

Calorimetry Measurements of Fusion, Hydration and Neutralization - Hess' Law

Prepared by Edward L. Brown, Lee University

To become familiar with energy changes and the use of a calorimeter to measure these changes and to explore the mathematical connections between mass, specific heat, ΔT and energy.

OBJECTIVE

APPARATUS AND CHEMICALS

A P P A R A T U S

Beaker to support the calorimeter

MeasureNet Temperature Probe

Styrofoam calorimeter and lid

C H E M I C A L S

Sodium Hydroxide (granules)

1.2 M HCl

Ice

Thermodynamics is the study of energy and its movement from one object to another object. When a chemical reaction is involved in the production or absorption of energy, the field of study is called **thermochemistry**. Experiments investigating both types of energy transfers are performed in the typical general chemistry lab. The amount of heat transferred between objects, or from reactants to their surroundings, is related to the mass (typically in grams) of the objects or the moles of the reactants. This “connection” allows one to determine energy from the amount of a substance. The SI unit for energy is the joule, J, named after James Joule, a British scientist devoted to the investigation of work and heat. If you recall the kinetic energy equation, $E = \frac{1}{2}mv^2$, 1 J of energy is possessed by a 2 kg object moving at a speed of 1 m/s. [$\therefore 1 \text{ J} = \text{kg}\cdot\text{m}^2/\text{s}^2$] A non-SI energy unit still widely used in the sciences is the calorie. A **calorie** (cal) is defined as the amount of energy needed to raise the temperature of 1 g of water 1° C (specifically from 14.5° C to 15.5° C). Its relationship to the joule is defined (not measured) so significant figures are not a concern when using the following conversion factor:

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Calorimetry Measurements of Fusion, Hydration and Neutralization

$$4.184 \text{ J} = 1 \text{ cal (exactly)}$$

You have probably “counted calories” in your diet; the nutritional Calorie (Cal) is equal to 1000 cal (or 1 kcal). One Calorie in your diet will supply you with enough energy to raise the temperature of 1000 g (1 L) of water 1° C. The energy stored in foods is converted into heat and work. Food molecules have **potential energy** stored in their bonds. The bonds in food molecules are weaker and higher in energy than the bonds in CO₂ and H₂O (the major products of metabolism). This **potential energy** is used by the body to maintain a temperature of 37° C (heat transfer) and to do work (simple movement or strenuous labor).

A. Heat Transfers

Energy is transferred as heat between two objects of *differing temperatures*. Heat flows out of the hotter object and into the colder object. Since energy is conserved according to the **First Law of Thermodynamics**, **Equation 1** can be written to mathematically describe the transfer of heat:

$$\text{- Heat lost by the hotter object} = \text{Heat gained by the colder object(s)}$$

The **negative sign** is needed because these are two *opposite* processes. The heat gained or lost by an object is dependent on three properties of the object: its mass, the temperature change experienced by the object, and the object’s intrinsic ability to absorb heat. We are all aware of the two different sensations experienced by our hand when we hold hot coffee in a Styrofoam cup (no handle) or a ceramic coffee mug (we need the handle). Styrofoam and ceramic have very different intrinsic heat-absorbing properties. Like Styrofoam and ceramic, all substances have a measurable intrinsic heat-absorbing property that can be represented by any of the following:

1. **Heat Capacity** – the temperature change experienced by “**an object**” when it absorbs a known amount of energy. The “**object**” can be an undefined amount of a pure substance, or a complex apparatus, called a calorimeter, used to measure the amount of heat transferred or absorbed during a process. You can recognize the heat capacity in a problem because it has units of **J/°C** and **it must be experimentally determined** for a defined “**object**”. The next two intrinsic heat-absorbing properties pertain to **pure substances**.

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2. **Molar Heat Capacity** – the energy needed to raise the temperature of 1 mole of a pure substance 1° C. You can recognize molar heat capacity in a problem because it has units of **J/mole·°C**.
3. **Specific Heat Capacity** (specific heat) - the energy needed to raise the temperature of 1 gram of a pure substance 1° C. You can recognize specific heat in a problem because it has units of **J/g·°C**. The specific heat of liquid water is 4.184 J/g·°C and must be memorized. All three phases (solid, liquid and gas phase) of a pure substance will have different specific heats. It is critical that you use the specific heat of the appropriate phase when performing energy calculations.

