

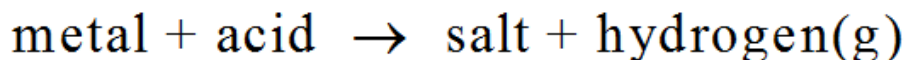
THE STOICHIOMETRY AND ATOMIC MASS OF ALUMINUM

Prior to the invention of the mass spectrogram by the British scientist, Francis William Aston in 1919, the atomic mass of an element was determined by stoichiometry. That is, measured masses of various substances were reacted. The atomic mass of the elements in the material was then deduced from the data obtained. The concepts of atoms and atomic “weight” were first introduced by another Englishman, John Dalton, in his atomic theory published in 1803. Dalton made three assertions which are the basis for much of the chemistry we learn today.

- (1) The chemical elements are composed of very minute individual particles of matter, called atoms, which preserve their individuality in all chemical changes.
- (2) All the atoms of the same element are identical in all respects, particularly in **weight**. Different elements have atoms differing in weight. Each element is characterized by the weight of its atom. (Dalton didn’t know about isotopes. Their existence wasn’t discovered until 1919 by F. W. Astor.)
- (3) Chemical composition occurs by the union of its atoms of the elements in simple numerical ratios.

As the quality of balances and the understanding the combining numbers improved, so did the accuracy of the atomic weights. In 1826, a Swedish chemist, John J. Berzelius, published a list of the atomic masses of fifty elements after more than ten years of work. While a number of atomic weights were published during that period, Berzelius’ values are the most consistent with the values given today. His atomic weight of aluminum changed from 54.88 grams in 1814, to 54.77 grams in 1818, and finally to 27.43 grams in 1826. Today’s value for the atomic mass of aluminum is 26.9815 grams.

The purpose of this experiment is to determine both the atomic mass of aluminum and the combining ratio between aluminum and hydrogen in much the same way as Berzelius did, with a gas measuring tube and a balance. This will be accomplished by studying the single replacement reaction that takes place when aluminum metal reacts with hydrochloric acid giving hydrogen gas as one of the products.



The gas generated will be collected by the downward displacement of water. This volume will be then converted into moles. Next, you will compare the ratio of the moles of hydrogen to aluminum in the balanced equation with the ratio determined experimentally. The atomic mass of aluminum will then be calculated from the mass of aluminum reacted and the number of moles of hydrogen collected.

For example, in a similar experiment, a student reacted a piece of aluminum foil with a mass of 0.041 g. When the reaction ceased, he recorded a volume of 54.0 ml of hydrogen at a water temperature of 21°C with an atmospheric pressure of 102.2 kPa. The student measured the volume of gas by displacing the water in a 60-ml syringe which required a correction of +1.2 ml. With only that information, he was able to determine both the stoichiometry of the reaction and the atomic mass of aluminum as follows:

1. The next step in the calculation is to determine the partial pressure of only the hydrogen gas in the mixture of hydrogen and water vapor collected. The pressure exerted by the hydrogen gas alone is then converted into atmospheres. The water temperature was 21°C and atmospheric pressure was 102.2 kPa. The vapor pressure of water is listed in the table to be 2.5 kPa at 21°C. The term “partial pressure” refers to the part of the total pressure exerted by each of the gases in a mixture. The relationship known as Dalton’s Law of Partial Pressures will be needed in this calculation. Where,

Total Pressure = Sum of the partial pressures of the individual gases in the mixture

$$\text{Total Pressure, } P_{\text{total}} = P_{\text{hydrogen}} + P_{\text{water vapor}}$$

$$102.2 \text{ kPa} = P_{\text{hydrogen}} + 2.5 \text{ kPa}$$

$$P_{\text{hydrogen}} = 102.2 \text{ kPa} - 2.5 \text{ kPa} = 99.7 \text{ kPa}$$

- 2.. The most important step in this calculation is to determine the number of moles of dry hydrogen gas generated by the reaction between aluminum metal and hydrochloric acid. This is accomplished using the ideal gas equation,

$$PV = nRT$$

Where, P = partial pressure of the hydrogen gas in kPa

V = volume of H₂ collected in the syringe in liters (volume of hydrogen in ml ÷ 1000 ml/liter)

n = number of moles of hydrogen in the gas measuring tube (the unknown here)

R = ideal gas constant, 8.31 J/deg·mol

T = the temperature of the gas in degrees Kelvin, K (°C + 273 = °K)

Using this student’s data,

P = the partial pressure of the hydrogen was calculated to be 99.7 kPa.

