1. Lab 1.1 Finding a Method of Collecting Oxygen Gas

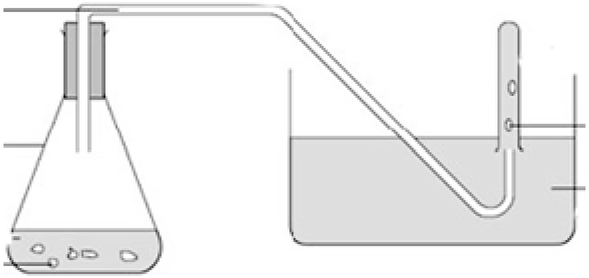
1. *What happened in the lab?*

- The reaction was: $2\ H_2O_{2(l)} \rightarrow O_{2(g)} + 2\ H_2O_{(l)}$

- **T**he catalyst KI was used to speed up an otherwise impractically slow reaction
- We found a way to trap and collect the oxygen

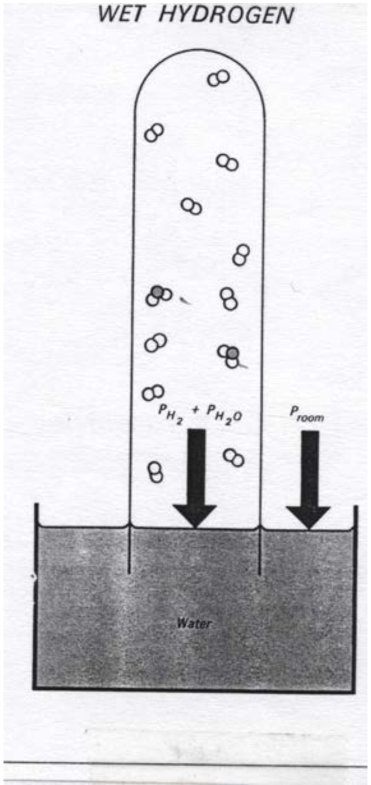
2. *How was it done?*

The test tubes were filled with water, held upside down, and the delivery tube from the reaction flask was inserted in them. The oxygen's pressure pushed the water out. When the tubes were empty of water, they were filled with oxygen. The glowing splint test confirmed their presence.



3. *What were some error sources?*

If the first tube or two were not discarded, then they simply filled with the original air from the flask and tubes. The error could be minimized by using more liquid hydrogen peroxide in the flask on the left. But too much of it can risk sending it into the tubes and making a royal mess.



2. Lab 1.2 The Reaction of Mg and HCl to Generate H_2 Gas

1. *What happened in the lab?*

- The reaction was: $Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$

We trapped Mg in a copper cage. Strong acid was added to a burette and topped with water. We inserted the cage into the water, stoppered the burette and turned it upside down. The denser acid slowly descended and reached the magnesium metal. The above reaction generated hydrogen gas, which rose to the top of the burette(see diagram) and pushed down on the water. We waited for the gas to cool to room temperature.

2. *Why was the burette placed in a giant cylinder?*

This was done to equalize the atmospheric pressure(room pressure) with the gas pressure(P_{total}) inside the burette. At this point:

$P_{H2} + P_{H2O} = P_{total}$ or $P_{H2} = P_{total} - P_{H2O}$

3. *What is P_{H2O}?*

Whenever you collect gas over water, part of the water evaporates and contaminates the gas. To get the true gas pressure one has to consult a table to know how much of the total collisions in the container are actually due to the evaporated water. The warmer the water, the bigger the contribution.

4. *How did we get the real volume of hydrogen collected from the partial pressure P_{H2}?*

$V_{H2} = (P_{H2}/P_{total})(V_{measured})$, where $P_{total} = P_{room}$.

5. *Why did we use V = nRT/P afterwards?*

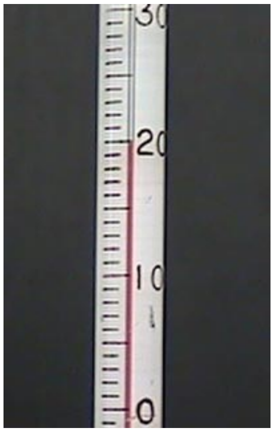
We did this to compare the theoretical volume with the volume obtained in #4

6. *Why did some of us get a yield that was greater than 100%?*

Since yield = $V_{actual}/V_{theoretical}$ *100% it meant that the actual volume measured was greater than it should have been due to the air bubble that was trapped in the burette at the beginning of the experiment.

3. Lab 1.3 Heat of Reaction

1. *What happened in the lab?*



For each of the three reactions we found the initial temperature of the aqueous environment, added NaOH in some form, stirred and recorded the maximum temperature.

