Essential Question:
1. Does the size of a gas molecule affect the amount of pressure it exerts on a container?
2. Does the ratio of gases in a mixture affect the pressure of the system?

Hypothesis:
Devise a hypothesis that indicates the correlation be size of gas molecules and pressure.

Procedure:
2. Click “Download.” Choose to “keep the file” when prompted. Open.
3. Click volume in constant parameters. Click on measurement tools → stopwatch. In the advanced options, you want to be sure that molecules collide and the temperature is set at 300K.

Part 1: Testing pressure of “Heavy” Gas
4. To establish a control, test the heavy species. Click on heavy species and enter 100 into the Gas Chamber.
5. Click the start button on the stopwatch. It will not start until you press the play button to the left of stopwatch.
6. Allow the gas molecules to move about the container for about 60 seconds. Then, press the pause button.
7. Record the pressure of the system as trial 1 for 100 particles of heavy gas.
8. Click “reset” before starting the next set of trials.
9. Repeat steps 5-9 for two more trials. Take an average of the trials.
10. To test varying amounts of gas, repeat steps 5-10 using the following amounts of gas 75, 50, 25 particles.
Part 2: Testing the “Lighter” Gas
1. To establish a control, test the light species. Click on light species and enter 100 into the Gas Chamber.
2. Click the start button on the stopwatch. It will not start until you press the play button to the left of stopwatch.
3. Allow the gas molecules to move about the container for about 60 seconds. Then, press the pause button.
4. Record the pressure of the system as trial 1 for 100 particles of light gas.
5. Repeat steps 5-8 for two more trials. Take an average of the trials.
6. To test varying amounts of gas, repeat steps 5-9 using the following amounts of gas 75, 50, 25 particles.

Part 3: Testing a Varying Gas Mixture
1. Click heavy species and enter 75 into the Gas Chamber. Then, enter 25 for the light species.
2. Click the start button on the stopwatch. It will not start until you press the play button to the left of stopwatch.
3. Allow the gas molecules to move about the container for about 60 seconds. Then, press the pause button.
4. Record the pressure of the system as trial 1 for 75-25 gas mixture.
5. Repeat steps 5-8 for two more trials. Take an average of the trials.
6. To test varying mixtures of gas, repeat steps 5-9 using the following amounts of gas 50-50 and 25-75 particles.

Data:

<table>
<thead>
<tr>
<th>Part 1: Testing pressure of “Heavy” Gas</th>
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<tbody>
<tr>
<td>Heavy Gas Particles</td>
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<tr>
<td>---------------------</td>
</tr>
<tr>
<td>100</td>
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<tr>
<td>75</td>
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<tr>
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<table>
<thead>
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<th>Part 2: Testing the “Lighter” Gas</th>
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<td>Light Particles</td>
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<tr>
<td>100</td>
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<tr>
<td>75</td>
</tr>
<tr>
<td>50</td>
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<tr>
<td>25</td>
</tr>
</tbody>
</table>
Part 3: Testing a Varying Gas Mixture

<table>
<thead>
<tr>
<th>Mixture H:L</th>
<th>Trial 1 Pressure (atm)</th>
<th>Trial 2 Pressure (atm)</th>
<th>Trial 3 Pressure (atm)</th>
<th>Average (atm)</th>
</tr>
</thead>
<tbody>
<tr>
<td>75:25</td>
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<tr>
<td>50:50</td>
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<tr>
<td>25:75</td>
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</tr>
</tbody>
</table>

Analysis:

1. Did you prove or disprove your hypothesis relating the size of a gas molecule to the pressure it exerts? Explain.

2. Add the pressure of 75 heavy particles to 25 light particles. How does it relate to the 75:25 mixture pressure?

3. Add the pressure of 50 heavy particles to 50 light particles. How does it relate to the 50:50 mixture pressure?

4. Add the pressure of 25 heavy particles to 75 light particles. How does it relate to the 25:75 mixture pressure?

5. How does the ratio of gases in a mixture affect the pressure of the system?
Extension:

1. A container holds three gases: oxygen, carbon dioxide, and helium. The partial pressures of the three gases are 2.00 atm, 3.00 atm, and 4.00 atm, respectively. What is the total pressure inside the container?

