Question: During a reaction, oxygen gas is collected over water (see **F12-4-3**). The data recorded is below. How many moles of O_2(g) were colleced? (Assume that all gases behave ideally.)

- Volume of gas collected: **55.7 mL**
- Temperature of room air and of water: **23.0℃**
- Barometric pressure: **742 mm Hg**
- Vapor pressure of water at 23.0℃: **21.1 mm Hg**

Concepts: To answer this question we need to know the process by which gas samples are collected over water and what the properties are of a mixture of gases.

Connections:

What is given? We are given specific experimental data from a laboratory experiment where a sample of gas was collected over water.

What do I want to know? We want to know how many moles of oxygen gas were collected as part of the experiment.

When a sample of gas is collected over water it is bubbled through the water and then trapped in a container above the water line. The height of the container is then adjusted so that the water inside is level with the water outside (like a manometer of with even heights). This works much like a manometer and tells us that the total pressure inside the container is equal to the atmospheric pressure at the time of the experiment. This is useful because directly measuring atmospheric pressure is much more easily done than directly measuring the pressure inside the sample container.

The temperature of the experiment is important because it directly relates to vapor pressure. Recall that the vapor pressure is the pressure of the vapor above the liquid sample (or solid though these are much lower). So, in the case of water it would be the pressure exerted by the water vapor above a sample of water. This is caused by having a few molecules of liquid with enough kinetic energy to overcome IMFs and become a vapor (recall speed distributions of gas molecules). As temperature goes up, more molecules have the energy required and vapor pressure increases. Vapor pressure tables are available based on temperature so you could look up the required value.

The mixture above the liquid water will therefore be composed of both water vapor and the oxygen gas collected. Dalton's Law of Partial Pressures tells us that the total pressure of this mixture will be equal to the sum of the individual partial pressures. Recall as well that the partial pressure of a gas in a mixture will be related to the number of moles of the gas of interest (remember that all gases in a mixture have the same volume!)

Be sure you understand everything above before moving on to the solution below.

Solution:

Let's organize the information given and determine a plan of attack:

Volume of gas collected: 55.7 mL

Temperature of room air and of water: 23.0°C

Barometric pressure: 742 mm Hg

Vapor pressure of water at 23.0°C: 21.1 mm Hg

We're asked for the number of moles of oxygen and a good starting point is usually the ideal gas law. Remember that the ideal gas law can be applied to mixtures as a whole or individual components of gas mixtures:

$P_{O2}V = n_{O2}RT$ since we want moles of oxygen we rearrange:

$P_{O2}V/(RT) = n_{O2}$

If we use the partial pressure of oxygen in the mixture we can get moles of oxygen.

We are given the volume and temperature and R is a constant. So, all we need is the partial pressure of oxygen. We know that the mixture is composed of oxygen and water vapor and that the total pressure is 742 mm Hg. Therefore we can say:

$P_t = P_{O2} + P_{H2O}$

The method of the experiment tells us that the total pressure of the mixture is the same as the atmospheric pressure:

742 mm Hg = $P_{O2} + P_{H2O}$

We also know that the partial pressure of water is going to be the **vapor pressure** of water:

742 mm Hg = P_{O2} + 21.1 mm Hg

Therefore, the partial pressure of oxygen must be:

$742 - 21.1$ mm Hg = 720.9 mm Hg

From our initial analysis we know:

$P_{O2}V/(RT) = n_{O2}$

and we now have the partial pressure of oxygen. We are using pressure units so R must be 0.0821 L-atm/mol-K

Our units tell us we need to have pressure in atm, volume in L and temperature in K so we must convert:

720.9 mm Hg / 760 mm Hg/atm = 0.948 atm

V = 55.7 mL / 1000 mL/L = 0.0557 L

T = 23.0°C + 273.15 = 296.15 K

$n_{O2} = P_{O2}V/(RT)$ = (0.948 atm) (0.0557 L)/(0.0821 L-atm/K-mol)(296.15 K)

n_{O2} = 0.00217 mol