

POST-LAB QUESTIONS

1. A student measured a book's cover and recorded the length as 28.1 cm and the width as 22.7 cm. The area, using a calculator, was reported to be 637.87 cm². What should have been reported by the student? Explain your answer.
2. A man weighed 152 lbs. Express the quantity in kilograms. Show your work.
3. Pike's Peak in Colorado is the 32nd highest peak in the world at 14,110 ft. in height. Express this height in meters (m) and kilometers (km). Show your work.
4. A 3.456 g sample of iron filings was added to a beaker weighing 83.65 g. What is the combined weight of the ~~paper~~^{beaker} and sample in grams (g) and milligrams (mg). Show your work.
5. Two students each weighed a 125-ml Erlenmeyer flask that had a true weight of 80.562 g. Each student recorded three weights for the flask and took an average. Below are the results.

Which set of results is more accurate? _____

Which set of results is more precise? _____

| Student A | Student B |
|---------------|---------------|
| 80.560 | 80.400 |
| 80.555 | 79.551 |
| <u>80.565</u> | <u>80.793</u> |

Average

What can be said of the ~~results from~~^{about the accuracy and precision for} Student A and Student B?

Experiment 3

POST-LAB QUESTIONS

1. All the metals listed in Table 3.1 can have their densities determined by the 'displacement-of-a-liquid' method, as in the **Procedure**. Why can you use this method?

2. Olive oil (density 0.918 g/cm^3) separates from the water-vinegar in a salad dressing and is visible as a layer. Is the olive oil the top or the bottom layer? Explain your answer.

Ethyl alcohol has a density of 0.791 g/cm^3 at 20°C . How many milliliters (mL) are needed to have 30.0 g of liquid? Show your work.

1. A student uses a volumetric pipet calibrated to deliver exactly 50.0 mL of a liquid and then measures the mass of the liquid. How is the density determination of the liquid affected in the situations which follow? Would the density increase, decrease, or not be affected? Explain your answer.

a. A dirty pipet is used and water droplets adhere to the inner walls of the pipet.

b. The student did not allow sufficient time for all the liquid to empty from the pipet.

c. The student blew out all the liquid from the pipet, including the small amount from the tip.

d. The student did not remove air bubbles from the pipet before delivering the liquid.

e. At the mark on the pipet, the student read the upper edge of the meniscus, not the lowest point on the curve.

5. Iron should sink in water since its density is greater than that of water (see Table 3.1). However, ships (for example, tankers and cruise ships) float even though their hulls are built of steel, an iron alloy. Explain why this is possible.

POST-LAB QUESTIONS

1. The weight of naphthalene was obtained by weighing the collected solid. Was this an accurate weight of the naphthalene present in the original sample? Explain your answer.
2. In step 14 the sand was not completely free of water when the weight was reported. How does this affect the percent composition of the sand in the mixture?
3. In your experiment did you recover 100% of the original mixture? What was your percent recovery?
4. Snow in the winter can slowly disappear even when the temperature is freezing. How do you account for this observation?
5. A sample of French fried potatoes weighing 150 g was extracted with the volatile organic solvent hexane. The recovered cooking oil weighed 12.3 g. What methods from this experiment could be used to separate and isolate the cooking oil after the extraction? What was the percent oil in the potatoes? Show your calculations.

POST-LAB QUESTIONS

1. Write the balanced equation for the formation of iron(II) oxide, FeO, from the elements of iron, Fe, and oxygen, O₂.
2. Write the two chemical equations that describe the conversion of magnesium into magnesium oxide.
3. What error in calculation would result if, in the procedure for forming the magnesium oxide, the fumes in the initial heating were allowed to escape?
4. Mercuric oxide decomposes when heated and forms free mercury metal and releases oxygen gas. If a sample of the oxide, 2.096 g, is decomposed, 1.942 g of mercury is obtained. Determine the empirical formula for mercuric oxide. (*Hint*: first determine the number of moles of mercury and oxygen in the sample.)

Experiment 8

POST-LAB QUESTIONS

Zinc metal, Zn, reacts with 0.5 M HCl, but mercury metal, Hg, does not. Why?

From the following list of chemicals, select two combinations that would lead to a double replacement reaction. Write the complete, balanced equations for the reactions of the chemicals in solution.



When ammonium carbonate [(NH₄)₂CO₃] decomposes, two gases are given off. Name the gases. Suggest a test to identify each one and the observation for the test.