2. A container with two gases, helium and argon, is 30.0% by volume helium. Calculate the partial pressure of helium and argon if the total pressure inside the container is 4.00 atm.

3. Hydrogen gas was collected over water. The total pressure of the container was 770.35 mmHg. If the water vapor pressure in the container is 55.30 mmHg, what is the partial pressure of hydrogen gas in the system?
Unit 9, Worksheet 2 —
Gases Again

1. A can of spray paint contains nitrogen gas as the propellant. The pressure of the gas is 3.5 atm when the temperature is 20°C. The can is left in the sun, and the temperature of the gas increases to 50°C. What is the pressure in the can?

2. A 90.0 mL volume of helium was collected under a pressure of 740 mmHg. At what volume would the pressure of this gas be 700 mm Hg? Assume temperature is constant.

3. A small bubble rises from the bottom of a lake, where the temperature is 8°C and the pressure is 6.4 atm, to the water’s surface, where the temperature is 25°C and pressure is 1.0 atm. Calculate the final volume (in mL) of the bubble if its initial volume was 2.1 mL.

4. Three gases are mixed in a 1.00 L container. The partial pressure of CO₂ is 250 mm Hg, N₂ is 375 mm Hg, and He is 125 mm Hg. What is the pressure of the mixture of gases?

5. What are the percentages, by moles, of the gases in the above mixture?
6. Our atmosphere is a mixture of gases (roughly 79% N₂, 20% O₂ and 1% Ar). What is the partial pressure (in mm Hg) of each gas at standard pressure?

7. A mixture of He and O₂ gases is used by deep-sea divers. If the pressure of the gas a diver inhales is 8.0 atm, what percent of the mixture should be O₂, if the partial pressure of O₂ is to be the same as what the diver would ordinarily breathe at sea level?

8. When you found the density of carbon dioxide gas, you collected the gas by displacing water in a bottle. The gas you collected was a mixture of CO₂ and H₂O vapor. If, on the day of the lab, the room pressure were 730 mm Hg and the partial pressure of water vapor were 21 mm Hg, what would be the partial pressure of the carbon dioxide gas? What fraction of the mixture was CO₂?

9. Suppose that when you reacted the zinc with the hydrochloric acid, you collected the hydrogen gas by water displacement. If the pressure in the room were 735 mm, and the partial pressure of the water were 22 mm Hg, what would be the partial pressure of the hydrogen gas? If the volume at this pressure were 25.0 mL, what would be the volume of the hydrogen gas alone at standard pressure?
Unit 9, Lab 1

Title:

Purpose/Question:

Procedure:
1. Record the room pressure and temperature. You will have to convert room pressure from inHg to mmHg (2.54 cm = 1 inch) and temperature from Celsius to Kelvin.
2. You should obtain a 50mL graduated cylinder, a 1000 mL beaker, a ring stand and clamp, distilled water, and a 3cm long piece of Mg metal.
3. Add 10 mL 3M HCl to the graduated cylinder.
4. Fill the rest of the graduated cylinder with distilled water.
5. Fill the 1000 mL beaker with tap water to the 800 mL mark.
6. Measure and record the length of your Mg metal (using significant figures).
7. Pull the copper wire through the hole in the rubber stopper. Create a “cage” so that the copper wire holds the Mg securely in place.
8. Push the stopper firmly in the graduated cylinder. Add additional water through the hole in the stopper until the entire graduated cylinder is filled with liquid.
9. Place your finger over the hole in the stopper and invert the graduated cylinder into the beaker of water. (The stopper end of the cylinder should be completely submerged.)
10. The inverted graduated cylinder can be clamped to the ring stand carefully (you do not want the stoppered end coming out of the water) so that you do not have to hold the cylinder.
11. Allow the reaction to proceed. (During this time, you will find a seat at the front of the room and begin your quiz.)
12. When the reaction is complete, you should adjust the height of the graduated cylinder in the beaker so that the water level in the cylinder is the same as in the beaker (do not allow the stoppered end of the cylinder to come out of the water).
13. Record the volume of the gas in the cylinder (use the scale where the “top” starts at zero).
14. Discard the contents of the tube in the sink. Bring your copper wire to the materials station to be reused.
15. Wash all glassware and wash your hands!
### Data

### Calculations

1. Convert the room pressure from inHg to mmHg. (2.54 cm = 1 inch)

   The hydrogen gas collected also contains some water vapor. The pressure of this mixture of gases equals the room pressure. Find the partial pressure of H₂ gas.