The mathematical relationship between these three intensive property measurements and heat is straightforward:

- Heat Capacity (J/°C) × ΔT (in °C or K) = Heat (in J)
- Molar Heat Capacity (J/mole·°C) × # moles × ΔT (in °C or K) = Heat (in J)
- Specific Heat (J/g·°C) × mass (in g) × ΔT (in °C or K) = Heat (in J)

If more than two substances are involved in the Heat Transfer, the First Law of Thermodynamics can be expanded to

$$\text{- Heat Lost By Substance A} = (\sum \text{Heat Gained By Substance B, Substance C, etc})$$

In today's lab, the heat lost by Substance A will be the Heat of Hydration (dissolving a solid) and the Heat of Reaction (an acid / base neutralization). This heat will be transferred to the surrounding water and calorimeter causing their temperature to increase. This gives the following specific equations:

$$\text{- Heat Lost By Hydration} = \text{Heat Gained By Water} + \text{Heat Gained By Calorimeter}$$

$$\text{- Heat Lost During Neutralization} = \text{Heat Gained By Water} + \text{Heat Gained By Calorimeter}$$

Experimentally, you will use a simple calorimeter constructed of nested Styrofoam cups, a Styrofoam lid and a temperature probe. Although Styrofoam is a good insulator, some heat will be absorbed by the calorimeter and some will be released to the surroundings. In actuality, this heat is small compared to the heat absorbed by the solution in the

calorimeter. **For this experiment, we will consider this heat loss to be negligible – the “Heat Gained By Calorimeter” will be dropped from the equations above.**

B. Heats of Reaction

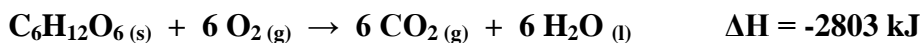
When reactants come together in a chemical reaction to form products, chemical bonds are broken (in the reactants) and formed (in the products). The breaking of **ANY** chemical bond requires the input of energy which has been assigned a “+” value. Imagine two magnets with North / South poles: pulling the two magnets apart requires an energy input from your muscles; however, placing the two magnets in close proximity to each other and then removing any restraining force, will result in a magnetic “bond” between the two magnets. Confusion can arise when biologists and biochemists refer to **high-energy bonds** (i.e. ATP) or “**getting energy**” by **breaking** bonds. High-energy bonds are very weak and easily broken with a minimal input of energy (a small + value). It is because the **products** of the reaction contain stable, low-energy bonds that allows one to “get energy” from the **overall** reaction. The heat given off or absorbed during a reaction is called the enthalpy of the reaction and is written as, ΔH_{rxn} . Exothermic reactions have a negative enthalpy change, while endothermic reactions have a positive enthalpy change. Several things of note concerning thermochemical equations (balanced chemical reactions with the associated ΔH_{rxn}):

- The $\Delta H_{\text{rxn}} = H_{\text{products}} - H_{\text{reactants}}$ (Enthalpies are found in tables at the end of most chemistry text books.
- The ΔH_{rxn} is given in units of J (or kJ) per mole. Therefore, the heat of the reaction depends on the amount of reactants involved.
- The $\Delta H_{\text{reverse rxn}}$ is equal to $-\Delta H_{\text{rxn}}$
- The $H_{\text{reactants}}$ and H_{products} are phase dependent. If a substance reacts as a gas, its enthalpy is different than if it reacts as a liquid or solid. **Make sure you note the phase of the substance when using enthalpy tables.**

Several types of reactions have historically been used to measure enthalpies of reaction due to the large evolution of heat that is subsequently trapped and measured in known quantities of water.

