

GAS LAWS

EXPERIMENT 17

OBJECTIVE

The objective of this experiment is to determine the molar volume of hydrogen gas STP and identify an unknown metal using the gas laws.

EQUIPMENT AND CHEMICALS

Hydrochloric acid (HCl - conc)	100 ml graduated cylinder
Copper wire	Funnel (stemless)
Magnesium (Mg ribbon)	Funnel (with stem)
1000 ml beaker	Thermometer (0 – 100 °C)

DISCUSSION

Gases have general characteristics that separate them from other states of matter (solids and liquids).

- Expansion: Gas expands indefinitely to fill all the space available.
- Indefinite Shape and Volume: Gases have no definite shape or volume and fill the container.
- Compressibility: Gases are highly compressible.
- Low Density: Gases have very low densities and are measured in g/l.
- Diffusion: Will mix homogeneously to fill the entire shape.

Gases exert pressure (force per unit area) on the walls of the container. Pressure is nothing more than the impact of the gas molecules on the walls. Gas pressure is measured in various units with torr, atmosphere (atm), and pounds per square inch (psi) being the most common units.

$$1 \text{ atm} = 760 \text{ torr} = 14.7 \text{ psi}$$

The air in the atmosphere exerts pressure on the earth and can be measured. Since the force of air pushing down on the earth varies with altitude, we use the pressure of air at 0°C and sea level as the reference. This reference is called *standard temperature and pressure* (STP). The conditions are 0°C (temperature) and 1 atmosphere (pressure).

GENERAL GAS LAW

Robert Boyle, an English chemist, carried out experiments in 1660 which showed that at a constant temperature (the temperature remains the same), the volume of a gas is inversely proportional to the pressure. Stated in another way, if the volume (V) of the gas doubles, the pressure (P) decreases to half.

$$\begin{array}{ccccccc} P_1 & \times & P_1 & = & P_2 & \times & V_2 \\ \text{old} & & \text{old} & & \text{new} & & \text{new} \end{array} \quad (\text{Boyle's Law})$$

Jacque Charles, a French physicist, showed in 1787 that if the pressure is held constant, the volume of a gas increases directly proportionally to the temperature of the gas when the temperature is measured in Kelvin ($^{\circ}\text{C} = 273\text{K}$). For example, if the temperature (T) in Kelvin is doubled, the volume (V) of the gas would double.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad (\text{Charle's Law})$$

Joseph Gay-Lussac, a French physicist, showed in 1802 that if the volume is held constant, the pressure increases directly proportionally with the temperature in Kelvin. For example, if the pressure (P) doubles, the temperature (T) in Kelvin would also double.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad (\text{Charle's Law})$$

COMBINED GAS LAW

The three general gas laws can be put together to form what is called the Combined Gas Law. The temperature (T) must be in Kelvin.

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \quad (\text{Combined Gas Law})$$

Example 1:

A scuba tank of compressed air that has a volume of 80 cubic feet, a pressure of 3000 psi, and a temperature of 27°C (300 K) are left in the sun. The temperature inside the tank rises to 50°C (323 K). The volume of the tank does not vary. What would the new pressure of air inside the tank be?

Solution:

$$\begin{array}{ll} P_1 = 3000 \text{ psi} & P_2 = ? \\ T_1 = 300 \text{ K} & T_2 = 323 \text{ K} \\ V_1 = 80 \text{ ft}^3 & V_2 = 80 \text{ ft}^3 \end{array}$$

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \quad (\text{Combined Gas Law})$$

$$P_2 = \frac{P_1 V_1 T_2}{V_2 T_1}$$

$$P_2 = \frac{(3000 \text{ psi})(80 \text{ ft}^3)(323 \text{ K})}{(80 \text{ ft}^3)(300 \text{ K})} = 3230 \text{ psi}$$

Example 2:

A scuba tank is sitting on a bench. The tank has a volume 80 ft^3 , a pressure of 15 psi, and is at 25°C (298K). What temperature would be necessary to increase the pressure to 3000 psi? (Assume the scuba tank will not melt.)