2. You collected H₂ gas at room temperature and at the pressure you found in #1. Determine what the volume of gas would be at STP.

3. Determine the mass, then the moles of Mg used.
4. Write a balanced equation to describe the reaction between the Mg and HCl. Determine the number of moles of H₂ gas in your sample.

5. Determine the volume of one mole of H₂ gas at STP.

6. Determine the % error in your experimental value.

7. Consider the model we have developed to represent molecules of a gas at room temperature. You may remember that the volume of a gas is roughly 1000x as great as that of a liquid. With this in mind, predict the volume of a mole of another gas, say N₂ or CO₂, at STP. Justify your prediction.

Your conclusion should include a discussion of 6 and 7 above. You should discuss (not just state in one sentence) at least 2 errors that could account for the % error above). Relate the conclusion back to the purpose. What is the importance of knowing the volume of a mole of gas?
Development of the Ideal Gas Law
1. What volume does 16.0 g of O2 occupy at STP?

2. A mixture contains 5.00 g each of O2, N2, CO2, and Ne gas. Calculate the volume of this mixture at STP.

3. A 250 mL flask of hydrogen gas is collected at 763 mm and 35°C by displacement of water from the flask. The vapor pressure of water at 35°C is 42.2 mmHg. How many moles of hydrogen gas are in the flask?

4. When calcium carbonate is heated strongly, carbon dioxide gas is evolved.

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

   If 4.74 g of calcium carbonate is heated, what volume of CO2 (g) would be produced when collected at STP?

   Equation:
   
   Before:
   
   Change
   
   After
5. Zinc metal reacts vigorously with chlorine gas to form zinc chloride. What volume of chlorine gas at 25˚C and 1.00 atm is required to react completely with 1.13 g of zinc?

Equation:

Before:

Change

After

6. Consider the following reaction:

\[ P_4(s) + 6 H_2(g) \rightarrow 4 PH_3(g) \]

What mass of \( P_4 \) will completely react with 2.50 L of hydrogen gas, at 0˚C and 1.50 atm pressure?

Equation:

Before:

Change

After

7. If water is added to magnesium nitride, ammonia gas is produced when the mixture is heated.

\[ Mg_3N_2(s) + 3H_2O(l) \rightarrow 3 MgO + 2NH_3(g) \]

If 10.3 g of magnesium nitride is treated with water, what volume of ammonia gas would be collected at 20˚C and 0.989 atm?

Equation:

Before:

Change

After

8. Nitrogen gas and hydrogen gas combine to produce ammonia gas (NH\(_3\)). What volume of hydrogen gas at 25˚C and 735 mm is required for the complete reaction of 10.0g of nitrogen?

Equation:

Before:

Change

After
1. Lithium reacts with water in a single replacement reaction forming hydrogen gas and lithium oxide. How many grams of lithium must be added to water in order to release 15.0 L of hydrogen gas at STP?

2. Potassium chlorate when heated strongly decomposes into potassium chloride and oxygen gas. 23.0 L of oxygen gas was released at 85°C and 100.5 kPa. What volume would the oxygen gas have at STP?
3. 224.5 grams of argon gas is contained inside a 15.0 L steel cylinder at a temperature of 24°C. What is the pressure inside the steel cylinder (in atm)?

4. When magnesium reacts with hydrochloric acid (HCl), hydrogen gas and magnesium chloride is produced. How many liters of hydrogen gas will be produced when 85.8 grams of pure magnesium metal reacts completely with excess hydrochloric acid. The laboratory temperature during this procedure was 24°C and the pressure was recorded at 765 mm Hg.
Unit 9, Worksheet 5—
Molarity of Solutions

Show your calculations and tell how you would prepare each of the following solutions in the laboratory. Study the first problem that has been completed as an example before beginning to work.

