mL of an HCl solution, the concentration of the new solution is 0.2354 M HCl. What was the concentration of the original HCl solution?

\[
\begin{align*}
\text{HCl} & \xrightarrow{31.52\text{ mL H}_2\text{O}} \text{HCl} \\
= & \frac{0.2354 \text{ M HCl}}{1000\text{ mL}} = 0.02354 \text{ M HCl} \\
\end{align*}
\]

a.) 0.1323 M
b.) 0.3677 M
c.) 0.4190 M
d.) 0.6544 M
e.) None of the above

11. A given set of d-orbitals contains how many degenerate orbitals?

a.) 1
d.) 6
c.) 10
b.) 3
e.) 5

\[
\begin{align*}
87.62\text{ mL} & \times 0.2354 \text{ mol HCl} = 0.02062157\text{ mol HCl} \\
= & \frac{0.02062157\text{ mol HCl}}{0.056105\text{ L}} = 0.367460 \text{ M HCl} \\
\end{align*}
\]
12. In the table below, the name or formula for a chemical compound is given. Fill in the table with the corresponding name or formula of the chemical compound. (4 points)

<table>
<thead>
<tr>
<th>NAME</th>
<th>FORMULA</th>
</tr>
</thead>
<tbody>
<tr>
<td>bromic acid</td>
<td>HBrO₃</td>
</tr>
<tr>
<td>ammonium carbonate</td>
<td>(NH₄)₂CO₃</td>
</tr>
</tbody>
</table>

13. In the space provided below give a.) the full electron configuration of the element that is underlined, b.) draw the Bohr representation of each element c.) calculate the effective nuclear charge (Z\text{eff}) felt by the valence electrons in each of the atoms, and d) state which atom is the largest. (10 points)

a.) Mg

\[ \text{Mg: } 1s^2 \text{2s}^2 \text{2p}^6 \text{3s}^2 \]

b.)

\[ \text{Al: } 1s^2 \text{2s}^2 \text{2p}^6 \text{3s}^2 \text{3p}^1 \]

c.)

\[ Z\text{eff} = 13 - 10 \]
\[ = 3 \text{ or } +3 \]

d.) Mg b/c smallest Z\text{eff}

14. Suppose both of the elements in question 16 above absorbed a photon of 400 nm light and had an electron promoted to an excited state.

a.) What color of light was absorbed? (2 points)

\[ \text{Violet} \]

b.) In which atom did the electron travel the longest distance? (2 points)

\[ \text{Mg} \text{ b/c smallest } Z\text{eff} \]
15. A sample of Co reacts with 22.50 mL of 0.1059 M HI to produce H₂(g) as described in the chemical equation below. The H₂(g) is collected in a container over water at 17.00°C. Assume the gas is at the same temperature as the liquid. The volume of the collected gas is 26.97 mL.

\[
\text{Co(s) + 2HI(aq) \rightarrow CoI}_2(\text{aq}) + \text{H}_2(\text{g})
\]

a.) How many moles of H₂ gas are produced? (4 points)

\[
\text{Moles of H}_2 = \frac{\text{Volume of H}_2 \times \text{Density of H}_2}{\text{Molar Volume}} = \frac{0.02697 \text{ L} \times 0.0898 \text{ g/L}}{0.014 \text{ L} \times \text{m}_2} = 0.001191 \text{ mol H}_2
\]

b.) What is the total pressure (in mmHg) inside the container holding the H₂ gas? (8 points)

\[
\rho V = nRT \quad \rho_{H_2} = \frac{n_{H_2} \times R \times T}{V} = \frac{0.001191 \text{ mol} \times 0.08206 \text{ L} \times \text{atm/molK} \times 298 \text{ K}}{0.02697 \text{ L}} = 1.051 \text{ atm}
\]

\[\rho_{H_2} = \rho_{H_2} + \rho_{H_2O}\]

\[P_{total} = 769.3 \text{ mmHg} + 14.53 \text{ mmHg} = 813.8 \text{ mmHg}\]

16. What is the condensed orbital box diagram for Mo? How many valence electrons does Mo contain? (3 points)

\[\text{Mo: [Kr] 5s}^1 4d^5\]

17. What is the full electron configuration for Cl? (3 points)

\[\text{Cl: } 1s^2 2s^2 2p^6 3s^2 3p^5\]

18. In the space provided below, draw a picture of a p-orbital. Make sure to point out the location of the nucleus. (3 points)

19. What is the condensed electron configuration of Fe⁴⁺? (3 points)

\[\text{Fe: [Ar] 4s}^2 3d⁶ \quad \text{Fe}^{⁴⁺}: [Ar] 3d⁶\]
20. Ammonia reacts with oxygen to form nitrogen dioxide and water as described in the equation shown below.