Solution:

$$\begin{array}{ll} P_1 = 15 \text{ psi} & P_2 = 3000 \text{ psi} \\ T_1 = 298 \text{ K} & T_2 = ? \\ V_1 = 80 \text{ ft}^3 & V_2 = 80 \text{ ft}^3 \end{array}$$

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \quad (\text{Combined Gas Law})$$

$$T_2 = \frac{P_2 V_2 T_1}{V_1 P_1}$$

$$T_2 = \frac{(3000 \text{ psi})(80 \text{ ft}^3)(298 \text{ K})}{(80 \text{ ft}^3)(15 \text{ psi})} = 59,600 \text{ K} = 59,327 \text{ }^\circ\text{C}$$

This shows that if pressure is increased, the temperature also increases proportionally.

Example 3:

The volume of a cylinder in a car is 50 cubic inches at STP (298 K and 14.7 psi). What would be the volume of the cylinder if the piston compressed the gas to 245 psi at 298 K?

$$\begin{array}{ll} P_1 = 14.7 \text{ psi} & P_2 = 245 \text{ psi} \\ T_1 = 298 \text{ K} & T_2 = 298 \text{ K} \\ V_1 = 50 \text{ in}^3 & V_2 = ? \end{array}$$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \quad (\text{Combined Gas Law})$$

$$V_2 = \frac{P_1V_1T_2}{T_1P_2}$$

$$V_2 = \frac{(14.7 \text{ psi})(50 \text{ in}^3)(298 \text{ K})}{(298 \text{ K})(245 \text{ psi})} = 3.0 \text{ in}^3$$

DALTON'S LAW OF PARTIAL PRESSURE

John Dalton, an English physicist, showed that each gas in a mixture of gases exerts a pressure equal to the pressure it would exert if it were the only gas in the same column; the total pressure of the mixture is then the sum of the partial pressure of all the gases present. In other words, if two gasses, oxygen and nitrogen, are present in a 1-liter container, the partial pressure of the oxygen is 121 torr, and the partial pressure of the nitrogen is 639 torr, the total pressure would be 760 torr.

$$P_{\text{total}} = P_1 + P_2 + P_3 + P_{\text{etc}} \quad (\text{Dalton's Law of Partial Pressure})$$

Water vapor is also a gas that exerts a partial pressure. If a gas were collected over water, the partial pressure of the water vapor would be constant at any given temperature. Therefore, the total pressure of the collected gas would be the partial pressure of the “dry gas” and the partial pressure of the water vapor.

$$P_{\text{total}} = P_{\text{dry gas}} + P_{\text{water}}$$

To obtain the vapor pressure of water at any temperature, refer to Table 1:

Temp (°C)	Press (torr)	Temp (°C)	Press (torr)
0	4.6	22	19.8
5	6.5	23	21.1
10	9.2	24	22.4
15	12.8	25	23.8
16	13.6	26	25.2
17	14.5	27	26.7
18	15.5	28	28.3
19	16.5	29	30.0
20	17.5	30	31.8
21	18.6	31	33.7

PARTIAL PRESSURE OF WATER VAPOR

Example 4:

Hydrogen is collected over water. The temperature of the gas is 25°C. The total pressure of hydrogen is 800 torr. What is the partial pressure of “dry” hydrogen gas?

Solution:

$$P_{\text{total}} = P_{\text{dry gas}} + P_{\text{water}}$$

$$P_{\text{total}} = P_{\text{H}_2} + P_{\text{water}}$$

$$P_{\text{H}_2} = P_{\text{total}} - P_{\text{water}}$$

Referring to Table 1, the partial pressure of water at 25 °C = 23.8 torr

$$P_{\text{H}_2} = 800 \text{ torr} - 23.8 \text{ torr} = 776.2 \text{ torr}$$

The partial pressure of the “dry” hydrogen is 776.2 torr.

MOLAR VOLUME

If a balloon were filled with one mole of any gas at 0°C and 1 atmosphere (STP), the volume of the balloon would be 22.4 liters.

$$1 \text{ mole of a gas} = 22.4 \text{ liters at STP}$$

The molar volume is useful for calculating the molecular weight of an unknown gas.

