Chemistry 100
Fall 2008

Lab 7 - Heat and Temperature

Purpose:

- Observe the ***first law of thermodynamics***: that energy is conserved.
- Investigate the ***transfer of heat***.
- Determine ***the heat of fusion*** of ice
- Determine ***the heat of neutralization*** of acids and bases.

Theory:

1. **Heat** is another **form of energy**, and can be quantified with familiar units, ***joules.***

2. The **transfer of heat** can be calculated by measuring **a change in the temperature**,

3. The **heat capacity** of a material tells you **how much heat energy** is **required** to **cause a change in temperature**. The heat capacity depends on **how much the material** there is and is therefore referred to as the **specific heat capacity**.

Some **specific heat capacities** of materials:

water has a relatively high specific heat (4.184 J/°C*gram)

copper has a relatively low specific heat (0.378 J/°C *gram)

****To raise the temperature of water by 1 °C would take more than **11×** the energy required to raise the temperature of the same mass of copper by 1 °C.

4. A **calorie** is the **amount of heat** required to **raise the temperature of 1 gram** of **water** by **1°C**. The **specific heat of water** is equal to **1 calorie/°C*gram**.

****We will make volumetric measurements using graduated cylinders, so the **effective heat capacity** for all our experiments is **1 calorie/°C*mL. (density of water is 1g/mL)**

5. **Heat of fusion** - how much **heat** is required to **melt ice**; therefore the heat absorbed for the action of melting an amount of ice. **(calories/gram)**

6. **Heat of neutralization** – amount of heat **given off** in the reaction of an acid with a base.

$$HCl\ (aq) + NaOH\ (aq) \rightarrow NaCl\ (aq) + H_2O\ (l) + heat$$

IMPORTANT EQUATIONS!!!

Equation 1 (Eq. 1) $\Delta T = T_{final} - T_{initial}$ It can be positive or negative!!!!!

Eq. 2 Heat transferred (calories) = Heat capacity (1calorie/ °C mL) x vol (mL) x ΔT (°C)

*****The **quantity of heat** that we observe in each experiment **is** the **product** of the **heat capacity, volume, and the change in temperature**.

TWO IMPORTANT POINTS TO REMEMBER:

1. The thermometers used in this lab only read the temperature of a liquid when they are in the liquid. When removed from the liquid, they read the temperature of the air.

2. **Always hold on to the thermometer** (unless it is lying on the bench top). **The Styrofoam cups are not heavy enough to prevent the thermometer and cup from falling over**.

In this experiment we consider heat gain and heat loss in a single container. Since **energy is conserved**, the heat gained by one fraction of the total volume is equal to the heat lost by the other fraction of the total volume.

Procedure 1: Record your data below!

1. Pour 50 mL of cold water from the tap into a graduated cylinder.
2. Transfer it into a Styrofoam cup and record its temperature* below.
3. Run "hot" water from the tap until it is actually hot.[†]
4. Measure 50 mL of hot water from the tap in the graduated cylinder.
5. Record the temperature* of the hot water.
6. **Pour the hot water into the Styrofoam cup with the cold water**.
7. Record the final temperature.
8. Calculate the heats involved.

* Allow the temperature to equilibrate.
[†] You may want to run the hot water into the graduated cylinder such that the walls of the cylinder are the same temperature as the water.

Heat Loss		Heat Gain	
Volume hot water	________	Volume cold water	________
Initial Temperature	________	Initial Temperature	________
Final Temperature	________	Final Temperature	________
ΔT (Eq. 1)	________	ΔT (Eq. 1)	________
Heat (Eq. 2)	________	Heat (Eq. 2)	________

(Include units in all entries)

REMINDER!!!!

Eq. 2 Heat transferred (calories) = Heat capacity (1calorie/ °C mL) x vol (mL) x ΔT (°C)

9. How does the heat loss compare to the heat gain?

Procedure 2:

Repeat the previous procedure, except pour **<u>50 mL of hot water into 100 mL of cold water</u>**. Make a prediction of the final temperature before mixing the hot and cold water.

Record your data below.

Heat Loss		Heat Gain	
Volume hot water	_________	Volume cold water	_________
Initial Temperature	_________	Initial Temperature	_________

Before you do the experiment, predict the final temperature: _____________

Final Temperature	_________	Final Temperature	_________
ΔT (Eq. 1)	_________	ΔT (Eq. 1)	_________
Heat (Eq. 2)	_________	Heat (Eq. 2)	_________

1. Is the heat calculated in the **heat loss** column **positive** or **negative**?

2. Is the heat calculated in the **heat gain** column **positive** or **negative**?

3. Is the **1st law of thermodynamics obeyed**? What is the *total* heat transferred (heat gained +heat lost)?

Procedure 3:

Repeat the previous, except pour **<u>50 mL of cold water into 100 mL of hot water</u>**. Make a prediction of the final temperature before mixing the hot and cold water.

Record your data below.

Heat Loss		Heat Gain	
Volume hot water	_________	Volume cold water	_________
Initial Temperature	_________	Initial Temperature	_________

Before you do the experiment, predict the final temperature: _____________

Final Temperature	_________	Final Temperature	_________
ΔT (Eq. 1)	_________	ΔT (Eq. 1)	_________
Heat (Eq. 2)	_________	Heat (Eq. 2)	_________

Procedure 4 **(Heat of fusion of ice):** Record your data below.