V = volume of hydrogen collected , 54.0 ml + the 1.2 ml correction for the syringe = 55.2 ml or 0.0552 liter (55.2 ml ÷ 1000 ml/liter)

T = temperature in degrees Kelvin, 21°C + 273 = 294 K

Substituting in the ideal gas equation,

$$PV = nRT$$

$$99.7 \text{ kPa} \times 0.0552 \text{ liter} = \# \text{ of moles of H}_2 \times 8.31 \frac{\text{J}}{\text{mol}\cdot\text{deg K}} \times 294 \text{ K}$$

$$\text{number of moles of H}_2 = 0.00225 \text{ mole H}_2$$

3. The student then calculated the number of moles of aluminum reacting from the mass of aluminum reacting calculated in step #1, 0.041 gram, and the atomic mass of aluminum, 27.0 grams/mol.

$$\text{number of moles of Al reacted} = \frac{\text{mass of Al reacted}}{\text{atomic mass Al}} = \frac{0.041 \text{ gram Al}}{27.0 \text{ g/mol}} = 0.00152 \text{ mol Al}$$

- 4.. Now that the number of moles of both hydrogen and aluminum are known, the ratio of the moles of hydrogen to moles of aluminum can be determined and compared with the expected ratio in the balanced equation. The empirical (experimental) ratio is then rounded to the nearest simple whole number ratio as expected in a balanced equation.

$$\text{Ratio} = \frac{\text{moles of hydrogen collected}}{\text{moles of aluminum consumed}} = \frac{0.00225 \text{ mol H}_2}{0.00152 \text{ mol Al}} =$$

$$\text{Ratio} = \frac{1.48 \text{ mol H}_2}{1 \text{ mol Al}} = \frac{3 \text{ mol H}_2}{2 \text{ mol Al}}$$

5. Let's calculate the percent error in our ratio. We expected a ratio of 3 moles of hydrogen to 2 moles of aluminum but calculated a ratio of only 1.48, a 0.02 difference. A 0.02 difference in a 1.50 ratio is a 1.3% error.

$$\text{percent error} = \frac{0.02}{1.50} \times 100 = 1.3 \text{ \% error}$$

7. The last calculation in this experiment will be to find the atomic mass of aluminum. The atomic mass of an element is the number of grams of that substance per mole of particles. Earlier, we calculated that when 0.41 grams of aluminum was reacted, it produced 0.00225 mol of hydrogen gas. Since we know that when 2 moles of aluminum produces 3 moles of hydrogen gas, we can use that ratio to find the number of moles of aluminum required to produce 0.00225 mol of hydrogen. Using this information, the atomic mass of aluminum was determined to be 27.3 g/mol.

$$0.00225 \text{ mol H}_2 \times \frac{2 \text{ mol Al}}{3 \text{ mol H}_2} = 0.00150 \text{ mol Al}$$

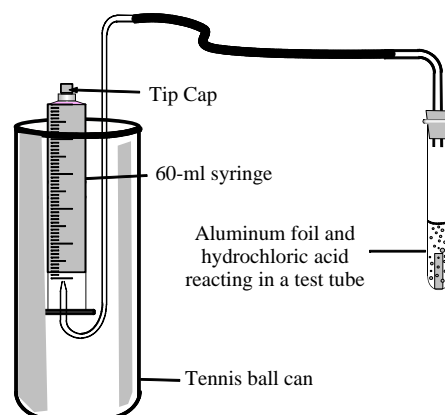
$$\text{atomic mass of Al} = \frac{0.041 \text{ gram Al}}{0.00150 \text{ mol Al}} = 27.3 \text{ g/mol}$$

Materials:

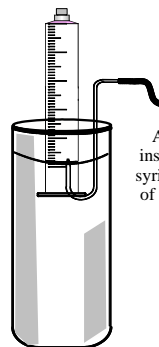
tennis ball can or tall beaker	thermometer	scissors or single edge razor blade
60-ml syringe	large (18 x 150 mm) test tube	small hypodermic syringe or calibrated transfer pipet
syringe tip cap	ruler	
250 ml beaker	4 M HCl	
gas generating apparatus	aluminum wire or foil	

Procedure:

1. Tear a 4 to 5 inch length of aluminum foil from the roll of aluminum foil. Using scissors, a utility knife, or a single edge razor blade, cut the foil into roughly 8 cm^2 pieces (2 cm x 4 cm).
2. Determine the mass of the piece of aluminum foil. If you have access to an analytical balance with a sensitivity of 0.001 gram or better, mass the metal directly. If your balance lacks milligram (0.001gram) or better sensitivity, you must calculate the mass of your sample. See the special calculation work sheet. The mass of your foil must **not exceed 0.42 gram.**
3. Fill the tennis ball can with tap water leaving a 1-2 cm space at the top. Then fill the syringe with water by pushing it below the surface of the water allowing the air to escape. Push a tip cap onto the open end to seal the syringe. Do not place the “J” tube into the water at this time.
4. Roll the foil into a loose coil and place it in a test tube followed by 3.0 ml of 4 M HCl. Press the stopper assembly firmly into the test tube with a twist to seal the system. Place the tip of the “J” tube into the bottom of the syringe. Check to be certain that all of the connections are tight to avoid a loss of hydrogen gas.
5. Swirl the test tube containing the acid and aluminum metal until the reaction begins to proceed rapidly. This may take 3-4 minutes depending on the acid temperature. Cool the reaction by placing the test tube into the water in the tennis ball can. Continue to swirl the test tube while it is in the water to cool the mixture and slowing the reaction. It may be necessary to hold the “J” tube against the syringe with one hand while swirling the test tube with the other. Wash any residual specks of un-reacted aluminum metal from the sides of the test tube by tilting the test tube slightly and then shaking the solution until it contacts the metal. Record your observations of the contents of the test tube.