Example 5:

A balloon has a volume of 11.2 liters at STP (0°C and 1 atm). The gas in the balloon has a mass of 16 grams. What is the molecular mass of the gas?

Solution:

$$1 \text{ mole} = 22.4 \text{ liters (STP)}$$

$$\text{mole} = \frac{\text{grams}}{\text{mol mass}}$$

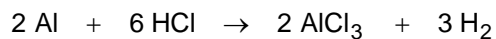
$$\text{mol mass} = \frac{\text{grams}}{\text{moles}}$$

$$\text{mol mass} = \frac{16 \text{ grams}}{0.5 \text{ moles}} = 32 \text{ amu} = \text{O}_2$$

Example 6

Aluminum (atomic mass = 26.98 amu) reacts with HCl to produce hydrogen gas. The volume of gas produced at STP was 44.8 liters. How much Al was used in the reaction?

Solution:



$$\text{moles H}_2 = 44.8 \text{ liters H}_2 \times \frac{1 \text{ mole H}_2}{44.8 \text{ liters H}_2} = 2 \text{ moles}$$

$$\text{moles Al} = 2 \text{ moles H}_2 \times \frac{2 \text{ mole Al}}{3 \text{ moles H}_2} = 1.33 \text{ moles of Al}$$

$$\text{grams Al} = 1.33 \text{ moles of Al} \times 26.98 \text{ amu} = 35.97 \text{ grams of Al}$$

IDEAL GAS LAW

The Ideal Gas Law combines the general gas laws and the molar volume of a gas into one equation.

$$PV = nRT \quad (\text{Ideal Gas Law})$$

The pressure (P) is in atm; the volume (V) is in liters; n is the number of moles; the temperature (T) is in Kelvin; and R is the gas constant which is equal to 0.0821 liter atm / mole K .

Example 7:

A balloon is to be filled with hot air. The air temperature is 100°C (398 K); the pressure is 760 torr (1atm), and there are 1000 moles of gas. What would be the volume of the hot air balloon?

Solution:

$$P = 1 \text{ atm}$$

$$n = 1000 \text{ moles}$$

$$V = ?$$

$$R = 0.0821 \text{ liter atm / mole K}$$

$$T = 398 \text{ K}$$

$$PV = nRT$$

$$V = \frac{nRT}{P}$$

$$V = \frac{(1000 \text{ moles}) (0.0821 \text{ liter atm / mole K}) (398 \text{ K})}{(1 \text{ atm})} = 32,676 \text{ liters}$$

DETERMINATION OF THE MOLECULAR WEIGHT OF A METAL

Using the gas laws, the molecular weight of a metal can be determined.

Example 8:

Calculate the molecular weight of a metal given the following information. A 0.55 gram sample of the metal was reacted with an acid and evolved 200 ml of hydrogen gas (H_2). Assume that one mole of the metal generates one mole of H_2 gas. The gas was collected over water at a temperature of $25^\circ C$ (298K) and at a pressure of 808 torr.

Solution:

The partial pressure of the “dry” hydrogen gas is determined at $25^\circ C$.

$$P_{\text{total}} = P_{\text{dry gas}} + P_{\text{water}}$$

$$P_{\text{total}} = P_{H_2} + P_{\text{water}}$$

$$P_{H_2} = P_{\text{total}} - P_{\text{water}}$$

Referring to Table 1, the partial pressure of water at $25^\circ C = 23.8$ torr

$$P_{H_2} = 808 \text{ torr} - 23.8 \text{ torr} = 784.2 \text{ torr}$$

The partial pressure of the “dry” hydrogen is 784.2 torr.

The volume of hydrogen is corrected to STP conditions (298K and 760 torr).

$$\begin{array}{ll} P_1 = 784.2 \text{ torr} & P_2 = 760 \text{ torr} \\ T_1 = 298 \text{ K} & T_2 = 273 \text{ K} \\ V_1 = 200 \text{ ml} & V_2 = ? \end{array}$$

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \quad (\text{Combined Gas Law})$$

$$V_2 = \frac{P_1 V_1 T_2}{T_1 P_2}$$

$$V_2 = \frac{(784.2 \text{ torr})(200 \text{ ml})(273 \text{ K})}{(298 \text{ K})(760 \text{ torr})} = 189 \text{ ml}$$

The volume of the hydrogen gas would be 189ml (0.189 l) if the gas were under STP conditions.

