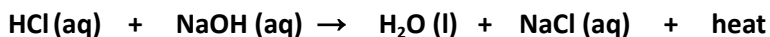


# Enthalpy of Neutralization

## Introduction

In the course of most physical processes and chemical reactions there is a change in energy. In chemistry what is normally measured is  $\Delta H$  (**enthalpy** change), the change in heat at constant pressure and ignoring any work done by the reacting system. If the reaction is exothermic, heat is given off and  $\Delta H$  has a negative value. When heat is absorbed from the surroundings during the reaction, the reaction is endothermic and  $\Delta H$  has a positive value. The change in enthalpy is a measure of the difference in energy between the bonds in the products and the bonds in the reactants. The absolute energy of a compound cannot be measured directly, but the change in enthalpy that occurs during a chemical reaction can be measured.

The reaction in today's experiment is acid / base neutralization which is an exothermic process.



The heat released by the reaction will be absorbed by the surroundings (aqueous solution). Coffee Cup Calorimetry will be employed to determine the amount of heat lost by the reaction and gained by the salt water solution. A calorimeter is simply a container used to measure the heat change. Coffee Cup Calorimetry just means that we will be measuring heat at constant pressure,  $\Delta H$ . The heat lost by the reaction will actually be transferred to both the salt water and its surroundings (the calorimeter.) The heat capacity of the calorimeter (in this case, two Styrofoam cups) usually would be calculated first; however, we have found that the heat capacity of the cups is so small that *it can be neglected*.

As the First Law of Thermodynamics applies,

Heat (q) lost by the reaction + heat (q) gained by the solution = 0

$$\text{or } q_{\text{rxn}} = -q_{\text{soln}}$$

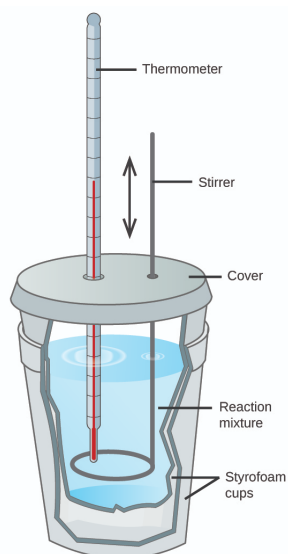
To find the heat lost by the reaction we use:

$$q = m \times C_s \times \Delta T$$

where: m = mass of the solution (The density of the solution is 1.04 g/mL)

$C_s$  = specific heat capacity of the solution (3.87 J/g°C)

$\Delta T$  = change in temperature of the solution ( $T_{\text{final}} - T_{\text{initial}}$ )



### Equipment

2 Styrofoam cups

Vernier lab quest and thermometer.

50 or 100 mL graduated cylinder

150 mL beaker

### Chemicals

2.00 M HCl

2.00 M NaOH

Image: OpenStax *Chemistry 2e*. We won't be using a cover or a stirrer.

### Procedure

1. Rinse and dry two Styrofoam cups. Place one cup inside the other. This is the calorimeter.
2. Measure out 25.0 mL of 2.00 M NaOH in a graduated cylinder and pour it into the calorimeter. Measure 25.0 mL of 2.00 M HCl in a clean graduated cylinder.
3. Determine the temperature of both the acid and the base to the nearest 0.1°C. Average the temperatures. Record this average as the initial temperature  $T_{\text{initial}}$ .
4. Pour the acid into the base quickly and carefully and gently stir with the thermometer probe. Start monitoring the temperature as soon as the two are mixed. Continue to stir and monitor the temperature. Record the maximum temperature that the solution reaches. This is  $T_{\text{final}}$ .
5. Repeat Steps 1 - 4 (above) for a second trial.

### Disposal

All reactants and products may be disposed of into the sink.