2. *What were the three reactions involved in this lab?*

We calculated ***ΔH per mole*** (molar enthalpy) for three different reactions:

1. $NaOH_{(s)} \rightarrow NaOH_{(aq)}$ + heat

- Here we dissolved 2.00 g of NaOH in water and recorded the ΔT.
- Then we used Q = *m*cΔT, where *m* was the mass of the 100.0 mL of water absorbing the heat, not the mass of NaOH
- ΔH = -Q
- then we obtained ΔH/*n* by getting *n* from the molar conversion of 2.00 g of NaOH

2. $NaOH_{(s)} + HCl_{(aq)} \rightarrow H_2O + NaCl_{(aq)}$ + heat

- Here we dissolved 2.00 g of NaOH in 100 mL of dilute acid and recorded the ΔT.
- Then we used Q = *m*cΔT, where *m* was the mass of the 100.0 mL of water (which equaled about 100 g) absorbing the heat (once the 100.0 mL of dilute acid is neutralized it'll become saltwater containing about 100.g of water) , not the mass of NaOH
- ΔH = -Q
- then we obtained ΔH/*n* by getting *n* from the molar conversion of 2.00 g of NaOH

3. $NaOH_{(aq)} + HCl_{(aq)} \rightarrow H_2O + NaCl_{(aq)}$ + heat

- Here we dissolved 50 ml of aqueous NaOH(1.0M) in 50 mL of dilute acid and recorded the ΔT.
- Then we used Q = *m*cΔT, where *m* was the mass of the 50 mL of dilute acid added to 50 mL of dilute NaOH(**which created about 100 g of water**) absorbing the heat, not the mass of NaOH.
- ΔH = -Q
- then we obtained ΔH/*n* by getting *n* from using CV = moles of NaOH. We were given 1.0moles/L for $NaOH_{(aq)}$

3. What did Hess' Law have to do with this lab?

We used n = 0.05 (NaOH) for all three reactions allowing us to compare their enthalpies fairly.

Since equation (1) and (3) were the following:

(1) $NaOH_{(s)} \rightarrow NaOH_{(aq)}$ + heat ΔH_1

(3) $NaOH_{(aq)} + HCl_{(aq)} \rightarrow H_2O + NaCl_{(aq)}$ ΔH_3

If we combine them algebraically then we obtain:

$NaOH_{(s)} + HCl_{(aq)} \rightarrow H_2O + NaCl_{(aq)}$, but this is reaction (2) from our experiment!

So therefore $\Delta H_1 + \Delta H_3 = \Delta H_2$

Now since we had obtained these enthalpy values from temperature measurements and calculations (from $Q = mc\Delta T$, as shown above), we could check the validity of Hess Law to see if indeed $\Delta H_1 + \Delta H_3 = \Delta H_2$.

4. *Why were some of our results off a bit?*

- A small error in temperature reading could lead to a large difference in ΔH.
- Losing water in the transfer could lead to a larger temp. increase than expected.
- Losing heat to the air and cup could lead to a smaller temp. increase than expected.
- Absorption of water by NaOH while weighing the solid could have led to reaction and heat loss outside the cup.

5. *How was the teacher easily able to tell if there was a mistake in measurements and not in the calculations?*

$\Delta H_1 + \Delta H_3 = \Delta H_2$

But since $\Delta H = - mc\Delta T$,

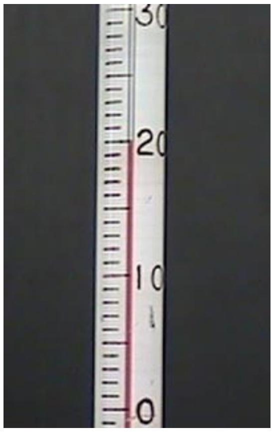
$- m_1c\Delta T_1 + - m_3c\Delta T_3 = -m_2c\Delta T_2$

But $m_1 = m_2 = m_3$ (in each case we had 100.0 ml of aqueous environment), so

$\Delta T_1 + \Delta T_3 = \Delta T_2$.

4. Measurements

1) The uncertainty is the smallest division divided by 2. In the case below, the smallest division is 1. The uncertainty becomes $\pm 1/2 = \pm 0.5$ So in this case any measurement has to be read to 1 decimal.



