

## Surviving the December 2011 Lab Exam

This may seem a little long, but it's very complete! A good look at this is also a great way to review for the theory part of the exam.

The practical part of the exam will involve one of the factors affecting rates. Then there will be 5 additional questions based on the following labs. (1 each) As you can see there are several sample questions for each lab.

### Lab 1.1 The Reaction of Mg and HCl to Generate H<sub>2</sub> Gas

1. *What happened in the lab?*

- The reaction was:  $\text{Mg}_{(s)} + 2\text{HCl}_{(aq)} \rightarrow \text{MgCl}_{2(aq)} + \text{H}_{2(g)}$

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 ( $P_{\text{total}}$ ) inside the burette. At this point:

$$P_{\text{H}_2} + P_{\text{H}_2\text{O}} = P_{\text{total}} \quad \text{or} \quad P_{\text{H}_2} = P_{\text{total}} - P_{\text{H}_2\text{O}}$$

3. *What is  $P_{\text{H}_2\text{O}}$ ?*

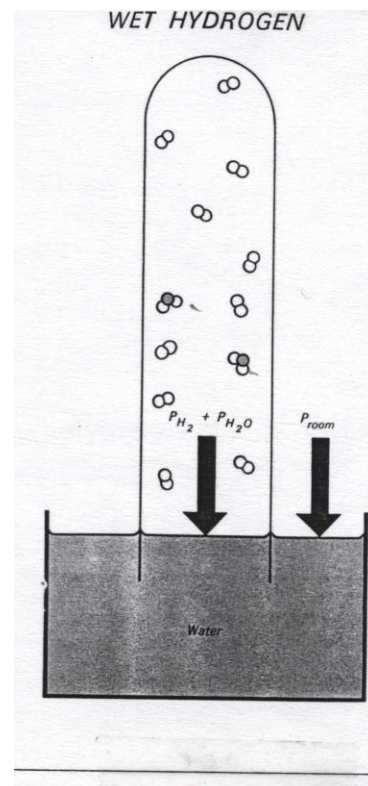
Whenever you collect gas over water, part of the water evaporates and contaminates the gas. To get the true gas pressure one has to consult a table to know how much of the total collisions in the container are actually due to the evaporated water. The warmer the water, the bigger the contribution.

4. *How did we get the real volume of hydrogen collected from the partial pressure  $P_{\text{H}_2}$ ?*

$$V_{\text{H}_2} = (P_{\text{H}_2}/P_{\text{total}}) (V_{\text{measured}}), \quad \text{where } P_{\text{total}} = P_{\text{room}}$$

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

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



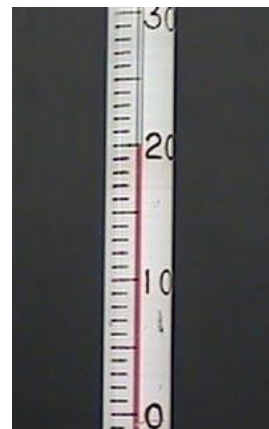
6. Why did some of us get a yield that was greater than 100%?

Since  $\text{yield} = V_{\text{actual}} / V_{\text{theoretical}} * 100\%$  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.

- **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  $\Delta H$  per mole (molar enthalpy) for three different reactions:



- Here we dissolved 2.00 g of NaOH in water and recorded the  $\Delta T$ .
- Then we used  $Q = mc\Delta T$ , where  $m$  was the mass of the 100.0 mL of water absorbing the heat, not the mass of NaOH
- $\Delta H = -Q$
- then we obtained  $\Delta H/n$  by getting  $n$  from the molar conversion of 2.00 g of NaOH



- Here we dissolved 2.00 g of NaOH in 100 mL of dilute acid and recorded the  $\Delta T$ .
- Then we used  $Q = mc\Delta T$ , 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
- $\Delta H = -Q$
- then we obtained  $\Delta H/n$  by getting  $n$  from the molar conversion of 2.00 g of NaOH



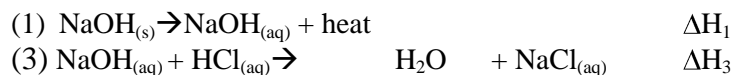
- Here we dissolved 50 ml of aqueous NaOH(1.0M) in 50 mL of dilute acid and recorded the  $\Delta T$ .

