Student Name

Professor's Name

Course

Date

Report for the Limiting Reactant Lab

Complete the following tables:

Table 1: Conservation of Mass

Reaction / Test tube	Mass of test tube and 5% CH_3COOH (grams) (A)	Mass of $NaHCO_3$ (grams) (B)	Mass of balloon and $NaHCO_3$ (grams) (C)	Total mass before the reaction (grams) (D = A + C)
1	20.25	0.3	1.7	21.95
2	20.25	0.3	1.7	21.95
3	20.25	0.7	2.1	22.35
4	20.25	0.7	2.1	22.35

Show calculation of total mass for one reaction below:

Total mass before the reaction (grams) (D = A + C) for test tube #1 = 20.25 + 1.7 = 21.95 g

Table 2: Moles of acetic acid (CH_3COOH)

Reaction / Test tube	Volume of vinegar (mL)	Mass of CH_3COOH (grams)	Moles of CH_3COOH
1	5.0	0.25	0.00416
2	5.0	0.25	0.00416

3	5.0	0.25	0.00416
4	5.0	0.25	0.00416

Show calculation of mass and moles of CH_3COOH for one reaction below:

Volume of CH_3COOH= 5%*5.0 mL = 0.25 mL

Mass of CH_3COOH = density*volume = 1 g/mL * 0.25 mL = 0.25 g

Moles of CH_3COOH = mass/molar mass of acetic acid = 0.25 g/60.052 g/mol= 0.00416 moles

Table 3: Moles of NaHCO3

Reaction / Test tube	Mass of $NaHCO_3$ (grams) (B)	Moles of $NaHCO_3$
1	0.3	0.00357
2	0.3	0.00357
3	0.7	0.00833
4	0.7	0.00833

Show calculation of moles of $NaHCO_3$ for one reaction below:

Moles of $NaHCO_3$ = mass/molar mass = 0.3 g/ 84.007 g/mol = 0.00357 moles

Moles of $NaHCO_3$ = mass/molar mass = 0.7 g/ 84.007 g/mol = 0.00833 moles

Table 4: Theoretical Yield of CO_2

Reaction / Test tube	Limiting Reagent	Moles limiting reactant from Table 2 or 3	Moles CO_2 (use balanced equation)	Mass CO_2 (grams) Theoretical yield
1	$NaHCO_3$	0.00357	0.00357	0.1571
2	$NaHCO_3$	0.00357	0.00357	0.1571
3	CH_3COOH	0.00416	0.00416	0.1831
4	CH_3COOH	0.00416	0.00416	0.1831

Show calculation of moles and grams of CO_2 for one reaction below:

Moles of CO_2 = moles of limiting reagent = 0.00357 moles

Mass CO_2 (theoretical yield) = moles* molar mass = 0.00357*44.01 g/mol= 0.1571 grams

Table 5: Volume of Balloon

Reaction / Test tube	Diameter of balloon (cm)	Volume of balloon (cm^3) (H)	Calculations
1	5.52	396.30	Volume = ¾*π*5.52^3= 396.30 cm^3
2	5.60	413.79	Volume = ¾*π*5.60^3= 413.79 cm^3
3	6.70	708.66	Volume = ¾*π*6.70^3= 708.66 cm^3
4	6.75	724.64	Volume = ¾*π*6.75^3= 724.64 cm^3

Table 6: Total Volume of Carbon Dioxide

Reaction / Test tube	Volume of liquid in test tube after reaction (mL) (J)	Volume of test tube (mL) (I)	Volume of gas space (mL) K = I – J	Total Volume (mL) = H + K
1	2	9	7	c
2	3	9	6	413.79+6 = 419.79
3	3	9	6	708.66+6= 714.66
4	2	9	7	724.64+7= 731.64

Show calculation of volume of gas space and total volume for one reaction below:

Volume of gas space (mL) = 9-2 = 7 mL

Total Volume (mL) = Total Volume (mL)

Data Table 7: Calculation of Percent Yield

Reaction / Test tube	Actual moles of CO_2 obtained Use Total volume (H + K) from Table 6, in L	Actual yield of CO_2 in grams	Percent yield (%)
1	0.00303	0.1212	77.15
2	0.00329	0.1316	83.77
3	0.00380	0.1520	83.30
4	0.00357	0.1428	77.99

Show calculation of actual yield and percent yield for one reaction below:

Actual yield of CO_2 in grams = moles*molar mass = 0.00303*40.01= 0.1212 grams

Percent yield = actual yield/theoretical yield= 0.1212 g/0.1571 g = 77.15%

Data Table 8: Conservation of Mass

Reaction / Test tube	Total Mass after reaction (grams) (E)	Mass difference F = D - E
1	21.65	0.2
2	21.85	0.1
3	22.05	0.3
4	21.95	0.4

Show calculation of mass difference for one reaction below:

Mass difference = F = D – E= 22.35-0.3 = 22.05 grams

Answer questions:

1. The chemical reaction you investigated was a two-step reaction. What type of reaction occurred in each step? How did you determine your answer?

In the first step, a double replacement reaction occurred where two ionic compounds exchanged ions. In the second step, a decomposition reaction occurred; a single compound formed two products.

2. Was your percent yield of CO_2 100%? What factors might have caused any error you found? Explain, citing specific examples from what you saw during the experiment. How did each error cause a greater % yield or lesser % yield and why?

None of the reactions resulted in a percent yield of 100% due to experimental errors. The notable source of error is a measurement error, which could have occurred when sodium hydrogen carbonate and vinegar were measured. Some of the sodium hydrogen carbonate could have stuck to the balloon or the paper, thus less than the measured mass would have taken part in the reactions.

3. Do you think it is common for scientists to get 100% yields? Why or why not?

No, it is not common for scientists to get 100% yields, as chemical reactions are affected by errors. Some errors may be out of the scientists' control, while others can be controlled. However, errors always affect the accuracy of experimental results, making 100% yield a rare happening.

4. The Law of Conservation of Mass states that mass is neither created nor destroyed in chemical reactions. Explain why you may have had a mass difference in Data Table 7. Give specific examples from your experiment.

Each experiment proceeded independently of the other three experiments. Therefore, the results obtained for each were not affected by those obtained for the others. The measurement of the sodium hydrogen carbon mass and volume of vinegar could have been subjected to different errors that would have led to differing sets of results.

5. If you had to do this lab again, what would you do differently to improve your results? Explain, using specific examples.

I would ensure that the balloon and the paper were completely dry to minimize the amount of sodium hydrogen carbon that stuck to the balloon or paper, thus increasing the yield obtained at the end of the experiment. The amount of sodium hydrogen carbonate that sticks to the walls of the test tube and the balloon has a large effect on the results, since we are dealing with small volumes and masses.