**We reuse the cups!** Rinse, dry and return to the front.

Name\_\_\_\_\_

## CHM112 Lab – Heat of Neutralization – Grading Rubric

| Criteria                                                                                                                              | Points possible | Points earned |
|---------------------------------------------------------------------------------------------------------------------------------------|-----------------|---------------|
| <b>Lab Performance</b>                                                                                                                |                 |               |
| Lab work performed correctly. Proper safety procedures followed and waste disposed of correctly. Work space and glassware cleaned up. | 3               |               |
| <b>Lab Report</b>                                                                                                                     |                 |               |
| Data table filled out correctly with units                                                                                            | 2               |               |
| Q1 (work shown in detail with units)                                                                                                  | 4               |               |
| Q2 - Q8 (1 point each)                                                                                                                | 7               |               |
| Q9 (work shown in detail with units)                                                                                                  | 2               |               |
| Q10 (work shown in detail with units)                                                                                                 | 2               |               |
| <b>Total</b>                                                                                                                          | 20              |               |

*Subject to additional penalties at the discretion of the instructor.*

|                                                    | Trial 1 | Trial 2 |
|----------------------------------------------------|---------|---------|
| Volume of 2.00 M HCl                               |         |         |
| Volume of 2.00 M NaOH                              |         |         |
| Volume of total solution                           |         |         |
| $T_{\text{initial}}$                               |         |         |
| $T_{\text{final}}$                                 |         |         |
| $\Delta T (T_{\text{final}} - T_{\text{initial}})$ |         |         |

**Remember that:**

$$q = m \times C_s \times \Delta T$$

where:  $m$  = mass of the solution (The density of the solution is 1.04 g/mL)

$C_s$  = specific heat capacity of the solution (3.87 J/g°C)

$\Delta T$  = change in temperature of the solution ( $T_{\text{final}} - T_{\text{initial}}$ )

1. Calculate the heat absorbed by the solution in each trial. ( $q_{\text{soln}}$ ) (Show all your work completely and clearly. Include proper units)

$q_{\text{soln}}$  (Trial 1) \_\_\_\_\_  $q_{\text{soln}}$  (Trial 2) \_\_\_\_\_ Average  $q_{\text{soln}}$  \_\_\_\_\_

2. Calculate  $q_{\text{rxn}}$  = \_\_\_\_\_

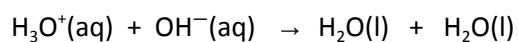
3. Calculate the number of moles of water formed (This is the same as moles of HCl):

moles of water = moles of HCl = \_\_\_\_\_

4. Calculate  $\Delta H_{\text{neutralization}}$  ( $\Delta H^\circ$ ) per mole of water (kJ/mole) formed. Make sure the sign is correct.

$\Delta H_{\text{neutralization}}$  = \_\_\_\_\_

5. Write the chemical reaction for the autoionization of water: (with equilibrium constant  $K_w$ ).
6. Find a [table](#) of  $K_w$  at different temperatures. Write the  $K_w$  value that is closest to your initial experimental temperature.
7. a) The net ionic reaction for the neutralization reaction is below. What is the relationship between the reaction below and the autoionization of water (Q5)?



- b) What is the equation for the equilibrium constant (K) for the reaction in today's lab?  $K = \frac{[\text{?}][\text{?}]}{[\text{?}][\text{?}]}$

8. Using the numerical value of  $K_w$  (Q6), calculate the value of K for our (neutralization) reaction.

9. Given the K calculated for the neutralization reaction, calculate  $\Delta G^\circ$  for the neutralization reaction.  
 $\Delta G^\circ = -RT \ln(K)$  (Use average initial T from the two trials for the reaction temperature.)  $R = 8.314 \frac{\text{Joule}}{\text{K mole}}$

10. From  $\Delta G^\circ$  above, and the  $\Delta H^\circ$  that you have calculated (Q4), calculate the  $\Delta S^\circ$  for the reaction in J/K. (Use the average initial T from the two trials for the reaction temperature.)