- Then we used  $Q = mc\Delta T$ , 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.
- $\Delta H = -Q$
- then we obtained  $\Delta H/n$  by getting  $n$  from using  $CV = \text{moles of NaOH}$ . We were given 1.0 moles/L for  $\text{NaOH}_{(aq)}$

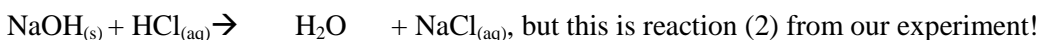
3. *What did Hess' Law have to do with this lab?*

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:



If we combine them algebraically then we obtain:



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  $\Delta H$ .
- 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.

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

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

## Lab 2.1 Energy in Chemical Reactions

### 1. What happened in the lab?

It resembled the previous lab, except that its purpose was different. We took calorimetric measurements for two different pure substances: LiCl and KNO<sub>3</sub>. A third measurement was taken for an unknown mixture of those same two substances. With a little algebra, we were then able to calculate the % of LiCl in the unknown. WOW!

### 2. How did we actually calculate the %?

Here's a summary of the calculations:

- The first part of the reaction revealed that the addition of LiCl to water is an exothermic reaction.
- The temperature change measured with a thermometer in a styrofoam cup was positive. Since the reaction released heat, the immediate environment (water) absorbs that heat, which is why  $mc\Delta T = Q$  was positive.
- We then expressed the Q per gram of LiCl which dissolved.
- The second part of the experiment revealed that the addition of KNO<sub>3</sub> to water is an endothermic reaction.
- The temperature change measured with a thermometer in a styrofoam cup was negative. Since the reaction absorbed heat, the immediate environment (water) lost energy, which is why Q was negative.
- We then expressed the Q per gram of KNO<sub>3</sub> which dissolved.
- The procedure was repeated for the unknown mixture consisting of LiCl(x) and KNO<sub>3</sub>(y)

Each Q per gram was multiplied by the mass (either x or y) of each solute in the mixture. The two expressions were added together, and their sum equaled  $mc\Delta T$  for the unknown.  $xW + yZ = mc\Delta T$ ,

where W = amount of heat from part 1

Z = amount of heat from part 2

m = mass of water

$\Delta T$  = temperature change caused by unknown

- Using a second equation in which the two unknown masses are known to add up to 1g:  $x + y = 1$ , we could solve for x and y

$$\% \text{ LiCl} = x/(x+y) * 100\%$$

(see sample problem in last year's exam)

## Lab 2.2 The Rate of Flow of Water

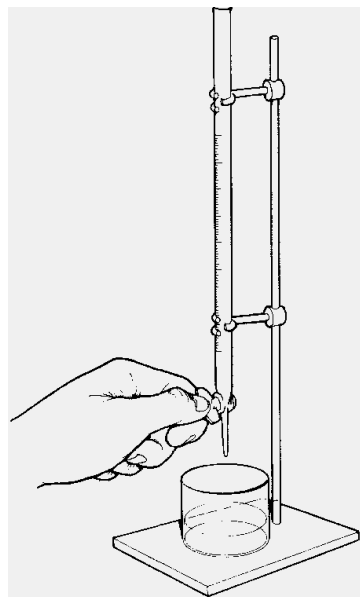


### 1. What happened in the lab?

We filled a burette with water and let it flow out while timing and measuring how much continuously actually flowed out every 10 seconds.

### 2. Why did the rate keep decreasing?

The weight of water above the outlet point kept decreasing. With less pressure from above, the rate decreased with every drop that flowed out. This situation mirrors the rates of chemical reactions, which keep slowing down because we keep losing reactants as they turn into products.



### 3. How do you obtain rates from the graph?

- A plot of volume versus time reveals a curve, not a straight line.
- The instantaneous rate, which is the rate at an exact moment in time, can be calculated by drawing a tangent line outside the curve. (see tangent at  $t = 180\text{s}$  in graph below.) The tangent must be drawn at the time on the graph that we are interested in. The instantaneous rate will be equal the slope of the tangent line.  $(35 - 50)/(100 - 250) = 0.10 \text{ ml/s}$
- The average rate on an interval ( $\Delta V/\Delta t$ ) from two points on the curve) will **not** necessarily equal the instantaneous rate just like your mark will not necessarily equal the class average.

ml of H<sub>2</sub>O

