7.A.1 Problems – Measuring Matter
Section 10.1 of your textbook.

1. How many atoms of potassium does 1 mol of potassium contain?

Determine the number of representative particles in each substance.

2. A. 0.250 mol of silver

   B. 8.56 E – 3 mol of sodium chloride

3. A. 35.3 mol of carbon dioxide

   B. 0.425 mol of nitrogen (N₂)

Perform the following conversions:

4. 1.51 E 13 atoms of Si to mol Si

5. 4.25 E – 2 mol of H₂SO₄ to molecules of H₂SO₄

6. 8.95 E 25 molecules of CCl₄ to mol of CCl₄

7. 5.90 mol of Ca to atoms of Ca
1. Which contains more atoms, a mole of silver atoms or a mole of gold atoms? Explain your answer.

2. Which has more mass, a mole of potassium or a mole of sodium? Explain your answer.

For the following two problems, calculate the mass of each element

3. 5.22 mol of He

4. 0.0455 mol of Ni

5. How many atoms are there in 2.22 grams of Ti?

6. Arrange from least to most in moles: 3.00 E 24 atoms Ne, 4.25 mol Ar, 2.69 E 24 atoms Xe, 65.96 g Kr.
How many moles of oxygen atoms are contained in each compound?

1. 2.50 mol of KMnO₄

2. 45.9 mol of CO₂

3. What is the molar mass of sodium nitrate, NaNO₃?

4. Garlic. Determine the molar mass of allyl sulfide, the compound responsible for the smell of garlic. The chemical formula of allyl sulfide is \((C₃H₅)₂S\).

How many moles are in 100.0 g of each compound?

5. dinitrogen monoxide (N₂O)

6. methanol (CH₃OH)

7. Iron. How many moles of iron can be recovered from 100.0 kg of Fe₂O₄? Hint: convert kilograms to grams first.
1. **Cholesterol.** Heart disease is linked to high blood cholesterol levels. What is the percent composition of the elements in a molecule of cholesterol (C_{27}H_{45}OH)?

2. What is the percent composition of the elements in calcium carbonate: CaCO_{3}? 

3. What is the percent composition of the elements in barium acetate, Ba(C_{2}H_{3}O_{2})_{2}?
1. Which of the following formulas represent the empirical and molecular formulas of the same compound? Explain your answer.

   NO      N₂O      NO₂      N₂O₄      N₂O₅

Determine the empirical formula for each compound:

2. Ethylene (C₂H₄)

3. Ascorbic acid (C₆H₈O₆)

4. Naphthalene (C₁₀H₈)

5. What is the formula of a compound with a percent composition of: Ba = 69.58%, C = 6.09%, O = 24.32%?
1. **Desiccants.** Why are certain electronic devices transported with desiccants?

Write the formula for the following hydrates:

2. nickel (II) chloride hexahydrate

3. cobalt (II) chloride hexahydrate

4. magnesium carbonate pentahydrate

5. sodium sulfate decahydrate

6. A 1.628-g sample of a hydrate of magnesium iodide (MgI₂) is heated until its mass is reduced to 1.072 g and all water has been removed. What is the formula of the hydrate?
In your lab groups, complete the tasks involving molar calculations. Use the balances to measure masses of objects, then make calculations using your knowledge of molar relations and molar masses. Show all calculations, or get a fee.

**Materials:**
- 2 of 5 different elements: Zn, Al, Pb, Cu, and Sn
- A paraffin candle – chemical formula = C_{25}H_{52}
- 100 mL beakers
- 100 mL graduated cylinders

**Task 1 (8 Points)** Select two different metals from the green cart and write their names down in the blank. Then, measure and record their masses. Next, determine:

1. How many moles of the metal you have in your sample, and
2. How many atoms of the metal you have in your sample.

**Metal 1:** ______________  **Molar Mass of Metal 1:** ______________

Mass of metal 1: ______________

1. Moles of metal 1: ______________
2. Atoms of metal 1: ______________

**Calculations:**

**Metal 2:** ______________  **Molar Mass of Metal 2:** ______________

Mass of metal 2: ______________

1. Moles of metal 2: ______________
2. Atoms of metal 2: ______________

**Calculations:**
**Task 2 (4 pts)** Measure and record the mass of your candle, then determine:
1. The molar mass of paraffin,
2. How many moles of paraffin your sample has,
3. How many moles of carbon your sample has, and
4. How many moles of hydrogen your sample has.