1. 2.0 L of 1.5 M potassium chloride

\[
2.0 \text{ L} \times \frac{1.5 \text{ mol KCl}}{1.0 \text{ L}} \times \frac{74.55 \text{ g}}{1 \text{ mol KCl}} = 220 \text{ g KCl dissolved in 2.0 L of solution}
\]

2. 3.0 L of 1.5 M lithium bromide

3. 0.500 L of 1.0 M aluminum nitrate

4. 1.0 L of 0.020 M magnesium hydroxide

5. 250.0 mL of 0.10 M beryllium nitrate

6. 500.0 mL of 1.50 M ammonium carbonate

7. 40.0 mL of 0.25 M rubidium fluoride
8. 100.0 mL or 0.80 M cesium nitrate

9. 500.0 mL of 6.0 M hydrochloric acid

10. A chemistry student obtained 75.0 mL of a 1.0 M solution of magnesium chloride. How many chloride ions are dissolved in this sample?

11. Draw a particle diagram that illustrates what is happening at the microscopic level for the magnesium chloride solution described in the previous problem #10. (HINT: Draw ions in the correct ratio with numbers that show the quantity of each ion in the solution.)

12. Draw two particle diagrams that illustrate (at the microscopic particle level) how a 1.0M solution of sodium bromide differs from a 2.0M solution of sodium bromide.
1. Solid magnesium bromide is dissolved in water producing 1.00 L of a 3.0 M solution.  
a. Write a balanced dissociation reaction for this reaction.

_________________________________________________________________

b. How many grams of solid magnesium bromide were mixed with water to make this solution?

c. Assuming 100% dissociation, how many Mg\(^{2+}\) ions are floating in this solution?

d. Assuming 100% dissociation, how many Br\(^{-}\) ions are in this solution?

2. A 45.3 g sample of potassium nitrate is dissolved in enough water to make 225 mL of solution. Determine the molar concentration of the potassium nitrate.

3. What is the molarity of a solution made by dissolving 275 g CuSO\(_4\) in enough water to make 4.25 L of solution?
4. An alcoholic iodine solution (“tincture” of iodine) is prepared by dissolving 5.15 g of iodine crystals in enough alcohol to make a volume of 225 mL. Calculate the molarity of iodine in the solution.

5. What final volume would be needed in order to prepare a 0.25 M NaCl solution from 5.2 g of NaCl (s)?

6. Draw a particle diagram of each of these ionic substances in solution. Then, calculate the molarity of each ion present in each of the following solutions.

a. 0.10 M AlCl$_3$

[Al$^{3+}$] = _________________________________

[Cl$^-$] = _________________________________

b. 0.40 M Na$_2$CrO$_4$

[Na$^+$] = _________________________________

[CrO$_4^{2-}$] = _________________________________
Unit 9, Worksheet 7—
Molarity and Stoichiometry

1. Draw a particle diagram of each of these ionic substances in solution. Then, calculate the molarity of each ion present in each of the following solutions.
   a. 0.0030 M Ca(OH)₂

   [Ca²⁺] = _______________________________
   [OH⁻] = ______________________________

   b. 0.250 M Na₃PO₄

   [Na⁺] = ______________________________
   [PO₄³⁻] = ______________________________

2. How many grams of silver nitrate are needed to prepare 750 mL of standard 0.200 M silver nitrate solution?

3. A chemistry student has 10.0 g of AgNO₃ and wants to prepare as much 0.10 M AgNO₃ solution as possible. How many milliliters of solution can be created?
4. Concentrated hydrochloric acid is made by pumping hydrogen chloride gas into distilled water. The gas dissolves in the water creating H+ and Cl- ions. The most concentrated hydrochloric acid solution possible contains 439 grams HCl for each one liter of solution. Adding more HCl gas to this solution does not increase its concentration because the gas escapes into the air. What is the molarity of concentrated hydrochloric acid?