2) The measurement is $20.0 \pm 0.5\ ^oC$

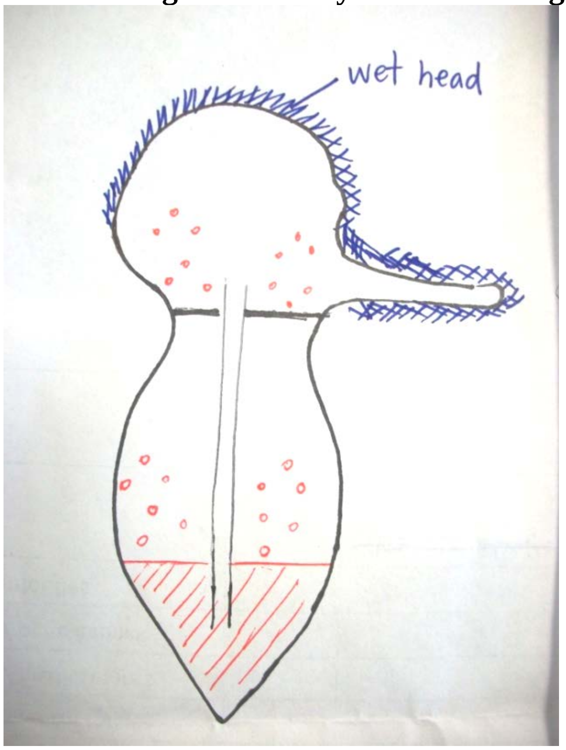
3) The error associated with the uncertainty becomes:

$$\frac{\pm uncertainty}{measurement} \times 100\% = error$$

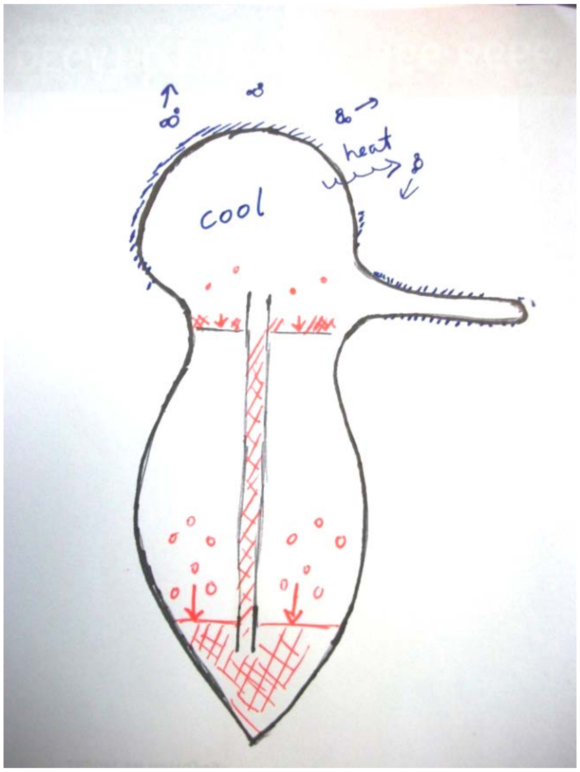
$$In\ the\ above\ case:$$

$\frac{\pm 0.5}{20.0} \times 100\% = 3\%$ (rounded from 2.5%)

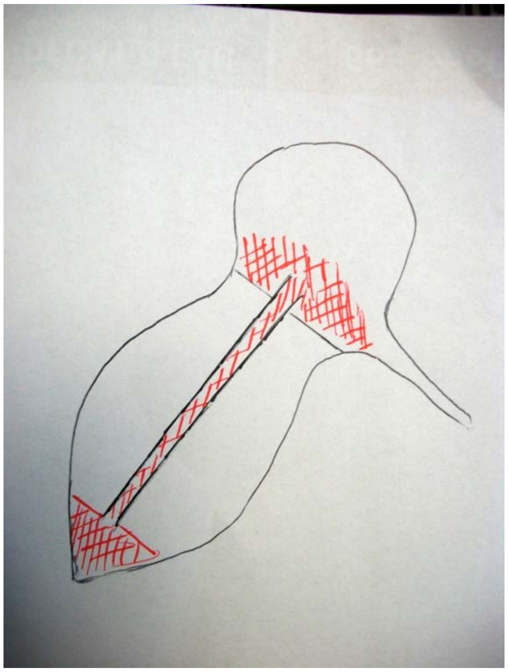
5. Explain the design and theory of the drinking bird.



There is equal pressure throughout the inside of the bird. Water has not evaporated yet. The pressure of the gas comes from the green liquid that evaporates easily.



(1)When water evaporates it steals heat from the environment. It's an endothermic process.
(2) Some of that heat will be lost by the gas inside the wet head.
(3) The heat loss will cause condensation and less gas pressure inside the bird's head
(4)This allows gas pressure from the lower body to push the liquid up through the column and into the head.



As the liquid moves into the head, it gets heavier and tips towards the glass of water, which is what got the head wet in the first place.