Candle’s mass: __________________

1. Molar mass of paraffin: ____________
2. Moles of paraffin: _______________
3. Moles of carbon: ________________
4. Moles of hydrogen:_______________

Calculations:

**Task 3 (4 pts)**
1. Calculate the molar mass of water.
2. Calculate the mass that 4.05 moles of water has.
3. Use the 100 ml beakers and balances to weigh out that mass. You will need to subtract the beaker’s mass from the water/beaker mass to get your exact water weight. Finally, pour the water from the beaker into one of the 100 mL graduated cylinders determine its volume.

1. Molar mass of water: ____________
2. Calculated mass of 4.05 moles of water: ___________________
3. Volume of 4.05 moles of water: _______________

Calculations:
Objective:
Experimentally determine the formula of a hydrated chemical.

Overview:
Many chemicals, called hydrates, have water molecules entrained within their crystal matrices. One of these is cobalt (II) chloride – CoCl₂ · ? H₂O. The question mark will be what you determine in this lab!

Safety:
1. Wear goggles throughout the lab – even when you are finished but others are still working.
2. Tie loose clothing and hair back.
3. You will need to move the hot test tube carefully with metal tongs.
4. Cool the test tube thoroughly before measuring its mass – hot items have an apparently greater mass than cold ones (how can this be?).

Materials:
Ring Stand
Bunsen Burner
Triple Beam Balance
Small Test Tube
Metal Tongs
Cobalt (II) Chloride

Procedure:
1. Read through the lab before starting.
2. Measure and record the mass of a small, dry test tube plus an attached clamp.
3. Obtain about 3 grams of cobalt (II) chloride using the balances. Use a piece of weighing paper with a crease in it to put your sample on, then use the crease as a channel to pour it into the test tube. Next, measure and record the mass of the test tube and cobalt (II) chloride.
4. Heat your sample in the test tube at an angle for 10 minutes, or until steam stops, using the metal tongs. Record your observations in question 1 as this process continues.
5. Let your tube cool down for 10 minutes.
6. Measure and record the mass of the tube and anhydrous (dry) cobalt (II) chloride.
7. **Clean up and Disposal:**
   a. Feel the temperature of the test tube with your hand (being sure first that it is not too hot).
   b. Fill the test tube halfway with deionized water, and then feel it again. Write down your observations in question 2.
   c. Let the water dissolve your hydrate for a few minutes. Heating it GENTLY will speed this up – but don’t boil it!
   d. Pour the re-dissolved cobalt chloride into the container labeled “Cobalt Chloride Waste.”
8. Make sure your area is **tidy** before you leave, so you don’t lose points.
Data Table: (5 Points)

As you go through the lab procedure, fill out the following data table. Measure masses to the hundredths place.

<table>
<thead>
<tr>
<th>Measured Item:</th>
<th>Mass (grams)</th>
<th>Moles</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. Test tube and clamp</td>
<td></td>
<td></td>
</tr>
<tr>
<td>2. Test tube, clamp, and CoCl₂ · ? H₂O</td>
<td></td>
<td></td>
</tr>
<tr>
<td>3. CoCl₂ · ? H₂O</td>
<td></td>
<td></td>
</tr>
<tr>
<td>4. Test tube, clamp, + anhydrous CoCl₂</td>
<td></td>
<td></td>
</tr>
<tr>
<td>5. Mass of water lost</td>
<td></td>
<td></td>
</tr>
<tr>
<td>6. Molar mass of water</td>
<td></td>
<td></td>
</tr>
<tr>
<td>7. Moles of water</td>
<td></td>
<td></td>
</tr>
<tr>
<td>8. Mass of anhydrous CoCl₂</td>
<td></td>
<td></td>
</tr>
<tr>
<td>9. Molar mass of anhydrous CoCl₂</td>
<td></td>
<td></td>
</tr>
<tr>
<td>10. Moles of anhydrous CoCl₂</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Data Analysis:

(2 pts for your calculations, 1 pt for proposed formula, 1 E. C. if your formula is correct)

Calculate the formula of hydrated cobalt (II) chloride. Round to the nearest whole number. Use your notes from section 7.A.6 to help you. **Be sure to show all your steps!** I’m grading on the process more than the result.