5. How many moles of solid lead (II) hydroxide can be formed when 0.250L of 0.450 M Pb(NO₃)₂ solution reacts with excess aqueous sodium hydroxide?

EQUATION: _________________________________________________________________

6. Barium nitrate reacts with aqueous sodium sulfate to produce solid barium sulfate and aqueous sodium nitrate. Abigail places 120.00 mL of 0.600 M barium nitrate in a flask. She has a bottle of 0.300M sodium sulfate solution available. What volume of sodium sulfate solution must she add to her flask of barium nitrate so she has no excess reactant left over?

EQUATION: _________________________________________________________________

7. Aqueous calcium chloride reacts with aqueous sodium carbonate forming solid calcium carbonate and aqueous sodium chloride. 350.0 mL of 0.100M CaCl₂ is mixed with 350.0 mL of 0.100M Na₂CO₃. What is the theoretical yield of solid calcium carbonate (in grams).

EQUATION: _________________________________________________________________
Energy of Combustion Lab

Introduction
In this experiment you will measure how much energy is released in the combustion reaction that occurs when candle wax and oxygen react to form carbon dioxide and water. The amount of energy released depends on the amount of wax burned. To compare the results from different lab teams, you will determine the quantity of heat released for each mole of the candle wax that was consumed.

Question/Purpose

Procedures

Data
Analysis

1. Determine the mass of candle wax burned.

2. Candle wax is a mixture of compounds. Assume that the formula of the wax you used is $C_{25}H_{52}$. Determine the number of moles of candle wax that were burned.

3. Determine the mass of water heated. Recall that the density of water is 1.0 g/mL.

4. Determine the temperature change ($\Delta t$) of the water.

5. Calculate the quantity of heat ($Q$) absorbed by the water in the can. $Q = mc\Delta t$
   Express the answer in kJ.

6. Calculate the quantity of heat that would be produced if one mole of candle wax were burned. Express your answer in kJ/mole.

Class Data:

<table>
<thead>
<tr>
<th>Lab team</th>
<th>Change in mass of candle wax consumed (g)</th>
<th>Moles of candle wax consumed</th>
<th>Q absorbed (kJ)</th>
<th>Heat of combustion (kJ/mol)</th>
</tr>
</thead>
</table>
Conclusions:

1. How did the value you obtained for the heat of combustion, ΔH, compare to that obtained by the other lab groups? The accepted value?

2. Write the balanced equation for the combustion of candle wax, including the quantity of energy measured in the experiment. Which stored more energy, reactants or products?

3. Had you used a more sophisticated calorimeter, would your answer be higher or lower? Explain why.

4. Discuss a minimum of two lab errors that contributed imperfect data.
Consider the reaction in which methane burns in air to produce carbon dioxide and water. The balanced equation for this reaction is

\[ \text{CH}_4 + 2 \text{O}_2 \rightarrow \text{CO}_2 + 2 \text{H}_2\text{O} \]

Experiments show that when one mole of \( \text{CH}_4 \) is burned, 890 kJ of energy are released.

It is possible to represent energy changes in chemical reactions in more than one way. Several meaningful representations are summarized below.

1. In Unit 6 an equation was written with the energy term listed as a product of the reaction as follows:

\[ \text{CH}_4 + 2 \text{O}_2 \rightarrow \text{CO}_2 + 2 \text{H}_2\text{O} + \text{energy} \]

Although this representation communicates that energy is released as a product of the reaction, it does not communicate how much energy is released. In Unit 8 students are able to write a chemical equation that includes the energy quantity. The more meaningful modified version of the chemical equation is shown below.

\[ \text{CH}_4 + 2 \text{O}_2 \rightarrow \text{CO}_2 + 2 \text{H}_2\text{O} + 890 \text{kJ} \]

2. Chemists routinely provide information related to energy transfers using what is called \( \Delta H \) notation. In this representation, the chemical reaction is written:

\[ \text{CH}_4 + 2 \text{O}_2 \rightarrow \text{CO}_2 + 2 \text{H}_2\text{O} \quad \Delta H = -890 \text{kJ} \]

The negative sign indicates that the chemical potential energy of the system decreases as energy is transferred from chemical potential energy to thermal energy. In other words, the products of the reaction have less chemical potential energy than was possessed by the reactants before the reaction began. Remember that the sign associated with \( \Delta H \) (+ or -) describes the net change in the chemical potential energy in a closed system.