6. When all the aluminum has been reacted, you must equalize the pressure inside the syringe and test tube with the atmosphere. With the tip of “J” tube still inside the syringe, move both the syringe and the “J” tube together until the open tip of the “J” tube is just above the level of the water in the tennis ball can. Now you can remove the “J” tube from the syringe. When recording the volume of the gas in the syringe, always move the syringe up or down until the level of the water inside is nearly the same as the level of the water in the tennis ball can. Record the volume of hydrogen gas to the nearest 0.5 ml. Be certain that you have read the volume of the hydrogen correctly. 60-ml syringes are calibrated in milliliters with the “0” at the top. Finally, record both the water temperature and the atmospheric pressure.



Always equalize the pressure inside the syringe by moving the syringe up or down until the level of the water inside the syringe is the same as the water level outside syringe the before removing the “J” tube.

7. Repeat the experiment as time permits.
8. Clean up by rinsing all glassware and returning the common materials to their designated areas.

Calculations and Questions

- Q1. (a) Determine the vapor pressure of water at the water temperature from a Table of the Vapor Pressures of Water at Various Temperatures.
- (b) Calculate the partial pressure of just the hydrogen in the gas measuring tube from Dalton's Law of Partial Pressures.
- Q2. (a) Correct the recorded volume of hydrogen in the syringe by adding 1.2 ml to your value.

$$\text{Corrected volume of hydrogen} = \text{recorded volume of hydrogen} + 1.2 \text{ ml}$$

- (b) Calculate the number of moles, n , of hydrogen gas contained in the syringe from the ideal gas equation:
- Q3. (a) Write the word equation and balanced formula equation for the reaction between the aluminum metal and the hydrochloric acid.
- (b) Calculate the actual number of moles of aluminum reacted from the mass of foil reacted and the atomic mass of aluminum given on the Periodic Table of the Elements.
- (c) One goal in this experiment is to compare the actual ratio of moles of hydrogen to moles of aluminum with the expected ratio predicted from the balanced chemical equation. Calculate the actual ratio of moles of hydrogen produced to the moles of aluminum metal reacted.

- (d) Calculate the percent error between the ratio of moles of hydrogen to moles of aluminum predicted in the balanced equation, Q3a, and the ratio calculated in part Q3c above.
- Q4. (a) Starting with the number of moles of hydrogen gas you calculated in question #Q2 and the balanced equation for the reaction, calculate the number of moles of aluminum metal which must have been reacted to produce the quantity of hydrogen collected.
- (b) Before modern instrumentation, the atomic mass of the elements was determined from experiments similar to this experiment. Calculate the atomic mass of aluminum from the mass of aluminum reacted and the number of moles of aluminum calculated in question #Q4a.
5. The final calculation is to find the molar volume of a gas at STP(101.3 kPa & 273 K) using the data from this experiment.
- (a) Calculate the number of moles of hydrogen expected from the moles of aluminum reacted and the theoretical ratio of moles of hydrogen to moles of aluminum.
- (b) Find the ratio of the volume of hydrogen collected in liters to the moles of hydrogen expected above. This will give you the molar volume (liters/mole) at the conditions of the experiment.
- (c) Correct the molar volume determined above from your experimental conditions to STP using the combination gas law equation. You must use the partial pressure of hydrogen corrected for the vapor pressure of water.
- (d) Find the percent error between your molar volume at STP and the expected value of 22.4 liters/mole.

Data:

	Trial #1	Trial #2	Trial #3
Mass of aluminum metal reacted	g.	g.	g.
Atmospheric pressure	kPa	kPa	kPa
Water temperature	°C	°C	°C
Volume of hydrogen collected	ml	ml	ml
Corrected Volume of hydrogen	ml	ml	ml
Vapor pressure of water (from the table below) [Q1a]	kPa	kPa	kPa
Partial pressure of the hydrogen [Q1b]	kPa	kPa	kPa
Moles of hydrogen collected [Q2]	mol	mol	mol
Moles of aluminum reacted [Q3b]	mol	mol	mol
Ratio of the moles of hydrogen collected to the moles aluminum reacted [Q3c]	$\frac{\text{mole H}_2}{\text{mole Al}}$	$\frac{\text{mole H}_2}{\text{mole Al}}$	$\frac{\text{mole H}_2}{\text{mole Al}}$
Percent error [Q3d]	%	%	%
Moles of Al reacted calculated from the moles of H ₂ collected [Q4a]	mol	mol	mol
Calculated atomic mass of aluminum [Q4b]	g/mol	g/mol	g/mol

°C	kPa	°C	kPa	°C	kPa	°C	kPa	°C	kPa
13	1.5	17	1.9	21	2.5	25	3.2	29	4.0
14	1.6	18	2.1	22	2.6	26	3.4	30	4.2
15	1.7	19	2.2	23	2.8	27	3.6	31	4.5
16	1.8	20	2.3	24	3.0	28	3.8	32	4.8