# General Chemistry Laboratory Fall 2011

# The molar volume of gases

(from General Chemistry Laboratory Manual, P.A.M. van Koppen, McGraw-Hill, 2003)

 In this lab you will learn how to quantitatively measure the volume of a gas, and how to calculate the molar volume of a gas at standard temperature and pressure. To do these things, you will apply the ideal gas law to measure the volume of oxygen at the experimental temperature and pressure, and you will then calculate the volume of the gas at standard temperature and pressure (STP; 0°C, 1 atm). It may help with the lab if you try to imagine yourself in an early 19th Century chemistry laboratory, pushing back the frontiers of science, wearing a powdered wig and a high, uncomfortable collar on your jacket.

 The ideal gas law, PV = nRT, is derived from three other laws (P = pressure, V = volume, n = number of moles, R = universal gas constant, T = temperature, in K):

* **Avogadro’s law** states that the volume of a gas (at constant temperature and pressure) is proportional to the number of moles of the gas, n: V α n
* **Boyle’s law** states that the volume of a fixed quantity of gas at a fixed temperature in inversely proportional to pressure (as volume increases, pressure decreases): P α (1/V)
* **Charles’ law** states that for a fixed quantity of gas at a constant pressure the volume increases as the temperature increases: V α T

The term ideal gas is used to evoke a collection of non-interacting particles in which a few particles move independently in a vast empty space – the least constrained form of matter. The universal gas constant is derived from a rearrangement of the ideal gas law: (PV)/(nT) = R. This universal gas constant can be measured experimentally, and is the same for all ideal gases because they all have the same molar volume (see table below). So if we know R and three of

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| --- | --- | --- | --- |
| **Gas** | **Density,****g/liter** | **Molar mass,****G/mol** | **Molar volume,****(liter/mol)** |
| CO | 1.250 | 28.010 | 22.406 |
| N2 | 1.250 | 28.013 | 22.403 |
| He | 0.178 | 4.003 | 22.429 |
| H2 | 0.0899 | 2.016 | 22.429 |
| CO2 | 1.977 | 44.010 | 22.262 |
| HCl | 1.639 | 36.461 | 22.244 |
| O2 | 1.429 | 31.999 | 22.392 |

the other variables, we can solve the ideal gas law for the fifth variable. For this lab we will also need Dalton’s law of partial pressures: *each gas exerts its own pressure, and that in a mixture of gases the total pressure is equal to the sum of the pressures of all of the gases: Ptotal = P1 + P2 + P3……. + Px* .

 In our experiment we will have to take into account the pressure of water vapor in our experimental apparatus. For example, we’ll collect oxygen gas over water, where Ptotal = PO2 + PH2O. So if we know Ptotal and PH20, we can calculate PO2. Fortunately, we can obtain PH2O from a graph of vapor pressure for water at different temperatures (see graph at end of protocol).

 To determine the molar volume of oxygen, you will use the decomposition of hydrogen peroxide to produce water and oxygen, catalyzed by FeCl3:

  . The law of mass conservation says the difference in the mass of reactants and the mass of products will equal the mass of oxygen produced in the reaction.

**Procedure –** **Production of oxygen gas**

Assemble an apparatus as shown in the following figure:

AB

BC

pinch clamp

stopper

tubing

A

C

B

250 ml flask (**A**) 500 ml flask (**B**) 600 ml beaker (**C**)

* Measure 20 ml of 3% H2O2 in a graduated cylinder and add it to flask A.
* Measure 4.5 ml of 3 M FeCl3 and pour into a small test tube
	+ Carefully place the test tube *upright* inside flask A
	+ Record the weight of flask A and its contents in the chart below:

|  |  |
| --- | --- |
| **Decomposition of H2O2 using FeCl3 catalyst** |  |
| Mass of flask A & contents before reaction |  |
| Mass of flask A & contents after reaction |  |
| Temp of gas in flask A (ie, temp of liquid) |  |
| Temp of gas in flask B (ie, temp of liquid) |  |
| Average temperature of gas |  |
| Volume of oxygen collected |  |
| Barometric pressure |  |
| Water vapor pressure (from table) |  |

* Fill flask B with tap water up to the neck
* Fill beaker C about one third with tap water
* Disconnect the stopper at point A and attach a pipette bulb to the tube
	+ With the pinch clamp open, force air into tube AB by squeezing the bulb, forcing water from flask B into beaker C
	+ Raise and lower beaker C to move water back and forth through tube BC to remove air bubbles
	+ With water halfway up the neck of flask B, reconnect the stopper at point A
* With the clamp open, test the system for leaks:
	+ Raise beaker C as high as possible – the water level in flask B should rise a little, then remain fixed
	+ If the water level in flask B continues to rise, then there is a leak that needs to be sealed
* Equalize the pressure inside and outside of flask B by raising beaker C until the water level in it and beaker C is the same – close the pinch clamp tightly
	+ Empty the water out of beaker C
	+ Put the tube back into beaker C and open the clamp
	+ A little water may run into the beaker – leave it there
* Tip flask A so that the FeCl3 catalyst spills into the H2O2 in the flask
	+ Avoid getting the solution on the stopper
	+ Record any changes you see
		- Also, feel the flask – is heat released from the reaction?
	+ Note the water level in beaker C
		- When the reaction stops, and it remains constant, the decomposition of the H2O2 is complete
* Raise beaker C until the water level in it and flask B is the same
	+ Close the pinch clamp and measure the volume of water in beaker C by pouring it into a graduated cyclinder
		- Record this volume of water displaced by oxygen
* Measure the temperature of the of the liquid in flasks A and B
	+ The average of these can be used to approximate the temperature of the gas produced in the reaction
* Disconnect flask A and weigh and record the flask and its content
* Obtain the barometric pressure from www.weather.com

**Calculating the molar volume of oxygen gas**

Since you measured the volume of O2 produced at known temperature and pressure, you can calculate the molar volume of O2 in liters/mol (volume of O2 produced / moles of O2 produced).

 As shown above, Ptotal = PO2 + PH2O . We can measure Ptotal, and we can find PH2O from the graph below (using the average temperature of the liquid in flasks A and B), so we can then calculate PO2: PO2 = Ptotal – PH2O .

 The state of a gas is defined by its pressure, volume, temperature and the number of moles of the gas. For the change in state from P1, V1, T1, n1 to P2, V2, T2, n2 we can write the equation:

  ; which can be rearranged, at constant number of moles, to:

  ; and then, to solve for V2:  .

 If you use your measurements as P1, V1 & T1, then you can calculate V2 at standard state: 0°C (T2) and 1 atm (P2).

|  |  |
| --- | --- |
| **Measure** | **Data** |
| Mass of oxygen |  |
| Moles of oxygen |  |
| Average temp of gas (°K) (= T1) |  |
| Pressure of O2, PO2 ( = P1) |  |
| Volume of O2 at STP (from your data, = V1) |  |
| Molar volume of O2 at STP (L/mol, = V2) |  |

 Compare your calculated value for the molar volume of O2 to the expected value given in the first table. Is this method for measuring gas volume reliable?