Calculations Work Area:

Write your proposed formula here: ________________________________
Questions:
1. (2 pts) What did you observe as you heated the test tube? Write complete observations down.

2. (2 pts) Explain thoroughly what happened when you poured water into your cooled, anhydrous sample.

3. (2 pts) Why do you think there was a temperature change when you added water?

4. (2 pts) Why might heating not be a suitable method for determining the water of hydration for all hydrates?
Solve these problems and enter your responses in the bubble boxes. Be sure to erase completely, and enter your student ID in the box on the right.

Each question is worth one point. If you turn this in late, you will only receive half credit for your work.

1. How many moles are there in 341 grams of beryllium nitrate?
   A) 2.56 moles  B) 2.87 moles  
   C) 3.02 moles  D) 3.14 moles

2. What is the molar mass of ammonium nitrate?
   A) 76.50 g/mole  B) 78.13 g/mole  
   C) 80.06 g/mole  D) 85.18 g/mole

3. How many formula units are there in 4.50 moles of sodium phosphate (Na₃PO₄)?
   A) 2.71 E 24 formula units  B) 1.45 E 24 formula units  
   C) 3.89 E 24 formula units  D) 1.99 E 24 formula units

4. How many molecules are there in 48.2 grams of carbon dioxide?
   A) 1.68 E 23 molecules  B) 2.87 E 23 molecules  
   C) 6.59 E 23 molecules  D) 4.56 E 23 molecules

5. Determine the empirical formula for methyl acetate, with a percent composition of 48.64% carbon, 8.16% hydrogen, and 43.02% oxygen.
   A) C₂H₆O  B) C₂H₅O₂  
   C) C₃H₆O  D) C₃H₆O₂

6. Determine the molecular formula of succinic acid, with a molar mass of 118.1 g/mol. Its percent composition is 40.68% carbon, 5.08% hydrogen, and 54.24% oxygen.
   A) C₂H₅O₂  B) C₄H₆O₄  
   C) C₂H₆O  D) C₃H₄O₂

7. Which of the titanium-containing minerals - rutile (TiO₂) or ilmenite (FeTiO₃) - has a larger percent by mass of titanium?
   A) rutile  B) ilmenite

8. What is the percent by mass of titanium in rutile?
   A) 59.93%  B) 45.67%  C) 31.55%  D) 67.88%
7.B.1 Problems – Defining Stoichiometry

Section 11.1 of your textbook.

1. What relationships can be determined from a balanced chemical equation?

Consider the following equation: \(4\text{Al} + 3\text{O}_2 \rightarrow 2\text{Al}_2\text{O}_3\) Interpret it in terms of:

2. Particles

3. Moles


Solid silicon dioxide (SiO\(_2\)), often called silica, reacts with hydrofluoric acid (HF) solution to produce the gas silicon tetrafluoride (SiF\(_4\)) and water.

5. Write the balanced chemical equation for the reaction.

6. List three mole ratios from the previous equation, and explain how you would use one of your ratios in a stoichiometric calculation.
1. What is the first step in all stoichiometric calculations?

2. Ethanol (C₂H₅OH), also known as grain alcohol, can be made from the fermentation of sugar (C₆H₁₂O₆). The unbalanced chemical equation is shown below. Balance the chemical reaction and determine the mass of C₂H₅OH produced from 750 g of C₆H₁₂O₆.

\[ \text{____ C₆H₁₂O₆ } \rightarrow \text{____ C₂H₅OH } + \text{____ CO₂} \]

Rocket Fuel. The exothermic reaction between liquid hydrazine (N₂H₂) and liquid hydrogen peroxide (H₂O₂) is used to fuel rockets. The products are nitrogen gas and water.

3. Write the balanced chemical equation.

4. How many grams of hydrazine are needed to produce 10.0 moles of nitrogen gas?
1. How is a mole ratio used to find the limiting reactant?

