Lab 3: Energy and Rates Analysis of Chemical Reactions

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Experiment Date: November 6th 2008.

Lab 3:
Energy and Rates Analysis of Chemical Reactions

Due Date: November 17th 2008.
Instructor: Mr. MacLean

Lab 3: Energy and Rates Analysis of Chemical Reactions

Questions

(i) How does the molar enthalpy of reaction of magnesium vary with different acids, namely, hydrochloric, sulfuric, and acetic acids?

(ii) How does the rate of reaction of magnesium vary with these acids?

Hypothesis

If magnesium is mixed into any of the aqueous acid solutions, it is predicted that there will be fizzing and bubbling. This bubbling will be due to the release of Hydrogen gas, which is produced when the hydrogen bonds are broken and replaced with the magnesium.

It is hypothesized that the test tubes will be warm, because the reaction is exothermic; therefore it releases energy, or heat into its surroundings. It is also expected that the carboxylic acid (Ethanoic acid) will have a faster rate of consumption for magnesium, because its carbon base will allow it to let go of its hydrogen atoms more readily than hydrochloric and sulfuric acid.

The molar enthalpy for Magnesium should only vary because of human error. All three calculated enthalpies should be very close, because the molar enthalpy of Magnesium is dependant upon only Magnesium's properties, not the other reactants'.

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Materials
- 1 mol/L Aqueous Ethanoic acid
- 1 mol/L Aqueous Hydrochloric acid
- 1 mol/L Aqueous Sulfuric acid
- Magnesium shavings
- 3 calorimeters
- 3 stop watches
- Three 50mL graduated cylinders
- Three 100mL beakers
- Digital scale
- Safety glasses

Procedure

1. Apply safety glasses, and gather all apparatus.
2. Using the digital scale, measure out three samples of approximately ¼ g of magnesium. Be sure to record the exact mass of each sample.
3. Measure out 50mL of each aqueous acid using the graduated cylinders. Using the digital scale, measure and record the mass of each sample.
4. Transfer each sample into separate, labeled calorimeters.
5. Using the thermometer, measure and record the temperature of each acid sample.
6. Add one of the ¼ g samples of magnesium to each calorimeter, and immediately put the lid on the calorimeter on to close the system. Each person in the group will watch a calorimeter, and record the highest temperature that it reaches.
7. Pour the products down the drain of a functioning sink. Wash calorimeter and thermometers.
8. Repeat steps 2 and 3.
9. Transfer each sample of aqueous acid into a separate, labeled 100mL beaker. Ensure that each acid is at approximately the same temperature.
10. Prepare the three stopwatches.
11. Add a sample of Magnesium to each beaker (be sure to record the exact mass of magnesium being added to each), and simultaneously start the stopwatches.
12. When the magnesium is completely dissolved in the first solution, stop the first watch. When the magnesium in the second solution is completely consumed stop the second watch, and when the last one has no visible magnesium left, stop the final watch. Record the three times.
13. Again, dispose of the products in the drain of a sink, and wash all apparatus.

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Evidence

Enthalpy test:

Sample	Acid	Initial Temp. (°C)	Final Temp. (°C)	ΔTemp (°C)	Mass (g)
1	Ethanoic	24	38	14	50.75
2	Sulfuric	24	40	16	53.25
3	Hydrochloric	24	41	17	51.20

(Mg) Sample	Mass(g)	Moles
1	0.26	1.07E-2
2	0.25	1.03E-2
3	0.25	1.03E-2

Rate test:

(Mg) Sample	Acid	(Mg) Mass (g)	(Mg) Moles	ΔTime (s)
4	Ethanoic	0.28	1.15E-2	1513
5	Sulfuric	0.27	1.11E-2	210
6	Hydrochloric	0.26	1.07E-2	573

Analysis

Q: How does the molar enthalpy of reaction of magnesium vary with different acids, namely, hydrochloric, sulfuric, and acetic acids?