\[
4\text{NH}_3(g) + 7\text{O}_2(g) \rightarrow 4\text{NO}_2(g) + 6\text{H}_2\text{O}(g)
\]

Molar Mass (g/mol)

\[
\begin{array}{c|c|c|c}
4\text{NH}_3 & 7\text{O}_2 & 4\text{NO}_2 & 6\text{H}_2\text{O} \\
17.03056 & 31.9988 & 46.0055 & 18.0153 \\
\end{array}
\]

a.) If 21.0 g of NH\(_3\) reacts with 21.0 g of O\(_2\), what mass of NO\(_2\) will be produced? (8 points)

\[
\begin{array}{c|c|c|c|c|c|c}
21.0 \text{ g NH}_3 & 1 \text{ mol NH}_3 & 4 \text{ mol NO}_2 & 46.0055 \text{ g NO}_2 \\
17.03056 \text{ g NH}_3 & 1 \text{ mol NH}_3 & 1 \text{ mol O}_2 & 17.2525 \text{ g O}_2 \\
\end{array}
\]

b.) How much of each reactant will be left after the reaction is complete? (4 points)

\[
\begin{array}{c|c|c|c|c|c|c}
21.0 \text{ g O}_2 & 1 \text{ mol O}_2 & 4 \text{ mol NO}_2 & 46.0055 \text{ g NO}_2 \\
31.9988 \text{ g O}_2 & 7 \text{ mol O}_2 & 1 \text{ mol NH}_3 & 17.2525 \text{ g NH}_3 \\
\end{array}
\]

\[
\text{NH}_3 \text{ left} = 21.0g - 6.38166g = 14.613g
\]

\[
\text{O}_2 \text{ left} = 21.0g - 17.2525g = 3.7475g
\]

c.) If the percent yield of the reaction is 79.4%, what mass of NO\(_2\) was actually produced? (2 points)

\[
\text{Yield} = \frac{\text{Actual yield}}{\text{Max yield}} \times 100\%
\]

\[
\frac{6.38166g}{17.2525g} \times 100\% = 79.4\%
\]

\[
\text{A. Y.} = 13.6198g
\]

\[
= 13.7g
\]
21. a. Under what temperature conditions does the ideal gas law fail to accurately calculate the properties of a gas? (2 points)

Low T

b. Using the diagram below and words, explain why the ideal gas law does not work under these conditions. (3 points)

At low temperatures, the gas particles move slower and they start to feel attractions for the other gas particles. The ideal gas law does not account for these attractions.

\[ \text{Pressure Meter} \]

\[ \text{lower } P \text{ than calculated} \]
### Conversion Factors, Constants, and Periodic Table

**Avogadro’s Number:**  \(6.022 \times 10^{23}\) particles/mole

**Pressure Conversion:**  \(760\text{ mmHg (torr)} = 1\text{ atm (exact)}\)

**Ideal Gas Constant:**  \(0.08206\ \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\)

**Temperature conversion:**  \(T\ [\text{K}] = 273.15 + T\ [\circ\text{C}]\)

**Percent Yield:**  \(\%\ Yield = \frac{A.Y.}{T.Y.} \times 100\%\)

**Gas Equations:**  
- \(P_A = P_{\text{tot}} \cdot \chi_A\)
- \(PV = nRT\)

### Partial Pressure of Water at Various Temperatures

<table>
<thead>
<tr>
<th>Temperature (°C)</th>
<th>Partial Pressure (mmHg)</th>
</tr>
</thead>
<tbody>
<tr>
<td>5.00</td>
<td>6.54</td>
</tr>
<tr>
<td>15.00</td>
<td>12.79</td>
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<td>17.00</td>
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<td>23.00</td>
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<td>30.00</td>
<td>31.82</td>
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<tr>
<td>50.00</td>
<td>92.51</td>
</tr>
</tbody>
</table>

### Periodic Table

| 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 | 11 | 12 | 13 | 14 | 15 | 16 | 17 | 18 | 19 | 20 | 21 | 22 | 23 | 24 | 25 | 26 | 27 | 28 | 29 | 30 | 31 | 32 | 33 | 34 | 35 | 36 | 37 | 38 | 39 | 40 | 41 | 42 | 43 | 44 | 45 | 46 | 47 | 48 | 49 | 50 | 51 | 52 | 53 | 54 | 55 | 56 | 57 | 58 | 59 | 60 | 61 | 62 | 63 | 64 | 65 | 66 | 67 | 68 | 69 | 70 | 71 | 72 | 73 | 74 | 75 | 76 | 77 | 78 | 79 | 80 | 81 | 82 | 83 | 84 | 85 | 86 | 87 | 88 | 89 | 90 | 91 | 92 | 93 | 94 | 95 | 96 | 97 | 98 | 99 | 100 | 101 | 102 | 103 |