2. Iron Production. Iron is obtained commercially by the reaction of hematite (Fe₂O₃) with carbon monoxide. How many grams of iron is produced when 25.0 mol of hematite reacts with 30.0 mol of carbon monoxide?

   \[ \text{Fe}_2\text{O}_3(s) + 3\text{CO}(g) \rightarrow 2\text{Fe}(s) + 3\text{CO}_2(g) \]

3. Lithium reacts spontaneously with bromine to produce lithium bromide. Balance the chemical reaction.

   \[ \underline{____} \text{Li} + \underline{____} \text{Br}_2 \rightarrow \underline{____} \text{LiBr} \]

   If 25.0 grams of lithium and 25.0 grams of bromine are present at the beginning of the reaction, determine:
   4. The limiting reactant.
   5. The mass of lithium bromide produced.
   6. The mass of excess reactant.
1. What is the difference between actual yield and theoretical yield?

2. (2 Points) Van Arkel Process. Pure zirconium is obtained using the two-step Van Arkel process. In the first step, impure zirconium and iodine are heated to produce zirconium iodide (ZrI₄). In the second step, ZrI₄ is decomposed to produce pure zirconium.

   \[ \text{ZrI}_4(s) \rightarrow \text{Zr}(s) + 2\text{I}_2(g) \]

   Determine the percent yield of zirconium if 45.0 g of ZrI₄ is decomposed and 5.00 g of pure Zr is obtained.

3. (2 Points) Chlorine forms from the reaction of hydrochloric acid with manganese (IV) oxide. The balanced equation is:

   \[ \text{MnO}_2 + 4\text{HCl} \rightarrow \text{MnCl}_2 + \text{Cl}_2 + 2\text{H}_2\text{O} \]

   Calculate the theoretical yield and the percent yield of chlorine if 86.0 g of MnO₂ and 50.0 g of HCl react. The actual yield of Cl₂ is 20.0 g.
### Objective:
Experimentally determine how much product is obtained from an amount of limiting reactant, and calculate the theoretical maximum product.

### Overview:
In chemical reactions, limiting reactants are consumed first. What remains is the excess reactant and products of the reaction.

In this lab you will predict the amount of copper metal you can make in a reaction of zinc and copper sulfate.

The balanced reaction between zinc and copper (II) sulfate is:

\[
\text{Zn}(s) + \text{CuSO}_4(aq) \rightarrow \text{ZnSO}_4(aq) + \text{Cu}(s)
\]

### Safety:
1. Wear goggles throughout the lab – even when you are finished but others are still working.
2. Tie loose clothing and hair back.
3. Copper (II) sulfate is mildly toxic and corrosive to skin, avoid contact with it.

### Materials:
- Ring Stand
- Bunsen Burner
- Triple Beam Balance
- Tweezers
- Copper (II) Sulfate Pentahydrate \( \text{CuSO}_4 \cdot 5 \text{H}_2\text{O} \)
- Spatula
- Funnel
- Filter Paper
- Stirring Rod
- 250 mL Beaker
- Deionized Water
- Zinc Metal - Zn

### Clean up and Disposal:
1. Pour all waste solutions down the drain.
2. Dispose of filter and copper remnants in the trash.
3. Place used zinc pellets in the receptacle at the front of the room.
Procedure: **Day 1.**
1. Obtain ~ 3 grams of zinc and record the exact mass in the data table.
2. Measure ~ 5 grams of copper (II) sulfate pentahydrate and record the exact mass.
3. In your 250 mL beaker, dissolve the copper (II) sulfate in 75.0 mL deionized water. Add the zinc metal, and record your observations in question 4 as time goes by.
4. Gently heat your beaker and stir the contents to get it in solution.
5. Heat the copper solution gently while stirring with the glass rod until the blue color fades to clear (about 10 – 15 minutes).
6. Use the metal spatula and the stirring rod to break the copper flakes off the zinc.
7. Decant as much of the clear solution down the drain as you can, being careful not to drop any copper into the sink!
8. Use the tweezers to remove the remaining zinc pieces from the solution – you can rinse the pieces with water to get all the remaining copper bits off. Put the zinc in the zinc disposal beaker at the front of the room.
9. Decant the water carefully out of your beaker down the drain, then rinse and decant with deionized water three more times to remove dissolved solids.
10. Get a piece of filter paper and label it with your group name and period. Measure its exact weight, and record it.
11. Build a funnel assembly, and filter your copper metal. Use the squirt bottles to rinse any residual copper out of the bottom of your beaker and into the funnel. When all the free water has gone through the funnel, pick up your filter with copper and place it in the container at the front of the room for drying.
12. Tidy up your lab area and work on homework or calculations for the rest of class.

