# EXPERIMENT 18

# Gaseous Diffusion

## MATERIALS AND EQUIPMENT

Solutions: concentrated ammonium hydroxide (in dropping bottle) and concentrated hydrochloric acid (in dropping bottle). Liquids: acetone and isopropyl alcohol. Glass diffusion tube, 65-80 cm long × 8 mm diameter (firepolished ends); two medicine dropper bulbs; utility clamp; rubber stopper, No. 4 or larger, one-hole, cut open on one side; metric ruler or meter stick; clock or watch with second hand (or stopwatch).

## DISCUSSION

The Kinetic Molecular Theory assumes that (1) the molecules of gases are in rapid random motion and (2) their average velocities (speeds) are proportional to the absolute (Kelvin) temperature. It also assumes that (3) at the same temperature, the average kinetic energies of the molecules of different gases are equal.

Graham's Law of Diffusion is based on the foregoing three basic assumptions of the Kinetic Molecular Theory and can be stated in either of two alternate forms: (1) The rates of diffusion of different gases are inversely proportional to the square roots of their densities, or (2) the rates of diffusion of different gases are inversely proportional to the square roots of their molar masses.

It is difficult to make experimental measurements of the velocities of individual gas molecules. But the **relative molecular velocities** of certain gases can be compared with the aid of some simple laboratory equipment. In this experiment you will determine the relative molecular velocities of ammonia ($NH_3$) gas and of hydrogen chloride (HCl) gas. This information will be used to calculate experimental values for (1) the molar mass of hydrogen chloride and (2) the ratio of the average molecular velocity of ammonia gas to that of hydrogen chloride gas.

Ammonia and hydrogen chloride gases react, upon contact, to form a white cloud (or smoke) of microscopic particles of solid ammonium chloride. This fact makes it fairly easy to **experimentally measure** the relative rates of diffusion of these gases. This is accomplished by simultaneously introducing ammonia and hydrogen chloride gases into the opposite ends of a glass tube and noting the time needed for the appearance of the faint white cloud of ammonium chloride. Experimental values for the relative rates of diffusion of $NH_3$ and HCl through the air in the tube can then be obtained by measuring the distance traveled by each gas and dividing by the time required for the appearance of the faint white cloud.

The kinetic molecular theory holds that, at the same temperature, the average kinetic energies of the molecules of different gases are equal. The kinetic energy (KE) of any moving body—regardless of whether it is a molecule of an automobile—is given by the equation, $KE = \frac{1}{2}mv^2$

(where $m$ is the mass of the body and $v$ is its velocity). Hence $KE_{(NH_3)} = KE_{(HCl)}$, and the masses and average velocities of $NH_3$ and HCl molecules must be related in this fashion:

$$\frac{m_{NH_3} v^2_{NH_3}}{2} = \frac{m_{HCl} v^2_{HCl}}{2} \tag{1}$$

By multiplying Equation (1) by 2 and rearranging, we obtain Equation (2)

$$\frac{v^2_{NH_3}}{v^2_{HCl}} = \frac{m_{HCl}}{m_{NH_3}} \tag{2}$$

By taking the square root of Equation (2) we obtain

$$\frac{v_{NH_3}}{v_{HCl}} = \frac{\sqrt{m_{HCl}}}{\sqrt{m_{NH_3}}} \tag{3}$$

Since the ratio of the molecular masses is proportional to the ratio of the molar masses, Equation (3) may be rewritten as

$$\frac{v_{NH_3}}{v_{HCl}} = \frac{\sqrt{M_{HCl}}}{\sqrt{M_{NH_3}}} \quad \text{(where } M \text{ represents molar mass)} \tag{4}$$

Equation (4) is, of course, the mathematical statement of Graham's Law of Diffusion in this form: The rates of diffusion of different gases are inversely proportional to the square roots of their molar masses.

From Equation (4) the ratio of the average molecular velocity of ammonia to that of hydrogen chloride can be calculated.

$$\frac{v_{NH_3}}{v_{HCl}} = \frac{\sqrt{M_{HCl}}}{\sqrt{M_{NH_3}}} = \frac{\sqrt{36.5}}{\sqrt{17.0}} = 1.47$$

This calculation tells us that, at a given temperature, the average velocity of ammonia molecules is theoretically 1.47 times greater than that of hydrogen chloride molecules.

## PROCEDURE

**Wear protective glasses.**

NOTES:

1. Use concentrated ammonium hydroxide and concentrated hydrochloric acid in dropping bottles or other containers designated by your instructor. Keep these containers tightly closed when not in use.

2. In performing Step 1, make sure that each liquid wets the entire inner wall of the tube. This is accomplished by rotating the tube while holding it at about a 45° angle while pouring the liquids through it.

3. Steps 4 and 5 are most efficiently performed by two persons working together. Therefore, unless otherwise directed by your instructor, arrange to cooperate with another student in the performance of these steps.

