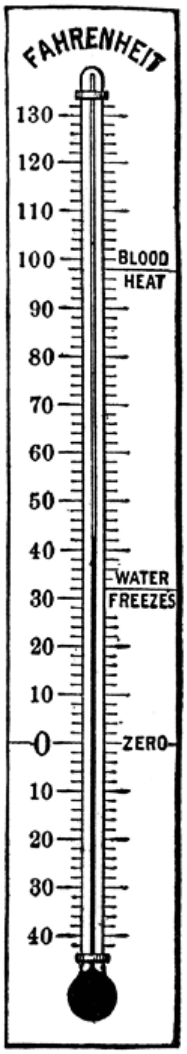
2012-05-19

**Chemistry Lab Exam Lab Review (keep scrolling down to page 6 for answers: recall that all four areas will be represented: (1) significant errors, measurements and lab tecjniques; (2) observations and interpretation of redox reactions; (3) collection of hydrogen gas—ideal gas law (4) Hess law lab)**

1. Your teacher squeezes the bulb to suck liquid into a 10.00 ml pipette. He lets the liquid rise to the brown line above the pipette's bulky part. He reads it at eye-level, making sure that the top of the meniscus matches the line. Then he lets it out, without forcing the small amount of liquid that seems to be stuck at the tip of the glassware. **What is he doing incorrectly**?

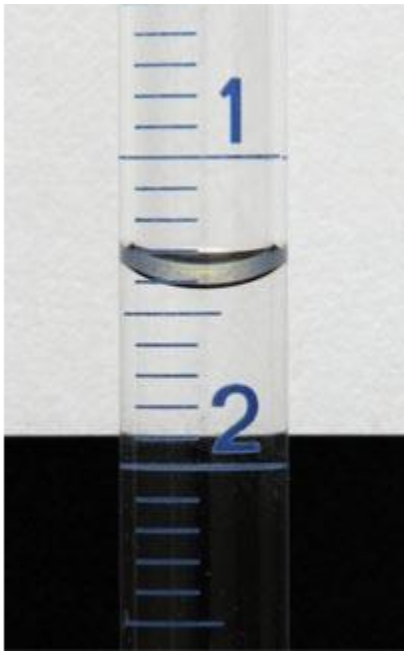
2. a. Your lab partner wants to measure the air temperature and is holding the thermometer. What could lead to an obvious error? (2 marks)



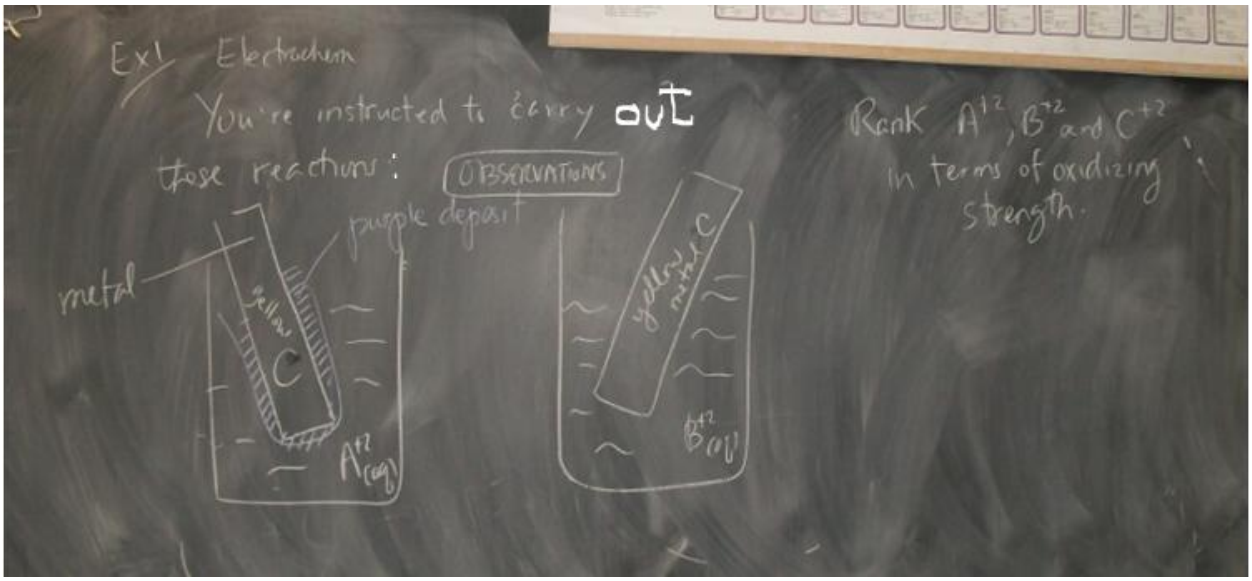

b. Is the uncertainty for this thermometer larger than usual? Why or why not?

