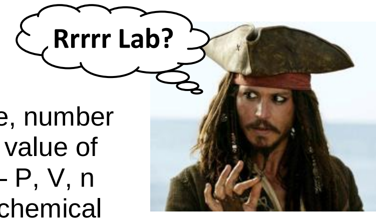
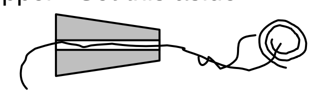
R Lab Name: ______________________



You have learned that "R" is the gas law constant that relates a gas sample's pressure, volume, number of moles and temperature. Given any three of these properties, R allows you to determine the value of the fourth. But how is R itself determined? In this lab you will measure – directly or indirectly – P, V, n and T, and then use them to determine an experimental value for R. To do this you will use a chemical reaction that produces a gas – then use stoichiometry to determine the number of moles of the gas, collect this gas in a long incremented test tube – so you will know the volume of the gas, look up the current local barometric pressure and make the necessary adjustments on it to determine the pressure of the gas, and of course use a thermometer to measure the temperature of the gas.

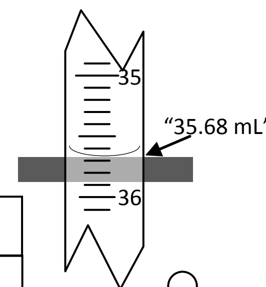
Procedure:

1) Obtain a piece of magnesium ribbon, and use a metric ruler to measure very precisely its lengths (top and bottom) in cm. Record these in the data table below. Then roll the ribbon up into a small coil as shown at right. Loop a small wire through the coil and pass the other end of the wire through the hole in the small end of a rubber stopper. Set this aside.



2) Holding the bottom end of gas collecting tube in the sink to catch any spillage, carefully add about 15-20 mL of 3 M HCl solution. Then add tap water on top of the HCl all the way to the very top of the tube. Clamp this tube vertically in the buret clamp. Now insert the magnesium coil and stopper into the tube. Be sure there is no air pocket. (If there is, remove the stopper, add some more water, and reinsert the stopper.)

3) Now, carefully flip the tube over, holding the stopper in place as you do so, and place the stoppered end in the plastic cup filled with water. Then clamp the tube in this position. Observe what happens as the more dense HCl solution streams through the water – this is an effect known as schlieren.* As you are waiting for the reaction to happen, go to one of the computers and record the current reported barometric pressure. Then do the moles calculation – #1 below.

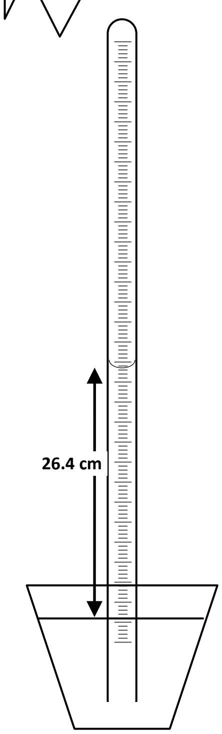


4) Once the reaction starts, observe the magnesium. Eventually enough will have reacted that what's left may break away and float to the top. This is OK, as long as it continues to react and doesn't get stuck on the side of the tube. Watch for this. Once the reaction is over, make a measurement of the volume – as always, use the black band to accentuate the meniscus and read from the bottom of the meniscus. Also, be careful: the scale is upside down. Record this volume in the table below.

5) Take the temperature of the water in the cup, which we will take to be the temperature of the gas. Record this in the table below.

6) Finally, measure the height of the water column beneath the gas sample in the tube. Record this in the table.

Length of Mg ribbon (cm) top|bottom		
Reported barometric pressure (inches Hg)		
Volume of gas collected (mL)		
Temperature of the water & gas (*C)		
Height of the water column (cm)		



Calculations:

<u>1) Moles:</u> Usually to determine the number of moles of a substance, you simply weigh the sample and divide the mass by the molecular weight, but since gases have such low densities, they are almost impossible to weigh accurately. So instead, we started with a small piece of a solid – magnesium – and used it to generate a gas. So, write the <u>balanced</u> equation for the reaction that happens between HCl and Mg (Hint: remember: it produces a gas.)

Balanced equation: ____.

Now, calculate how many moles of gas were produced. You could have weighed the magnesium first, then simply converted grams Mg → moles Mg, then used the balanced equation to convert moles Mg →moles H_2. If you had done this, however, you would have found that the tiny piece of Mg was so light that even on our best scale that reads to the thousandth of a gram, it wouldn't have given us a very precise mass – probably only one or two sig figs. So instead, the length and mass of a very long magnesium ribbon were measured and this information is posted on the board. Now, take the average of the two lengths you measured, and – using the conversion posted on the board – convert from cm Mg → g Mg → moles Mg → moles H_2. This will require three factor label steps:

