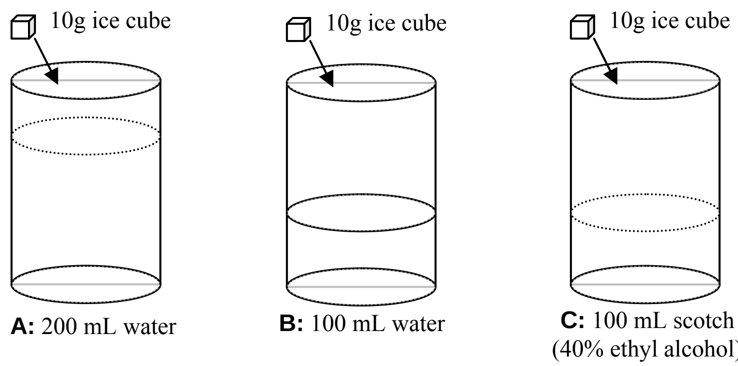
CHEM 1111 Tutorial **Week 7 - Fall 2007**

Thermochemistry - I

INTRODUCTION

Scenario: Your dad is hanging out with some of his friends at your house. He gets up to make some drinks for them. You watch as he fills three glasses. He fills the first with 200 mL of water, the second with 100 mL of water, and the third with 100 mL of scotch (which is at least 40% ethyl alcohol by volume). He adds one ice cube to each drink.



Normally, you wouldn't pay any attention to this, but you're working on your chemistry homework and realize that this is a real-life example of a calorimeter experiment. It gets you thinking . . . how much heat is involved in melting those ice cubes? What data would you need to collect to figure that out? Which drink is going to be the coolest once the ice cube melts? Warmest?

Your Challenge: To apply your understanding of several thermochemistry concepts to an everyday example of calorimetry. Thinking through this relatively simple scenario involving a *physical change* should provide a template for your thinking about systems that involve *chemical changes*.

Part I: Identifying the system and surroundings

The scenario is described and illustrated above. Identify the system and surroundings. Justify your choice.

The System	***The Surroundings***

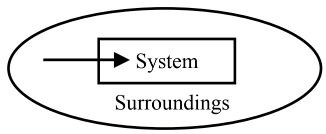
<u>Part II: How is heat involved in this process?   How *much* heat is involved?</u>

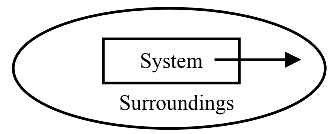
✹✹ To answer these questions, we monitor the temperature change ***of the surroundings*** ✹✹

Qualitative approach: use your hand to monitor the temperature change of the surroundings.

a) If you were holding the glass while the ice cube melted, what would you feel? *(By the way, is your hand now involved in the process . . . . ?)*

b) What does this tell you about the direction of heat flow (circle one)?



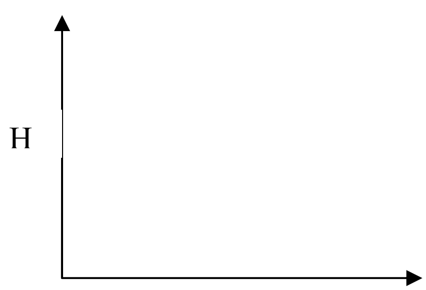


c) Thus, ***the system*** is (circle one): *releasing heat* *absorbing heat*

This is considered to be (circle one): *exothermic* *endothermic*

d) Represent this process by writing a chemical equation. Include "heat" as a reactant or product.

e) Represent this process another way. Draw an enthalpy diagram that shows the ***initial state*** and ***final state*** of the system, and ***ΔH*** for the process. Clearly label your diagram.



Quantitative approach: Use the equation $q=mc\Delta T$ to determine the quantity of heat involved.

First, consider the surroundings. Assume each drink starts at room temperature. *Make some predictions about the relative values* for q and ΔT for each drink (A, B, and C). (You won't actually calculate any values.)

a) Compare $\mathbf{q_{surr}}$ for glasses A, B, and C. Your answer should address both the magnitude AND sign of q. What information do you need to answer this question? Describe your reasoning.

b) Compare $\mathbf{\Delta T}$ for glasses A, B, and C. *(Which should be warmest after the ice melts? Coolest? Or will they be the same . . .?)* Your answer should address both the magnitude AND sign of ΔT. Describe your reasoning.

Then, consider the system. Assuming you had the temperature data for the surroundings, you could calculate q_{surr}. How does this relate to q for the system (which is, ultimately, what you're interested in)?

APPLICATION

Use the ideas from parts I and II to answer the following question.

As you begin to set up the problem, make sure you identify the system and the surroundings. This will help you with the signs on your answer.

1. A coffee-cup calorimeter contains 150.0 g of water at 25.1ºC. A 121-g block of metal is heated to 100.0ºC by placing it in a beaker of boiling water. The hot metal is added to the calorimeter, and after a time the contents of the cup reach a constant temperature of 30.1ºC. The heat capacity of the calorimeter is 23.4 J/ºC. Identify the metal used in the experiment. As part of your answer draw a picture that identifies the system, the surroundings, and the direction of heat flow when the metal is placed into the calorimeter.

Note: *You will need to use the equation $q=cm\Delta T$ twice to answer this question.*

What are you finding the first time you use the equation?

Do you expect your answer to have a positive or negative value? Why?

What are you finding the second time you use the equation?

Do you expect your answer to have a positive or negative value? Why?

This activity asked you to think about a relatively simple system based on a physical change. It was probably easy for you to identify the system and surroundings, partly because you can actually see the ice cube floating in the drink.

Try to apply some of the thinking involved in this activity to the systems you encounter in lab this week.