In an open system, the reaction described above eventually result in a transfer of energy from system to surroundings. A negative value for \( \Delta H \) is a reminder that the total energy of the system decreases while the total energy of the surroundings increases while these transfers are taking place.

3. The energy bar graph for this chemical reaction provides additional insights into the energy changes that occur.
Note in this representation, the chemical potential energy of the system decreases as the reactant molecules rearrange to form product molecules. Initially, the thermal energy increases (system gets hotter) as energy is conserved. Eventually, the warmer system transfers energy to the cooler surroundings.

This reaction is described as **exo-thermic** because eventually energy **exits the system** and is transferred into the surroundings.

Experimental evidence has shown that the chemical reaction summarized below requires 72.8 kJ of energy to proceed.

\[
H_2 + Br_2 \rightarrow 2HBr
\]

DIRECTIONS: Represent this reaction in three different ways.

1. Write a balanced equation with the energy quantity in the equation.

2. Use $\Delta H$ notation to describe the energy changes that occur in the reaction.

3. Complete the energy bar graph diagram, showing how $E_{th}$ and $E_{ch}$ change in the system. Will energy eventually flow into or out of the system? Why?
1. How many kJ of heat will be released when 4.72 g of carbon react with excess oxygen gas to produce carbon dioxide?

\[ \text{C} + \text{O}_2 \rightarrow \text{CO}_2 \quad \Delta H^\circ = -393.5 \text{ kJ} \]

2. How much heat should be transferred when 38.2 g of liquid bromine reacts with excess hydrogen gas to form hydrogen bromide? Is the heat being transferred from the system to the surroundings or from the surroundings to the system?

\[ \text{H}_2 + \text{Br}_2 \rightarrow 2 \text{HBr} \quad \Delta H^\circ = 72.80 \text{ kJ} \]

3. How many kJ of heat would you expect to be transferred when 6.44 g of sulfur react with excess oxygen gas to produce sulfur trioxide? Is this reaction endothermic or exothermic?

\[ 2 \text{S} + 3 \text{O}_2 \rightarrow 2 \text{SO}_3 \quad \Delta H^\circ = -791.4 \text{ kJ} \]

4. Nitrogen gas and oxygen gas can combine to produce nitric oxide, NO. If such a reaction absorbs 88.0 kJ of heat from the surroundings, how many grams of nitrogen gas do you predict were consumed in the reaction?

\[ \text{N}_2 + \text{O}_2 \rightarrow 2 \text{NO} \quad \Delta H^\circ = 180 \text{ kJ} \]
5. Ammonia gas combines with excess oxygen gas to produce nitric oxide and water. If 256 kJ of energy were released in such a reaction, how many grams of ammonia gas were reacted?

\[ 4 \text{NH}_3 + 5 \text{O}_2 \rightarrow 4 \text{NO} + 6 \text{H}_2\text{O} \quad \Delta H^\circ = -1170 \text{ kJ} \]

6. Carbon in the form of graphite combines with excess hydrogen gas to form benzene, C\textsubscript{6}H\textsubscript{6}. In the following reaction 3.95 kJ of heat were transferred. Calculate the grams of graphite reacted. Is the reaction endothermic or exothermic?

\[ 6\text{C (graphite)} + 3 \text{H}_2 \rightarrow \text{C}_6\text{H}_6 \quad \Delta H^\circ = 49.03 \text{ kJ} \]

7. How much heat will be released if 30.0 g of octane (C\textsubscript{8}H\textsubscript{18}) is burned in excess oxygen?

\[ \text{C}_8\text{H}_{18} + 12 \frac{1}{2} \text{O}_2 \rightarrow 8 \text{CO}_2 + 9 \text{H}_2\text{O} \quad \Delta H = -5483.4 \text{ kJ} \]

How much heat would be released by burning one gallon of octane? The density of octane is 0.703 g/mL. 1 gallon = 3.79 liters.
Unit 9— Extra Practice Problems

1. A mixture of gases contains 82.00 grams each of NO\textsubscript{2}, CO\textsubscript{2}, and SO\textsubscript{2} at a pressure of 1.08 atmospheres and at a temperature of 15°C.

   a. Assuming that the mixture of gases occupies a container that is capable of expansion and contraction, calculate the total volume of the gas mixture in liters.

   b. Using < (less than) symbols, rank the three gases in order of their partial pressures.