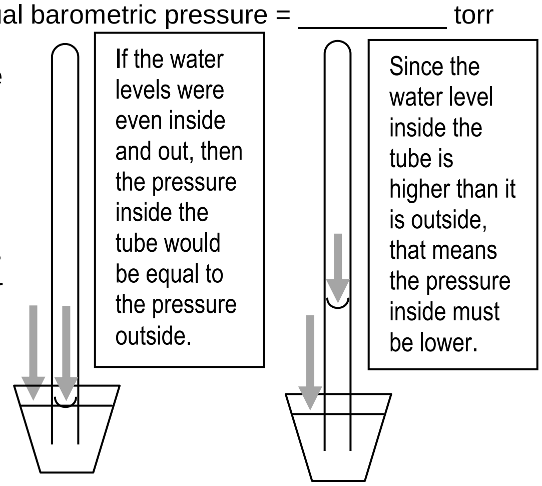
moles H_2 = _______

*schlieren = visible streaks produced in a transparent medium as a result of variations in the medium's density leading to variations in refractive index

2) Pressure: There are four corrections that need to be made to get us from a reported barometric pressure (in inches of Hg) to the actual pressure of the H_2 gas in the tube (in torr).
A) First, convert the reported barometric pressure from inches Hg to mmHg. Use factor label: 1 inch = 25.4 mm. And since a "mm Hg" is the same thing as a "torr." We now have reported barometric pressure in torr.

Show work: Reported barometric pressure = _________ torr

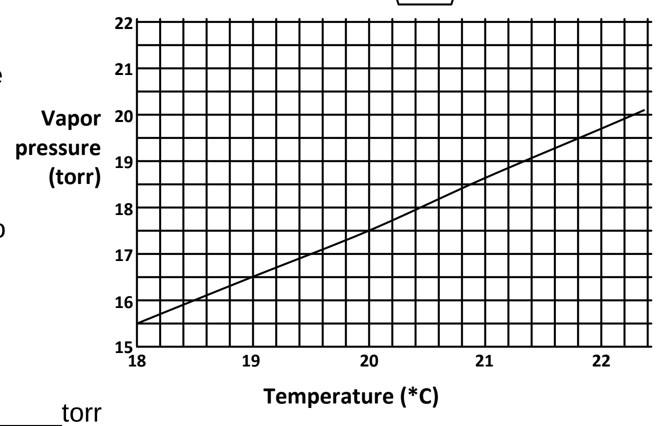
B) As it turns out, the reported barometric pressure is not the actual barometric pressure. (That's right! The meteorologists are lying to us!) Reported barometric pressures have all been adjusted to sea level elevation. In other words, the reported barometric pressure in Kirkwood at a given time is what the pressure *would be* if Kirkwood were at sea level. (The reason for doing this will be discussed in the follow-up questions.) But Kirkwood is obviously not at sea level. KHS, for example, is at an elevation of 222 m (639 ft). As you move upward through the atmosphere, the air gets progressively thinner and thinner and this decreases the pressure by 0.0866 torr for every meter. So a meteorologist measures the pressure in Kirkwood and adds 222 m x $(^{0.0866\text{ torr}}/_{1\text{ m}})$ = 19.2 torr to it to report what the pressure would be at sea level. So to determine the actual barometric pressure, simply take the value from A above and subtract 19.2 torr from it. Actual barometric pressure = __________ torr



C) So… that's the actual pressure in the room, but it is not the pressure inside the gas collecting tube. If the pressures were the same, the liquid levels inside and outside the tube would be even. Since the level inside the tube is in fact higher, it must mean that the pressure outside the tube is greater – enough to push the water up by ________ cm. And this is equal to _________ mm. And since water's density is 1 g/mL and mercury's density is 13.6 g/mL, simply divide the mm water by 13.6 to convert this column of water into an equivalent column of mercury: ________ mmHg
And since a mmHg is the same as a torr, subtract this number from the actual barometric pressure in the room to get the actual pressure in the tube.

Pressure in tube: _______torr

D) Finally, we would be done with all these pressure adjustments if the gas inside the tube were pure hydrogen, but since the hydrogen was collected over water, and water is always evaporating, there is a little water vapor mixed in. The pressure of that vapor (known as "vapor pressure") is related to the temperature of the water by the graph at right. Use this graph along with the water temperature you recorded to find what the vapor pressure was inside the tube: __________ torr



Now subtract that vapor pressure from the total pressure in the tube (the one derived in part C above) to determine the pressure of just the hydrogen gas in the tube

Hydrogen gas pressure: ________torr

3) Volume: Easy, just convert mL into L (you should know how to do this!): Volume = ___________ L

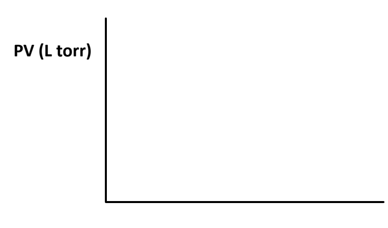
4) Temperature: Also easy, just convert *C into K (you should know how to do this too!): Temperature = _________ K

5) R: Now that you have all the pieces… Since PV = nRT, we can rewrite this as **R =** $^{PV}/_{nT}$

So… calculate your PxV = ___________ L torr and your n x T = _____________ mol K.