POST-LAB QUESTIONS

1. When do real gases approximate the behavior of an ideal gas? Do the conditions for this experiment meet these criteria? Explain.
2. Instead of measuring the volume of the Erlenmeyer flask as directed in step no. 15, a student recorded the volume written on the flask, in this case 125 mL. Recalculate the molar mass using this value for the volume and determine the percent error that results.
3. How will the calculated molar mass be affected if all of the liquid in the flask did not completely vaporize?
4. A student took the atmospheric pressure to be 1 atm instead of recording the actual barometric pressure (which in the student's case was 750 mm Hg). Will the calculated molar mass be too high, too low, or show no change?
5. A student had an unknown liquid and wished to determine its molecular weight. The following data was collected using the procedure of this experiment:
 - a. Weight of the 125-mL Erlenmeyer flask and cover (aluminum foil and copper wire), 79.621 g;
 - b. Weight of cooled flask and cover containing condensed liquid, 79.983 g;
 - c. Temperature of boiling water, 100.0°C;
 - d. Volume of the 125-mL Erlenmeyer flask, 128.0 mL;
 - e. Atmospheric pressure, 760.0 mm Hg.

Using this data, calculate the liquid's molar mass and identify the unknown from Table 14.1. Show all of your work.

Experiment 16

POST-LAB QUESTIONS

1. On the basis of your observations, are HCl solutions electrolytes or nonelectrolytes? Explain your answer.
2. Both acetic acid and HCl solutions react as acids. However, only the HCl solutions caused the light bulb to shine brightly while the acetic acid solutions caused the light bulb to shine dimly. How do you account for these observations?
3. Why does naphthalene dissolve in petroleum ether (an organic solvent similar to gasoline) but not in water?
4. Distilled water shows no conductivity since the light bulb does not go on and thus, is a nonelectrolyte. However, tap water does show some conductivity since the light goes on faintly. Explain this result.

POST-LAB QUESTIONS

1. If, in the reaction between magnesium and HCl, we would have used 3 M HCl rather than 6 M HCl, would the rate of reaction increase, decrease, or remain the same?
2. If in the reaction between 6 M HCl and Mg had we used globular chunks of magnesium instead of ribbon (both having the same weight), would the rate of reaction increase, decrease, or stay the same? Explain.
3. Assume that the zinc ribbon you added to the test tubes in Section 5 was 0.5 g. Calculate the rate of reaction as moles of Zn per min. for each temperature.
4. In general, we expect that doubling the concentration of a reactant will approximately double the rate of reaction. Was this expectation justified in the iodine clock reaction?
5. Assume that we do a reaction of zinc with 6 M HCl at room temperature (20°C). How much ~~faster will~~ ^{time will it take} these two chemicals ^{to} react at 40°C (see Section 5 of the **Procedure**)?

Experiment 20

POST-LAB QUESTIONS

1. In the first experiment with the copper–ammonia complex, you added ammonia to change the color and, later, equal strength HCl to change it back to blue. Did you require more, less, or an equal number of drops from each to accomplish the color change? On the basis of stoichiometry, what was your expectation?
2. Adding HCl to the $\text{H}_2\text{PO}_4^-/\text{HPO}_4^{2-}$ mixture is called a common ion effect. Explain why.
3. Silver ion reacts with chloride ion in solution to form the precipitate AgCl. What would happen to the color of a dilute solution containing FeCl_4^- if we added to it a solution containing silver ion?
4. The manufacture of ammonia from nitrogen and hydrogen is an exothermic reaction. Which temperature would give a greater yield of ammonia: room temperature or 100°C ? Explain.

Experiment 22

POST-LAB QUESTIONS

1. Assume that your vinegar contained a small amount of citric acid (a triprotic acid). Using the same experimental data, would you expect the molarity of this sample to be the same as or different than a sample which contained only pure acetic acid?
2. Assume that the tip of your buret was not properly filled with NaOH solution. It contained an air bubble which was eliminated during the titration. Would the calculated molarity of the vinegar be smaller, larger, or the same as the true molarity? Explain.
3. We added about 10 mL of distilled water to the vinegar, but we did not use this volume in the calculation of molarity. Why do you think we can ignore this volume in the calculation?
4. According to your results, how many grams of acetic acid are in a 250-mL bottle of vinegar?
5. You have a 10.0-mL sample of formic acid, HCOOH , to be titrated with 0.200 M NaOH. The initial reading of the volume of NaOH in the buret was 23.1 mL. The reading of the buret at the end point of the titration was 41.8 mL NaOH. What is the % w/v of formic acid in your sample? Show your work.
6. You have a 5% formic acid solution and a 5% acetic acid solution (vinegar). You titrate 5 mL of each with 0.1868 M NaOH. (a) Which will need more NaOH to reach the end point? (b) How much more? Show your calculation.