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- The most obvious are the **combustion reactions**. Oxygen combines with many elements and compounds to form tremendous amounts of energy. This energy, if harnessed, can be converted into work. Most of the energy used by man is produced from the combustion of fuels like coal, natural gas, propane, gasoline, diesel, etc. These “fossil fuels” contain C and/or H and/or O and produce carbon dioxide and water when burned. In an advanced chemistry laboratory you may use a “bomb calorimeter” to study the heat produced in a particular combustion reaction. In an experiment of this type, a metal vessel called a “bomb” is charged with a combustible substance and pressurized with oxygen. The “bomb” is surrounded by water in an insulated container. An electrical current starts the reaction and the heat from the combustion causes an increase in the temperature of the surrounding water. Calculations are performed which allow for the determination of the enthalpy of combustion. It is somewhat ironic that the human body serves as the calorimeter for the “**controlled**” **combustion reaction**:



Of the 2803 kJ of energy produced from each mole of glucose, only ~ 40% is used for muscle contraction; the remaining 60% is used to maintain body temperature at 37 °C year round.

- **Neutralization reactions** also generate significant amounts of heat and have been exploited in the general chemistry laboratory. Since the acid and base are presented to each other as aqueous solutions, the heat emitted when the salt and water are formed is absorbed by the surrounding water molecules. This causes an increase in temperature that provides the heat of the reaction when multiplied by the specific heat and mass of the solution. In some experiments the heat capacity of the calorimeter is measured and becomes part of the equation.
- The energy associated with hydrating (dissolving in water) a solute can be endo- or exothermic. Again, since water molecules surround the solute, any heat absorbed in the **hydration** process will come from the water. A corresponding decrease in temperature will be noted. Likewise, any heat emitted will raise the temperature of the surrounding water molecules. This change in temperature is used to calculate the enthalpy of hydration when multiplied by the specific heat and mass of the solution.

C. Heat of Fusion

Phase changes present additional opportunities to calculate heat transfers and specifically **heat transfers that do not cause a temperature change**. The energy needed to melt (or freeze) 1 mole of a pure substance when the pressure is 1 atm is called the **heat of fusion**. The energy associated with this solid \leftrightarrow liquid phase change has units of **kJ/mole** (note the absence of a temperature unit) and can be calculated from the following equation:

$$\text{Heat absorbed (melting) or released (freezing)} = \text{Heat of Fusion} \times \# \text{ moles}$$

Like melting (or freezing), vaporization (or condensation) also occurs at a defined temperature and the energy requirements at that temperature are described by the **heat of vaporization (kJ/mole)** of 1 mole of pure substance. The heat absorbed by a liquid going to a gas or the heat released when a gas condenses to a liquid is calculated from an equation similar to the one above where the heat of fusion is replaced with the heat of vaporization.

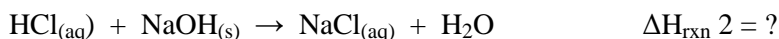
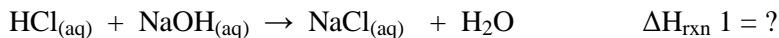
D. Hess' Law

We can measure the heat in the above reactions by “trapping” it in a pure substance with a known specific heat (i.e. water). Because there are countless known chemical reactions, it would be impractical to set up an experiment to measure each and every reaction. Fortunately, we can generate millions of values for the ΔH_{rxn} from tables containing only a few hundred entries. Thanks go to Germain Henri Hess (1802 - 1850) for his thermochemical studies, which led him to discover an empirical law based on assigning **state function** status to enthalpy. If a thermochemical property is determined to be a state function, then only the quantity of the initial and final states are used in computations. **Hess' law** states that the heat evolved or absorbed in a chemical process is the same whether the process takes place in one step ($A \rightarrow D$) or in several steps ($A \rightarrow B \rightarrow C \rightarrow D$).

In today's lab, the heat of hydration of sodium hydroxide and the heat of reaction between sodium hydroxide and hydrochloric acid will be investigated separately, prior to being coupled to provide proof of Hess' law. This coupled process should produce an energy change that is equal to the sum of the energies from the two individual processes.

Calorimetry Measurements of Fusion, Hydration and Neutralization

Measuring the energy produced or consumed by a reaction can be as important as analyzing the reactants and products. In the lab today, you will investigate the three reactions below:



Notice that the above listed reaction 1 and reaction 3 together can be considered as alternative steps to reaction 2. If we can determine the ΔH of reactions 1 and 3 per mole of NaOH present, we can calculate the $\Delta H_{\text{rxn } 2}$:

$$\left(\frac{\Delta H_{\text{rxn } 1} (\text{J})}{\text{moles NaOH}} \right) (\text{moles NaOH}) + \left(\frac{\Delta H_{\text{rxn } 3} (\text{J})}{\text{moles NaOH}} \right) (\text{moles NaOH}) \text{ should equal } \Delta H_{\text{rxn } 2} (\text{J})$$

This calculated value will then be compared to an experimentally determined value as a proof of Hess' Law.