Procedure: **Day 2.**
1. Get your filter from the container and record its exact mass.
2. Subtract the empty filter mass from the filter + copper mass and record the copper mass in your data table.
3. You are ready to do the calculations for the lab, so make sure your area is tidy before you leave, and get to work.

**Data Table: (5 Points)**
Measure masses to the hundredths place.

<table>
<thead>
<tr>
<th>Measured Item:</th>
<th>Mass</th>
<th>Molar Mass</th>
<th>Moles</th>
</tr>
</thead>
<tbody>
<tr>
<td>Zinc</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Copper (II) Sulfate Pentahydrate</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Dry Filter Paper</td>
<td></td>
<td></td>
<td>X</td>
</tr>
<tr>
<td>Filter Paper and Copper</td>
<td></td>
<td></td>
<td>X</td>
</tr>
<tr>
<td>Copper</td>
<td></td>
<td>X</td>
<td>X</td>
</tr>
</tbody>
</table>
Calculations and Questions: Answer the following in complete sentences, and show all work for calculations.

1. (2 Points) What was the limiting reactant in this lab, zinc or copper sulfate? What evidence do you cite to support your answer?

2. (3 Points) Use stoichiometry to calculate the mass of copper possible from the exact mass of copper sulfate pentahydrate that you used. Show all your work, and box your answer.

3. (2 Points) What was the percent yield in your experiment? The equation is as follows:

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

Theoretical Yield = the value from question 2
Actual Yield = the weighed mass of copper.

4. (2 Points) Describe what happened when you added the zinc metal to the solution of copper sulfate.
1. When solid copper is added to nitric acid, copper (II) nitrate, nitrogen dioxide, and water are produced. List four mole ratios for the reaction (2 points).

\[ \text{Cu}_3 + 4 \text{HNO}_3 \rightarrow \text{Cu(NO}_3)_2 + 2 \text{NO}_2 + 2 \text{H}_2\text{O} \]

2. (3 points total). Carbon dioxide is released into the atmosphere through the combustion of heptane (C\(_7\)H\(_{16}\)) in gasoline. Balance the reaction (1 point) and calculate the mass of heptane needed to release 5.00 moles of CO\(_2\) (2 points).

\[
\begin{align*}
\underline{\text{C}}_7\text{H}_{16} & \quad \underline{\text{O}}_2 \\
\rightarrow & \quad \underline{\text{CO}}_2 + \underline{\text{H}}_2\text{O}
\end{align*}
\]

3. (2 points) Chloroform (CHCl\(_3\)) is produced by a reaction between methane and chlorine.

\[ \text{CH}_4 \quad + \quad 3\text{Cl}_2 \quad \rightarrow \quad \text{CHCl}_3 + 3\text{HCl} \]

How many grams of CH\(_4\) are needed to produce 50.0 g CHCl\(_3\)?

4. (3 Points Total) The decomposition of ammonium nitrate during heating releases nitrogen gas, water vapor, and oxygen gas. Balance this reaction (1 point), and determine how many moles of ammonium nitrate you would need to decompose to produce 158 moles of oxygen gas (2 points).

\[
\begin{align*}
\underline{\text{NH}}_4\text{NO}_3 & \quad \rightarrow \\
\underline{\text{N}}_2 + \underline{\text{H}}_2\text{O}_\text{(g)} + \underline{\text{O}}_2\text{(g)}
\end{align*}
\]
One proposed equation for the reaction of sugar with potassium nitrate is:

$$C_6H_{12}O_6 + 4 \text{KNO}_3 \rightarrow 4 \text{CO}_2 + 2 \text{CO} + 6 \text{H}_2\text{O} + 2 \text{K}_2\text{O} + 2 \text{N}_2$$

1. (2 Points) If you mixed 3.5 moles of sugar and 13.0 moles of potassium nitrate and ignited them, which would be the limiting reactant?