The moles of hydrogen gas are now calculated using the molar volume relationship.

1 mole of a gas STP = 22.4 liters

$$\text{moles H}_2 = 0.189 \text{ liters of H}_2 \times \frac{1 \text{ mole}}{22.4 \text{ liter}} = 0.0084 \text{ moles}$$

If we only have 0.189 liters collected, the number of moles of H₂ would be very small.

Assuming that one mole of metal = one mole of H₂, there would be 0.0084 moles of metal.

Moles are equal to the grams of metal divided by the atomic mass of the metal. We reacted 0.55 grams of metal.

$$\text{mole of metal} = \frac{\text{grams}}{\text{mol mass}}$$

$$\text{mol mass} = \frac{\text{grams of metal}}{\text{moles of metal}}$$

$$\text{mol mass} = \frac{0.55 \text{ grams}}{0.0084 \text{ moles}} = 65.4 \text{ amu} = \text{Zn}$$

PROCEDURE

PART A SET UP OF APPARATUS

The laboratory apparatus used to collect the hydrogen evolved when the magnesium is reacted with hydrochloric acid will be set up.

1. Obtain a piece of pre-weighed magnesium (Mg) ribbon from your instructor. Record the mass of the magnesium on your Report Sheet.
2. Roll the magnesium ribbon into a loose jellyroll and wrap the magnesium ribbon with several turnings of copper wire. Leave approximately a one-foot length of copper wire for anchoring the piece of Mg later on.
3. Fill a 1000 ml beaker half-full of water.
4. Place the magnesium ribbon on the bottom of the beaker and place a stem-less funnel on top of the wire.
5. Run the copper wire underneath the funnel and over the side of the beaker. This will prevent the magnesium from floating up to the surface of the water during the reaction.
6. Fill a 100 ml graduated cylinder with water and place a stopper in the end.

7. Invert the funnel and place the lip with the stopper underneath the surface of the water in the 1000 ml beaker and remove the stopper.
8. If any air bubbles get into the graduated cylinder, refill the graduated cylinder and repeat until no air bubbles are present.
9. Place a long stem funnel inside the beaker. This funnel will later be used for the addition of concentrated hydrochloric acid.

PART B GENERATION AND MEASUREMENT OF HYDROGEN GAS

Hydrochloric acid (HCl) is added to the beaker and because of its greater density sinks to the bottom of the beaker. The HCl reacts with the magnesium ribbon to produce hydrogen gas. The volume of hydrogen gas is then measured after equalizing to atmospheric pressure.

1. Carefully add 30 ml of concentrated HCl to the bottom of the beaker using a long stem funnel.
2. Allow the magnesium ribbon to completely react with the HCl until all the magnesium ribbon is gone. If hydrogen bubbles stop and magnesium ribbon remains, add an additional 30 ml of concentrated HCl.
3. After all the magnesium is reacted; adjust the level of the water in the graduated cylinder to match the level of water in the beaker by adding additional water to the beaker, or raising the cylinder upward.
4. Record the volume of gas in the graduated cylinder on your report sheet when the two water levels are equal.
5. Using a thermometer, record the temperature of the water in the beaker on your report sheet. This temperature will be the same temperature as the gas in the graduated cylinder.
6. Obtain the barometric pressure of the atmosphere from your instructor and record on your Report Sheet.
7. Carefully pour the acidic water in your beaker down the sink and flush with water. Rinse sink and the apparatus with water.

PART C IDENTIFICATION OF AN UNKNOWN METAL

An unknown metal will be obtained from your instructor and the metal identified by the calculation of its molecular weight.

1. Obtain a pre-weighed unknown metal from your instructor and record the weight.
2. Repeat the experiment using the metal. Assume a one mole of metal = one mole of hydrogen unless the instructor indicates otherwise.