In this part of the experiment we will determine the heat of fusion of ice. In other words, you will determine how much heat is required to melt ice.

1. Pour 100. mL of room temperature (about 20^oC) water into a styrofoam cup.
2. Record all of your data below.
3. Slowly add ice, monitoring the temperature, and stir until the ice no longer melts.
4. **You want as small of an excess of ice as possible**
5. The final temperature should be approximately 2^oC.
6. Pour the volume of water from the styrofoam cup into a graduated cylinder, and measure the final volume.
7. Calculate the volume and mass of ice that was added.

In an insulated system such as this, **no heat is lost or gained to the surroundings**, so the heat given off by the water must be absorbed by the ice. The heat of fusion is the number of calories needed to melt 1.00 gram of ice.

HEAT LOSS

Initial volume ___________

Initial Temperature ___________

Final Temperature ___________

Temperature Change ___________

HEAT GAIN

Heat given off (calories)___________ (Equals) Heat absorbed ___________

Final Volume ___________

(Initial Volume) ___________

Volume Change ($V_f – V_i$) _________

(The density of water is 1gram/mL.) ---------------------------→Mass of ice melted ___________

8. Using your data, calculate the heat of fusion of ice (the units will be calories/gram) :

Heat of Fusion = heat absorbed / mass melted =

9. The accepted literature value for the heat of fusion of ice is 80. calories/gram. Is your value greater or less than the literature value? Why is it different?

Procedure 5 (Heat of reaction, strong acid): Record your data below

In this part we will investigate the heat released in a chemical reaction.

$$HCl\ (aq) + NaOH\ (aq) \rightarrow NaCl\ (aq) + H_2O\ (l) + heat$$

Heat is evolved in this acid/base neutralization, and heat is considered a product of the reaction along with the dissolved salt ions (aqueous NaCl), and liquid water (H_2O).

1. Pour 25. mL of 1.0 M HCl into one styrofoam cup (1.0 M is a concentration [moles/liter]).
2. Pour 25. mL of 1.0 M NaOH into the graduated cylinder
3. Record the temperature of these solutions below. Both are initially at room temperature.
4. **Mix the two solutions** and record the final temperature.
5. Note that the final volume is **50. mL**. This is the volume of water that is heated by the chemical reaction.

Remember, Heat = Effective heat capacity x Volume of solution x ΔT.

Initial Temperature __________

Final Temperature __________

ΔT (T_f-T_i) __________

Heat (calories) ___________

6. In this experiment you made about 0.45 g of water. **How many calories are released per gram of water?**

$$\left[\frac{heat(calories)_{your\_measurement}}{[0.45grams]}\right] = \left[\frac{??calories}{gram}\right] = ?$$

7. What would be the **heat evolved** in a **scaled up reaction** that produces **10. grams** of water?

8. What is the **molar heat of reaction** (the amount of heat released per mole of water produced)? The molar mass (molecular weight) of water is necessary in the calculation (as shown below).

$$\left[\frac{heat(calories)_{your\_measurement}}{[0.45grams]}\right] \times \left[\frac{?grams}{mole}\right]_{MW(water)} = \left[\frac{??calories}{mole}\right] = ?$$

9. What would be the **heat evolved** in a **scaled up reaction** that produces 10. moles of water?

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Procedure 6 (Heat of reaction, weak acid): Record your data below

Repeat the previous experiment but use 1.0 M acetic acid, $C_2H_3O_2H$, and 1.0 M NaOH. Acetic acid is a weak acid, in contrast to HCl, the strong acid used in procedure 5.

$$C_2H_3O_2H\ (aq) + NaOH\ (aq) \rightarrow C_2H_3O_2Na\ (aq) + H_2O + heat$$

Initial Temperature __________

Final Temperature __________

ΔT (T_f-T_i) __________

Heat __________

1. As in procedure 5, you made about 0.45 g of water that is in addition to the 50.0 mL. **How many calories are released per gram of water?**

$$\left[\frac{heat(calories)_{your\_measurement}}{[0.45 grams]}\right] = \left[\frac{??\, calories}{gram}\right] = ?$$

2. What would be the **heat evolved** in a **scaled up reaction** that produces **10. grams** of water?

3. What is the **molar heat of reaction** (the amount of heat released per mole of water produced)? The molar mass (molecular weight) of water is necessary in the calculation (as shown below).

$$\left[\frac{heat(calories)_{your\_measurement}}{[0.45 grams]}\right] \times \left[\frac{?\, grams}{mole}\right]_{MW(water)} = \left[\frac{??\, calories}{mole}\right] = ?$$

4. What would be the **heat evolved** in a **scaled up reaction** that produces **10. moles** of water?

5. How do the **molar heats of reaction compare** for **strong and weak acids** (compare the results of procedures 5 and 6)?

6. Procedures 5 and 6 were examples of **exothermic reactions**. What **was the sign (+ or -) of ΔT for these reactions?**

7. **What sign of ΔT** would you expect for **endothermic reactions**?