2. (3 Points) How many grams of water would you end up with if 14 grams of sugar reacted completely? Use the given balanced equation for mole ratios.

3. (3 Points) How many grams of potassium nitrate would you need in order to consume 0.30 grams of sugar completely? Use the equation above for mole ratios.
1. When aluminum is mixed with iron (III) oxide, iron and aluminum oxide are produced along with a large amount of heat. The balanced reaction is: \( \text{Fe}_2\text{O}_3(s) + 2\text{Al}(s) \rightarrow 2\text{Fe}(s) + \text{Al}_2\text{O}_3(s) \). What mole ratio would you use to calculate moles Fe possible if mol \( \text{Fe}_2\text{O}_3 \) is known?

A) 2 mol Al / 1 mol Fe\(_2\)O\(_3\)  
B) 2 mol Fe / 1 mol Fe\(_2\)O\(_3\)  
C) 2 mol Fe / 2 mol Fe\(_2\)O\(_3\)  
D) 1 mol Fe\(_2\)O\(_3\) / 2 mol Fe

2. One component of pearls is calcium carbonate. If pearls are put into acid, they dissolve according to the reaction: \( \text{CaCO}_3(s) + 2\text{HCl}(aq) \rightarrow \text{CaCl}_2(aq) + \text{H}_2\text{O}(l) + \text{CO}_2(g) \). How many mol \( \text{CaCO}_3 \) can be dissolved with 0.0250 mol HCl?

A) 0.033 mol \( \text{CaCO}_3 \)  
B) 1.0 mol \( \text{CaCO}_3 \)  
C) 0.0250 mol \( \text{CaCO}_3 \)  
D) 0.0125 mol \( \text{CaCO}_3 \)

3. Sulfuric acid reacts with zinc metal to produce zinc sulfate and hydrogen gas according to the reaction: \( \text{H}_2\text{SO}_4 + \text{Zn} \rightarrow \text{ZnSO}_4 + \text{H}_2 \). Calculate how many moles of hydrogen can be made if 15.0 grams of zinc react with an excess of sulfuric acid.

A) 0.203 mol \( \text{H}_2 \)  
B) 0.229 mol \( \text{H}_2 \)  
C) 0.50 mol \( \text{H}_2 \)  
D) 0.65 mol \( \text{H}_2 \)

4. An alkaline battery produces electrical energy according to the following balanced equation: \( \text{Zn}(s) + 2\text{MnO}_2(s) + \text{H}_2\text{O}(l) \rightarrow \text{Zn(OH)}_2(s) + \text{Mn}_2\text{O}_3(s) \). Determine the limiting reactant if 25.0 grams of Zn and 30.0 grams of \( \text{MnO}_2 \) are used. Assume that an excess of water is available to react.

A) Zn  
B) \( \text{MnO}_2 \)  
C) \( \text{Zn(OH)}_2 \)  
D) \( \text{Mn}_2\text{O}_3 \)

5. An alkaline battery produces electrical energy according to the following balanced equation: \( \text{Zn}(s) + 2\text{MnO}_2(s) + \text{H}_2\text{O}(l) \rightarrow \text{Zn(OH)}_2(s) + \text{Mn}_2\text{O}_3(s) \). Determine the mass of \( \text{Zn(OH)}_2 \) produced (the theoretical yield).

A) 14.8 g \( \text{Zn(OH)}_2 \)  
B) 16.5 g \( \text{Zn(OH)}_2 \)  
C) 17.1 g \( \text{Zn(OH)}_2 \)  
D) 20.3 g \( \text{Zn(OH)}_2 \)

6. An alkaline battery produces electrical energy according to the following balanced equation: \( \text{Zn}(s) + 2\text{MnO}_2(s) + \text{H}_2\text{O}(l) \rightarrow \text{Zn(OH)}_2(s) + \text{Mn}_2\text{O}_3(s) \). Once you determine the theoretical yield of \( \text{Zn(OH)}_2 \) produced, determine the percent yield if a chemist is able to produce 15.6 grams of \( \text{Zn(OH)}_2 \) in an actual experiment.

A) 65%  
B) 78%  
C) 87%  
D) 91%