c. Read the volumes with the correct number of decimal places.

d. What is the percent error associated with the first measurement if the uncertainty is $\pm 0.05\ ml$?



3.

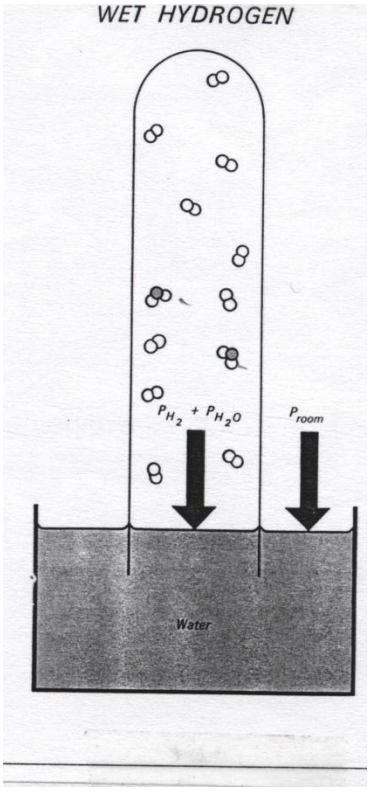



a) Rank $A^{2+}$, $B^{2+}$, $C^{2+}$ in terms of oxidizing strength.

b) What would have happened if you had placed metal plate B in $C^{2+}_{(aq)}$?

**4. Questions and Notes on two other labs you are responsible for:**

**a) Lab 1.1 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 $P_{total} = P_{room}$.

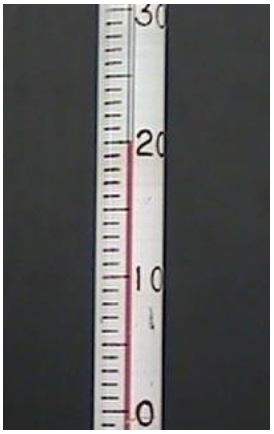
5. *Why did we use V = nRT/P afterwards?*

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.

**b) 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 $\Delta T$.
- 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

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 $\Delta T$.
- 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

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 $\Delta T$.
- 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 = 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 $\Delta H_1$

(3) $NaOH_{(aq)} + HCl_{(aq)} \rightarrow H_2O + NaCl_{(aq)}$ $\Delta H_3$

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 $\Delta H$.
- 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

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

**Answers to first three questions (the rest are within question 4).**

1. The mistake was in reading the top of the meniscus. The bottom has to be read. (He should not have forced out the extra bit of liquid. The instrument is calibrated to deliver the exact amount while excluding what is stuck at the narrow tip. )
2. A) There could be a transfer of heat from the hand to the thermometer, which will exaggerate the temperature reading.
   B) The uncertainty is larger than normal because they are using only 5 divisions(lines) for every 10 $^{o}$ on the scale.
   C) 1.40 ml and 1.42 ml. Notice that they must have the same number of decimal places.
   D) 0.05/1.40 * 100% = 4% (1 SF only in % due to 0.05's 1SF). Notice that the error is not too bad since the measurement itself did have three SF's.

3. In the first reaction the purple deposit came from the solution's $A^{2+}_{(aq)}$ which presumably stole electrons from the metal plate C to become neutral A. ( Neutral metals do not become negative, so C would not have further raised $A^{2+}_{(aq)}$' s charge!)

   So the reaction was: $C + A^{2+}_{(aq)} \rightarrow C^{+2}_{(aq)} + A$, meaning that $A^{2+}_{(aq)} > C^{+2}_{(aq)}$

   In the second beaker there is no reaction, meaning that $B^{2+}_{(aq)} < C^{2+}$

   Overall then, $A^{2+}_{(aq)} > C^{+2}_{(aq)} > B^{2+}_{(aq)}$