PROCEDURE

Heat of Hydration of NaOH

1. Obtain a dry Styrofoam calorimeter with a Styrofoam cap [Figure 1]. Place it in an appropriately sized beaker for support.
2. Take the dry calorimeter and its cap to the balance and determine its mass to the nearest 0.001 g [Data Sheet Q1].
3. Obtain 50 mL \pm 5 mL **tap water** in a graduated cylinder. Pour this water in the calorimeter and snugly replace the Styrofoam cap. Determine the mass of the calorimeter and water to the nearest 0.001 g [Data Sheet Q2].
4. Determine the mass of the water [Online Report Sheet Q3].
5. Place the calorimeter in its beaker and insert the MeasureNet temperature probe. To record the temperature of a solution using the temperature probe, press the **MAIN MENU** button, followed by **F2** and **F1**. Press the **DISPLAY** button to view

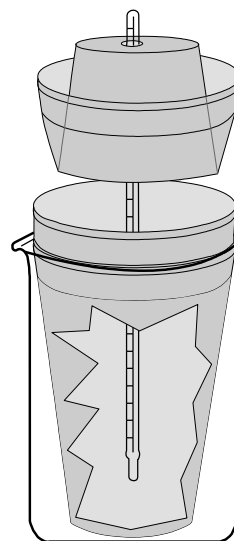


Figure 1

Calorimetry Measurements of Fusion, Hydration and Neutralization

the temperature or press the **START/STOP** button to view the temperature over time on the built-in graph.

6. Allow the water and calorimeter to equilibrate for 3 minutes and record the temperature (to the nearest 0.01 °C) [**Data Sheet Q4**].
7. One member of your team will handle the solid NaOH and will need to *put on gloves and wear them for the remainder of the lab*.
8. Take the calorimeter, its cap and your vial of solid NaOH to the balance area. The person wearing the gloves will add $\sim 2 \text{ g} \pm 0.2 \text{ g}$ solid NaOH to the calorimeter by *following the directions below*.

- ✓ **NaOH absorbs water from the air, so this part must be done quickly!!**
- ✓ Tare the balance.
- ✓ With the cap on the NaOH vial, record its mass [**Data Sheet Q5**].
- ✓ Remove the NaOH vial and tare the balance.
- ✓ Place the calorimeter (without its cap) on the balance and **Tare the balance (it should read 0.000 g)**.
- ✓ Open the vial containing the solid NaOH and add $2 \text{ g} \pm 0.2 \text{ g}$ to the calorimeter. **If the NaOH is in pellet form, pour the pellets into your GLOVED hand and add one pellet at a time. CAP THE VIAL IMMEDIATELY WHEN FINISHED!!**
- ✓ Quickly place the calorimeter cap snugly in place to minimize heat loss to the air – have your partner take the calorimeter back to your lab station and insert the MeasureNet temperature probe. **They will proceed to Step 9 while you complete the following steps.**
- ✓ Place the **CAPPED** NaOH vial on the balance and record its mass [**Data Sheet Q6**].
- ✓ Determine the mass of the added NaOH [**Online Report Sheet Q7**].
- ✓ **NaOH is very caustic – avoid contact!!** If suspicious of contact, wash affected area immediately with plenty of water.

9. Continuously swirl the contents of the calorimeter (by swirling the beaker).
10. Record the maximum temperature (to the nearest 0.01 °C) reached in the calorimeter [**Data Sheet Q8**].
11. Determine the ΔT of the solution during the hydration of solid NaOH [**Online Report Sheet Q9**].