   __________________________________________________________________________

   Is it possible that all three gases exert equal partial pressures? __________ Why or why not?

   c. What would have been the total volume of this same mixture of gases if the gases had been at STP?

2. Nitrogen gas and hydrogen gas combine to produce ammonia gas (NH\textsubscript{3}). What volume of hydrogen gas at STP is required to react completely with 1.50 moles of nitrogen gas?
3. Consider the following reaction: \[ \text{P}_4(\text{s}) + 6 \text{H}_2(\text{g}) \rightarrow 4 \text{PH}_3(\text{g}) \]
   a. What volume of hydrogen gas at laboratory conditions of 25.0°C and 0.981 atm is required to react completely with 42.85 grams of \( \text{P}_4 \)?

   b. What volume of \( \text{PH}_3 \) gas will be produced in this reaction assuming that the reaction provides the normal and expected percent yield of 82.0%?

4. Solid tin metal reacts with chlorine gas forming liquid tin (IV) chloride according to the reaction shown below:
   \[ \text{Sn} (\text{s}) + 2\text{Cl}_2(\text{g}) \rightarrow \text{SnCl}_4(\text{l}) \quad \Delta \text{H} = -511 \text{ kJ} \]

   What volume of chlorine gas (at laboratory conditions of 760.0 mm Hg and 25°C) is required in order for the reaction to release 328 kJ of heat energy?

5. What mass of \( \text{Na}_2\text{CO}_3 \) is needed to prepare 1.25 L of a 0.75M solution?
6. Sucrose (C₁₂H₂₂O₁₁) solution is produced by dissolving table sugar in distilled water. Show the math and explain in words step-by-step how to mix 2.0 liters of a 0.200M sucrose solution. (NOTE: The laboratory only stocks half-liter and one liter volumetric flasks.)

7. What is the molarity of a solution that contains 125 grams of NaCl in 4.00 L of solution?

10. Determine the molarity of a solution containing 4.67 moles of lithium sulfite dissolved in 2.04 liters of solution.

11. What is the molarity of a solution containing 0.0348 grams of lead (II) chloride dissolved in 45.0 mL of solution?

12. How many grams of magnesium carbonate are needed to make 3.00 L of a 0.500 M solution?

13. Which solution would contain the greater mass of solute: 6.2 L of a 3.76 M sodium oxide solution or 5.4 L of a 4.75 M sodium oxide solution? Show your work and give a written explanation, based on data, for your choice.
14. Solid iron (III) chloride can be produced by reacting aqueous iron (III) oxide with a hydrochloric acid solution. How many milliliters of a 6.00M HCl solution are needed to react with excess Fe₂O₃ to produce 16.5 grams of solid FeCl₃? (NOTE: The second product is liquid water.)

EQUATION: _______________________________________________________________

B

C

A

Ideal Gas Law Practice Problems:

1) If I have 4 moles of a gas at a pressure of 5.6 atm and a volume of 12 liters, what is the temperature?

2) If I have an unknown quantity of gas at a pressure of 1.2 atm, a volume of 31 liters, and a temperature of 87 °C, how many moles of gas do I have?

3) If I contain 3 moles of gas in a container with a volume of 60 liters and at a temperature of 400 K, what is the pressure inside the container?

4) If I have 7.7 moles of gas at a pressure of 0.09 atm and at a temperature of 56 °C, what is the volume of the container that the gas is in?