A: The three molar enthalpies were all relatively close. The difference between the enthalpies is most likely due to human error, or other complications during the experiment. Such as the inaccuracy of the scale, thermometer, calorimeter or any other apparatus used. A major source of error could have been using the specific heat capacity for water for each acid. Ideally the three molar enthalpies should be the same, because the molar enthalpy for magnesium is dependant upon magnesium, not the other reactant. (In Ethanoic acid: 2.8E2 kJ/mol, In Sulfuric acid: 3.5E2 kJ/mol, In Hydrochloric acid: 3.5E2 kJ/Mol)

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Q: How does the rate of reaction of magnesium vary with these acids?

A: The rate of reaction varies noticeably in the three acids. The rate was fastest in Sulfuric acid, at 5.29E-5 mol/s, second fastest in Hydrochloric acid at 1.87E-5 mol/s, and slowest in Ethanoic acid at 7.61E-6 mol/s. Since the surface area of the magnesium and the temperature of the acid was essentially the same in every sample, the only reason that the rates would vary this much is due to their different chemical properties.

Evaluation

As predicted before the experiment, there was a lot of fizzing and bubbling during the reaction. Also, as predicted, the system increased in temperature in all segments of the experiment, due to the exothermic nature of the reaction.

Contrary to the hypothesis, the Carboxylic acid had the slowest rate of reaction with magnesium. This could be due to the slight polarity of Ethanoic acid. Since magnesium is a single atom molecule, it is not polar; therefore it would dissolve more readily in a non-polar solution. This would account for its higher reaction rate with less polar hydrochloric acid and Sulfuric acid.

As hypothesized, the molar enthalpy for magnesium did not change a lot, and the differences are so minute that they can be accounted for by error. The molar enthalpy change should be the same for each segment because the acid is only being used to see how much heat is given off by magnesium. If the specific heat capacities were more accurate, the enthalpies would be closer to each other. Another source of error could be how well the thermometers were being read; digital thermometers would have given a more accurate reading. The calorimeters were not 100% closed systems. There could have been minor energy loss through the hole in the top, or the sides.

Synthesis

Before the investigation was carried out, it was predicted that Ethanoic acid would react faster than Hydrochloric acid and Sulfuric acid with the magnesium ribbon. After carrying out the lab, this was proven wrong. The carboxylic acid was actually the slowest of the three. As previously discussed, this could be due to the fact that it is more polar than Hydrochloric and Sulfuric acid. This could also be due to the more simplistic structures of Hydrochloric acid and Sulfuric acid, therefore there would be less intermediates and a more straightforward reaction.

The rate in the experiment was measured in respect to consumption of magnesium. This seems to be the easiest and most straightforward method of measuring the rate of the reaction. Another way that the rate could have been measured is in respect to production of hydrogen gas. The investigation would have been carried out the exact same, but this time after the reaction had finished, the timer would be stopped, and the remaining solution would be measured in terms of mass. This mass would be subtracted from the sum of the original mass of the aqueous acid and the mass of the magnesium. This difference would b the mass of the hydrogen gas produced. This number would then be divided by 2.02, the molar mass of hydrogen gas. The quotient would be the number of moles of hydrogen gas produced. By dividing this by the time the reaction took to finish, you would obtain the rate of production for hydrogen gas in mol/s.

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The more heat that is added to a system, the faster the molecules move, therefore there would be more molecules with the necessary amount of energy to create an activated complex or have a successful collision. Exothermic reactions produce heat energy, therefore the solution of acid and Magnesium would get warmer the more it reacted. Since the solution is getting hotter, the molecules are increasing in energy levels, and therefore the rate of the reaction is increasing. A way to control the exothermic nature of this reaction would be cooling it down as the reaction goes on. Adding more Freon or ice to the surrounding atmosphere to absorb the heat given off by the reaction. Not enough to cool down the solution, just enough to keep it at a stable temperature.

Something like this would be effective in providing heat in a hand warmer, but for multiple reasons, unsafe. In order to generate enough heat to keep somebody's hands warm in the winter, the acid would have to be pretty concentrated. This concentrated acid would be a hazard to the person if it ever leaked, caustic acid can cause irritation to skin and in some cases severe burns. Also there is a considerable amount of hydrogen gas produced in this reaction, which needs to go somewhere. The package with the acid and Magnesium in it would have to have ventilation for the gas to escape. This would be a place for the acid to leak out also. Acid and magnesium is not the ideal exothermic reaction to provide heat for hand warmers due to the risk of acid burns.

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Calculations