12. **Pour the calorimeter contents into a clean beaker AND SET IT ASIDE!! – you will use it in Step 36.** Continue with the next part of the experiment (Step 16) while this solution cools to room temperature.
13. Determine the mass of the NaOH solution [**Online Report Sheet Q10**].
14. Determine the heat (**in J**) absorbed by the NaOH solution in the calorimeter [**Online Report Sheet Q11**]. Assume the specific heat of the solution is the same as pure water and use the mass in **Online Report Sheet Q10** as the mass of the solution in this calculation.
15. Determine the heat of hydration of NaOH (J/mole NaOH) [**Online Report Sheet Q12**].

Heat of Fusion of Water

16. Rinse the calorimeter with water, dry it completely and determine its mass with the cap attached [**Data Sheet Q13**]. This will be slightly different than [**Data Sheet Q1**] since the Styrofoam will trap a little water from the first experiment.
17. Place 100 ± 5 mL of hot water (temperature = 50 ± 5 °C) in the Styrofoam calorimeter and snugly affix the Styrofoam cap.
18. Record the mass of the calorimeter, cap, and water [**Data Sheet Q14**].
19. Determine the mass of the water [**Online Report Sheet Q15**].
20. Place the calorimeter in its beaker and insert the MeasureNet temperature probe into the calorimeter.
21. At the back of the class is an ice machine containing ice that is at 0 °C (you will get 30 – 40 grams **in Step 22**). Go ahead and record 0 °C as the temperature of the ice. [**Data Sheet Q16**]. **Perform Step 22 as quickly as possible.**
22. Use a 100 mL beaker to:
 - tare the 100 mL beaker and add ice until the balance reads a number between 30 and 40 g. **Quickly** take the ice to your lab bench;
 - as you arrive your partner should record the temperature of the warm water to the nearest 0.01 °C [**Data Sheet Q17**]and
 - **quickly** remove the calorimeter cap / MeasureNet probe;
 - **quickly** add the ice – **it is fine if a few pieces remain in the beaker;**
 - recap the calorimeter
23. Gentle swirl the beaker holding the calorimeter and record the lowest temperature reached by the system (to the nearest 0.01 °C) once all the ice has melted [**Data Sheet Q18**].

Calorimetry Measurements of Fusion, Hydration and Neutralization

24. Remove the temperature probe. Place the calorimeter and its contents on the balance and determine the total mass [**Data Sheet Q19**].
25. Determine the mass of the ice [**Online Report Sheet Q20**].
26. Determine the heat (in J) lost by the warm water as it cooled to the final temperature recorded in **Data Sheet Q18** [**Online Report Sheet Q21**].
27. Determine the temperature change experienced by the H₂O that was introduced into the calorimeter as ice which then melted into water [**Online Report Sheet Q22**].
28. Determine the heat (in J) gained by the water (formerly the ice) as it warms from 0 °C to the final temperature recorded in **Data Sheet Q18** [**Online Report Sheet Q23**]. Remember to use only the mass of the ice (now water) in this calculation.
29. Using **Equation 1**, determine the Heat of Fusion of Water in J/g [**Online Report Sheet Q24**] and in kJ/mole [**Online Report Sheet Q25**].
30. The accepted value for the Heat of Fusion of water is 6.008 kJ/mole. Determine the % Error of your experimental data [**Online Report Sheet Q26**].

$$\% \text{ Error} = \frac{|6.008 - Q25|}{6.008} \times 100$$

Heat of Neutralization (NaOH + HCl)

31. Empty the contents of the calorimeter and dry the calorimeter with paper towels.
32. Determine the mass of the calorimeter with the cap attached [**Data Sheet Q27**].
33. Place ~ 60 ± 5 mL of 1.2 M HCl in the calorimeter (use a graduated cylinder). **Record the mass of the calorimeter, cap and HCl solution** [**Data Sheet Q28**].
34. Determine the mass of the HCl solution [**Online Report Sheet Q29**].
35. Place a temperature probe in the solution and after 3 minutes record the temperature (to the nearest 0.01 °C) [**Data Sheet Q30**].
36. Remove the temperature probe from the calorimeter, dry it completely with a paper towel and place it in the beaker containing the NaOH solution **from Step 12**. When the NaOH solution is within 0.5 °C of the temperature recorded for the HCl solution, add the NaOH solution to the HCl solution in the calorimeter and snugly replace the temperature probe and Styrofoam cap.
37. Determine the mass of the neutralization solution (the HCl solution plus the NaOH solution) [**Online Report Sheet Q31**].
38. Swirl the contents of the calorimeter and record the maximum temperature (to the nearest 0.01 °C) reached during this neutralization reaction [**Data Sheet Q32**].