R = ________________

6) Pooling class data. Go to the desktop computer and enter your nxT and then your PxV values. These will be graphed (PV as a function of nT). At right, sketch what you think this graph will look like. What purpose would this graph serve?



Follow up questions:

1. One thing that was not mentioned was that the long piece of magnesium ribbon was buffed with piece of steel wool until it was completely shiny. Why was this important to do?

2. Should this buffing of the magnesium be done before or after the ribbon's length and mass were measured? ________ Why?

3. The 3M HCl solution has a density of about 1.04 g/mL, and water has a density of 1.00 g/mL If the HCl had been less dense than the water, why would the procedure not have worked the same?

4. Magnesium has a density of 1.74 g/mL – which is pretty low for a metal, but still much greater than the density of the water or the HCl solution. So why did the little piece of Mg float up to the top of the tube when it broke free of the wire loop.

5. When Sammy flipped the tube over, a bubble of air somehow got in. "Oh great, he said, "Now we have to start all over!" "Oh no we don't!" said his astute lab partner Sally… What did Sally realize they could do?

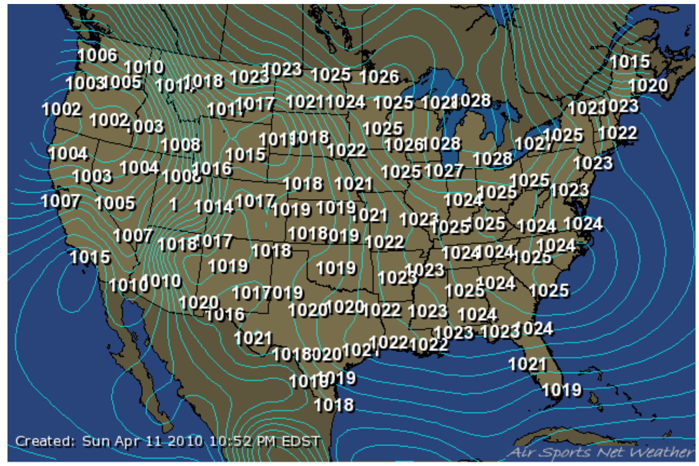
6. You repeat the entire experiment in Denver and record the following data. Determine R based on this information.
SHOW ALL WORK:

Lengths of Mg ribbon (cm)………	3.56 | 3.62
Volume of H_2 collected (mL)……………	37.87
Height of water column (cm)…………….	16.75
Temperature of water & gas (*C)………..	24.7
Mass of 250.0 cm of Mg (g)……………..	1.876
Altitude of Denver (m)…………………….	1672
Reported bar. Pressure(inches Hg)…..	29.85

R = ____________

7. Why magnesium? Give at least two reasons for this choice.

8. So why do meteorologists adjust all their reported barometric pressures to sea level?   Hint, look at the map at right and consider how it would look if the pressures were not adjusted.



9. For each of the following potential error sources, write "**H**" if it would have made the R value come out too high, **"L"** if it would have made the R value come out too low, or **"X"** if it would not have had any effect at all.

__ Not all the Mg reacted; some of it got caught on the side of the tube.

__ A bubble of air got in when the tube was flipped over but no one noticed.

__ The magnesium wasn't buffed enough before- hand.

__ Some of the hydrogen dissolved in the water as it bubbled up in the tube.

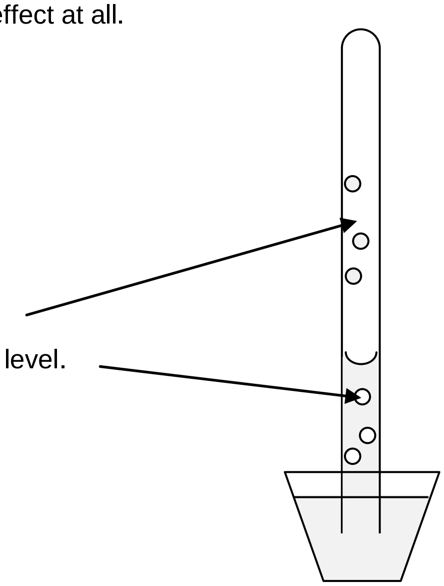
__ There were drops of water clinging to the inside of the tube above the water level.

__ There were bubbles of hydrogen clinging to the inside of the tube below the water level.

__ You accidentally recorded 33.27 mL as "33.72 mL."

__ You forgot to change mL into L.

__ You forgot to subtract out the vapor pressure



__ 10. Consider the following measurement errors
A) being 1 cm off in the measuring of the water column height
B) being 1 cm off in the measuring of your Mg ribbon lengths
C) being 1*C off in the measuring of the temperature of the water/gas
D) being 1 mL off in the measuring of the volume of gas collected
E) being 1 inch off in the recording of the reported barometric pressure

Rank these five from greatest impact on your results (the one that would throw off your experimental R value the most) to lowest impact.   ___ ___ ___ ___ ___