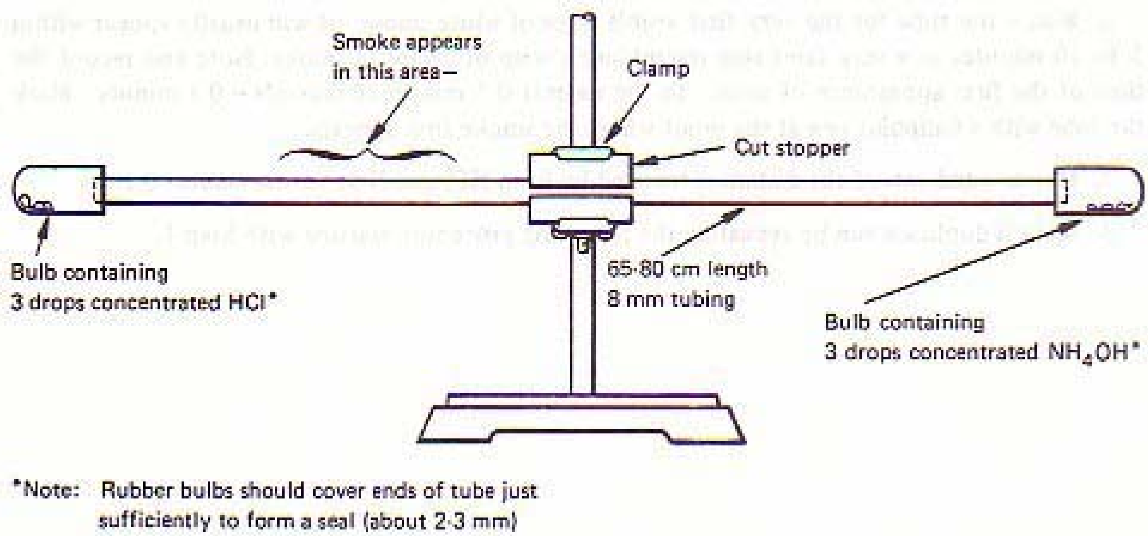


**Figure 18.1. Setup for Gaseous Diffusion.**

⚠ PRECAUTION: Isopropyl alcohol and acetone are volatile and highly flammable liquids.

1. Clean the glass diffusion tube by:

   (a) pouring warm water through it and rinsing with 5 mL of distilled water;

   (b) pouring about 3 mL of isopropyl alcohol through the tube;

   (c) pouring about 3 mL of acetone through the tube and wiping the outside clean with a towel or tissue;

   (d) drawing air through the tube using an aspirator pump or vacuum line for 5 minutes. **The tube must be perfectly dry for satisfactory results.**

2. Clean two medicine dropper bulbs by rinsing with water, then with 1 mL isopropyl alcohol, and finally with 1 mL acetone. Dry the bulbs thoroughly.

3. Slip a cut stopper in place at the center of the glass tube. Clamp the tube by means of this stopper and a utility clamp so that it is in a horizontal position about four inches above the desk top as shown in Figure 18.1.

4. Put **three drops** of conc. hydrochloric acid solution into one of the dropper bulbs, and at the same time have your partner put **three drops** of conc. ammonium hydroxide into the other bulb. Keep the bulbs turned open-end up so these liquids can't escape.

5. Without delay note the position of the second hand on your watch or clock; and as it crosses the next half minute or minute mark, working with your partner, simultaneously place the dropper bulb containing the ammonium hydroxide on the right end of the tube and the bulb containing hydrochloric acid on the left end. The bulbs should be **gripped lightly by the upper or rim portions,** and put on the ends of the glass tubing **only far enough to form a seal** (2 or 3 mm). **Do not squeeze the bulbs or push them very far onto the glass tube** because either of these actions will force air/gas mixtures into the glass tube and have an adverse effect on the experiment. Be sure to note and record the exact time when the bulbs were put on the tube.

6. Watch the tube for the very first visible trace of white smoke—it will usually appear within 5 to 10 minutes as a very faint ring resembling a wisp of cigarette smoke. Note and record the time of the first appearance of smoke to the nearest 0.1 minute (6 seconds = 0.1 minute). Mark the tube with a ballpoint pen at the point where the smoke first appears.

7. Measure and record the distances traveled by both HCl and $NH_3$ to the nearest 0.1 cm

8. Make a duplicate run by repeating the foregoing procedure starting with Step 1.

NAME ____________________

SECTION ________ DATE ________

INSTRUCTOR ____________________

## *REPORT FOR EXPERIMENT 18*

# Gaseous Diffusion

**Data Table**

| | Run 1 | Run 2 |
|---|---|---|
| Start Time | | |
| Finish Time (first visible smoke) | | |
| Elapsed Time (to nearest 0.1 min.) | | |
| Distance traveled by $NH_3$ (to nearest 0.1 cm) | | |
| Distance traveled by HCl (to nearest 0.1 cm) | | |
| $NH_3$ diffusion rate (cm/min.) | | |
| HCl diffusion rate (cm/min.) | | |

### CALCULATIONS

Show calculation setups and answers.

1. Experimental molar mass of HCl. Calculate the molar mass of HCl from the experimental data using Equation (4). Use 17.03 as the molar mass of $NH_3$. Suggestion: First square both sides of Equation (4) to remove the square root signs.

Run 1 ________ Run 2 ________

2. Percentage error in experimentally determined HCl molar mass (accepted value is 36.5).

$$\frac{\text{(Experimental value)} - \text{(Accepted value)}}{\text{Accepted value}} \times 100 = \text{Percentage error}$$

Run 1 ________ Run 2 ________

3. Experimental velocity ratio:

$$\frac{v_{NH_3}}{v_{HCl}}$$

Run 1________ Run 2________

4. Percent error in experimental ratio from No. 3 above. The accepted value of the ratio is 1.47–see Discussion.

Run 1________ Run 2________

### QUESTIONS AND PROBLEMS

1. Write the chemical equation for the reaction that caused the white smoke to appear inside the tube.

2. If 5 cm of the tube had been wetted with water on the inside at the $NH_3$ end, would the molar mass of HCl calculated from the experimental data probably be (1) greater than 36.46; (b) less than 36.46; or (c) equal to 36.46? Explain.

3. Which gas will diffuse faster, $O_2$ or $CH_4$? How many times faster?

4. In an experiment similar to the one which you performed, substance X was substituted for HCl. Ammonia was found to have a diffusion rate of 5.0 cm/min., and substance X, a rate of 4.6 cm/min. Calculate the molar mass of substance X. Show setup.

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