Calorimetry Measurements of Fusion, Hydration and Neutralization

39. Determine change in temperature [**Online Report Sheet Q33**].
40. Determine the heat (in J) absorbed by the solution [**Online Report Sheet Q34**]. Assume the solution in the calorimeter has the same specific heat as pure water.
41. Determine the $\Delta H_{\text{neutralization}}$ in J/g NaOH [**Online Report Sheet Q35**] and in J/mole NaOH [**Online Report Sheet Q36**].

Hess' Law: Heat of Hydration + Heat of Neutralization

42. Pour the contents of the calorimeter down the sink with water and dry the calorimeter with paper towels.
43. Determine the mass of the calorimeter with the cap attached [**Data Sheet Q37**].
44. Place $\sim 60 \pm 5$ mL of 1.2 M HCl in the calorimeter (**use a graduated cylinder**). **Record the mass of the calorimeter, cap and HCl solution** [**Data Sheet Q38**].
45. Determine the mass of the HCl solution [**Online Report Sheet Q39**].
46. Place a temperature probe in the solution and after 3 minutes record the temperature (to the nearest 0.01 °C) [**Data Sheet Q40**].
47. Take the calorimeter and its cap to the balance where you will add the solid NaOH to the calorimeter by *following the directions below*..

- ✓ **NaOH absorbs water from the air, so this part must be done quickly!!**
- ✓ Tare the balance.
- ✓ With the cap on the NaOH vial, record its mass [**Data Sheet Q41**].
- ✓ Remove the NaOH vial and tare the balance.
- ✓ Place the calorimeter (without its cap) on the balance and **Tare the balance (it should read 0.000 g)**.
- ✓ Open the vial containing the solid NaOH and add $2 \text{ g} \pm 0.2 \text{ g}$ to the calorimeter. **If the NaOH is in pellet form, pour the pellets into your GLOVED hand and add one pellet at a time. CAP THE VIAL IMMEDIATELY WHEN FINISHED!!**
- ✓ Quickly place the calorimeter cap snugly in place to minimize heat loss to the air – have your partner take the calorimeter back to your lab station and insert the MeasureNet temperature probe. **They will proceed to Step 48 while you complete the following steps.**
- ✓ Place the **CAPPED** NaOH vial on the balance and record its mass [**Data Sheet Q42**].
- ✓ Determine the mass of the added NaOH [**Online Report Sheet Q43**].

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- ✓ **NaOH is very caustic – avoid contact!!** If suspicious of contact, wash affected area immediately with plenty of water.
48. Determine the mass of the solution (the 1.2 M HCl + the solid NaOH) [**Online Report Sheet Q44**].
 49. Continuously swirl the contents of the calorimeter (by swirling the beaker).
 50. Record the maximum temperature (to the nearest 0.01 °C) reached during this neutralization reaction [**Data Sheet Q45**].
 51. Determine change in temperature [**Online Report Sheet Q46**].
 52. Determine the heat absorbed (in kJ) by the 1.2 M HCl / solid NaOH solution [**Online Report Sheet Q47**]. Assume the solution in the calorimeter has the same specific heat as pure water.
 53. Calculate the heat (in kJ) that would be predicted to be released if the mass of NaOH in **Online Report Sheet Q43** were hydrated. Use your experimental heat of hydration determined in **Online Report Sheet Q12** to perform this calculation [**Online Report Sheet Q48**].
 54. Calculate the heat (in kJ) that would be predicted to be released if the mass of NaOH in **Online Report Sheet Q43** were neutralized. Use your experimental heat of neutralization determined in **Online Report Sheet Q36** to perform this calculation [**Online Report Sheet Q49**].
 55. Determine the % Error of your experimental data [**Online Report Sheet Q50**].

$$\% \text{ Error} = \left| \frac{Q47 - (Q48 + Q49)}{|Q47|} \right| \times 100$$

Waste Disposal: All solutions can be flushed down the sink with plenty of water.