5) If I have 17 moles of gas at a temperature of 67 °C, and a volume of 88.89 liters, what is the pressure of the gas?
6) If I have an unknown quantity of gas at a pressure of 0.5 atm, a volume of 25 liters, and a temperature of 300 K, how many moles of gas do I have?

7) If I have 21 moles of gas held at a pressure of 78 atm and a temperature of 900 K, what is the volume of the gas?

8) If I have 1.9 moles of gas held at a pressure of 5 atm and in a container with a volume of 50 liters, what is the temperature of the gas?

9) If I have 2.4 moles of gas held at a temperature of 97 °C and in a container with a volume of 45 liters, what is the pressure of the gas?

10) If I have an unknown quantity of gas held at a temperature of 1195 K in a container with a volume of 25 liters and a pressure of 560 atm, how many moles of gas do I have?

11) If I have 0.275 moles of gas at a temperature of 75 K and a pressure of 1.75 atmospheres, what is the volume of the gas?

12) If I have 72 liters of gas held at a pressure of 3.4 atm and a temperature of 225 K, how many moles of gas do I have?
Ideal Gas Law and Stoichiometry Problems:

1) For the reaction \(2 \text{H}_2(g) + \text{O}_2(g) \rightarrow 2 \text{H}_2\text{O}(g)\), how many liters of water can be made from 5 L of oxygen gas and an excess of hydrogen?

2) How many liters of water can be made from 55 grams of oxygen gas and an excess of hydrogen at STP?

3) How many liters of water can be made from 55 grams of oxygen gas and an excess of hydrogen at a pressure of 12.4 atm and a temperature of 85° C?

4) How many liters of water can be made from 34 grams of oxygen gas and 6.0 grams of hydrogen gas at STP? What is the limiting reactant for this reaction?
Molarity Practice Problems:

1) How many grams of potassium carbonate are needed to make 200 mL of a 2.5 M solution?

2) How many liters of 4 M solution can be made using 100 grams of lithium bromide?

3) What is the concentration of a 450 mL solution that contains 200 grams of iron (II) chloride?

4) How many grams of ammonium sulfate are needed to make a 0.25 M solution at a concentration of 6 M?

5) What is the concentration of a solution that has a volume of 2.5 L and contains 660 grams of calcium phosphate?

6) How many grams of copper (II) fluoride are needed to make 6.7 liters of a 1.2 M solution?

7) How many liters of 0.88 M solution can be made with 25.5 grams of lithium fluoride?
8) What is the concentration of a solution that with a volume of 660 that contains 33.4 grams of aluminum acetate?

9) How many liters of 0.75 M solution can be made using 75 grams of lead (II) oxide?

10) How many grams of manganese (IV) oxide are needed to make a 5.6 liters of a 2.1 M solution?

11) What is the concentration of a solution with a volume of 9 mL that contains 2 grams of iron (III) hydroxide?

12) How many liters of 3.4 M solution can be made using 78 grams of isopropanol (C₃H₇O)?

13) What is the concentration of a solution with a volume of 3.3 mL that contains 12 grams of ammonium sulfite?

**Molarity and Stoichiometry Problems:**

1) How many mL of a 0.124 M sodium hydroxide are needed to react with 15.4 mL of 0.108 M H₂SO₄?
2) You add 500 mL of a 0.100 M AgNO₃ solution containing an excess of sodium chloride solution. How much silver chloride precipitate will form?

3) If you mix 200 ml of 0.100 M lead (II) nitrate and 300 ml of 0.200 M magnesium chloride, how much lead (II) chloride precipitate will you form?

4) How many L of a 0.100 M hydrochloric acid solution would be required to react completely with 5.00 g of calcium hydroxide?

5) If I combined 15.0 g of calcium hydroxide with 75.0 mL of a 0.500 M hydrochloric acid solution, how many grams of calcium chloride would be formed?
6) How many litres of 0.100 M HCl would be required to react completely with 5.00 grams of calcium hydroxide?

7) How many grams of calcium hydroxide are needed to react with 69.50 ml of 0.350 M hydrochloric acid?

8) If 350.0 mL of 0.250 M calcium chloride were produced, what mass of calcium hydroxide was required? (Assume an excess of HCl)