Lab Report: Once you have turned in your Instructor Data Sheet, lab attendance will be entered and lab attendees will be permitted to access the online data / calculation submission part of the lab report (click on **Lab 7 – Calorimetry Measurements**). Enter your data accurately to avoid penalty. The lab program will take you in order to each calculation. If there is an error, you will be given additional submissions (the number and penalty to be determined by your instructor) to correct your calculation.

Calorimetry Measurements of Fusion, Hydration and Neutralization

Post-Lab Questions: The questions for this lab can be found at <http://www.Chem21Labs.com>. **Do Not Wait Until The Last Minute!!!! Computer Problems and Internet Unavailability Happen, But Deadlines Will Not Be Extended!!** On the Internet, complete any Post Lab Assignments. The computer program will check your answer to see if it is correct. If there is an error, you will be given additional submissions (the number and penalty to be determined by your instructor) to correct your answer.

Late Submission: Late submission of the lab data / calculations is permitted with the following penalties: **- 10 points** for submissions up to 1 day late, **- 20 points** for submissions up to 2 days late.

Laboratory 7

Student Data Sheet

Heat of Hydration of NaOH

- | | | |
|---|-------|----|
| 1. Mass of dry calorimeter + cap | _____ | g |
| 2. Mass of calorimeter + cap + water | _____ | g |
| 4. Temperature of water and calorimeter | _____ | °C |
| 5. Mass of NaOH vial | _____ | g |
| 6. Mass of NaOH vial after adding ~ 2 g | _____ | g |
| 8. Maximum temperature reached | _____ | °C |

Heat of Fusion of Water

- | | | |
|---|-------|----|
| 13. Mass of calorimeter + cap | _____ | g |
| 14. Mass of calorimeter + cap + water | _____ | g |
| 16. Temperature of the ice | _____ | °C |
| 17. Temperature of the warm water | _____ | °C |
| 18. Temperature of the system after the ice melts | _____ | °C |
| 19. Mass calorimeter + cap + water + ice | _____ | g |

Heat of Neutralization

- | | | |
|--|-------|----|
| 27. Mass of calorimeter + cap | _____ | g |
| 28. Mass of calorimeter + cap + HCl | _____ | g |
| 30. Temperature of the HCl solution | _____ | °C |
| 32. Maximum temperature reached in the calorimeter | _____ | °C |

Hess' Law: Heat of Hydration + Heat of Neutralization

- | | | |
|--|-------|----|
| 37. Mass of calorimeter + cap | _____ | g |
| 38. Mass of calorimeter + cap + HCl | _____ | g |
| 40. Temperature of calorimeter and contents | _____ | °C |
| 41. Mass of NaOH vial | _____ | g |
| 42. Mass of NaOH vial after adding ~ 2 g | _____ | g |
| 45. Maximum temperature reached in the calorimeter | _____ | °C |

Name: _____

Laboratory 7

Instructor Data Sheet

Heat of Hydration of NaOH

1. Mass of dry calorimeter + cap _____ g
2. Mass of calorimeter + cap + water _____ g
4. Temperature of water and calorimeter _____ °C
5. Mass of NaOH vial _____ g
6. Mass of NaOH vial after adding ~ 2 g _____ g
8. Maximum temperature reached _____ °C

Heat of Fusion of Water

13. Mass of calorimeter + cap _____ g
14. Mass of calorimeter + cap + water _____ g
16. Temperature of the ice _____ °C
17. Temperature of the warm water _____ °C
18. Temperature of the system after the ice melts _____ °C
19. Mass calorimeter + cap + water + ice _____ g

Heat of Neutralization

27. Mass of calorimeter + cap _____ g
28. Mass of calorimeter + cap + HCl _____ g
30. Temperature of the HCl solution _____ °C
32. Maximum temperature reached in the calorimeter _____ °C

Hess' Law: Heat of Hydration + Heat of Neutralization

37. Mass of calorimeter + cap _____ g
38. Mass of calorimeter + cap + HCl _____ g
40. Temperature of calorimeter and contents _____ °C
41. Mass of NaOH vial _____ g
42. Mass of NaOH vial after adding ~ 2 g _____ g
45. Maximum temperature reached in the calorimeter _____ °C