hs=32.3 kJ/mol=0.91 mol

I. Heat of Sublimation and Ice Cream

Writeup: I. Title, II. Prelab Q , III. Purpose, IV. Data

II. Prelab: Watch & answer the following from the video: https://www.youtube.com/watch?v=oc0ypeDELb0

1. 1. What is one difference between CO and CO_2 ,chemically?

2. What is the heat of sublimation?

3. What are intermolecular forces between molecules?

4. Looking at the phase change diagrams for water and carbon dioxide, what is the biggest, noticable difference?

5. Melting and sublimation require the ___________ of heat.

6. The molar heat of fusion is a __________thermochemical process. (endothermic or exothermic).

7. The molar heat of sublimation is a _________ thermochemical process.

8. Calculate the amount of heat required to completely sublime 40.0 g of solid dry ice (CO2) at its sublimation temperature. The heat of sublimation for carbon dioxide is 25.23 kJ/mol.

9. From the background information, draw a diagram that represents the processes that are occurring when ice absorbs heat.

10. How are internal energy (Δ E) and enthalpy (H) different?

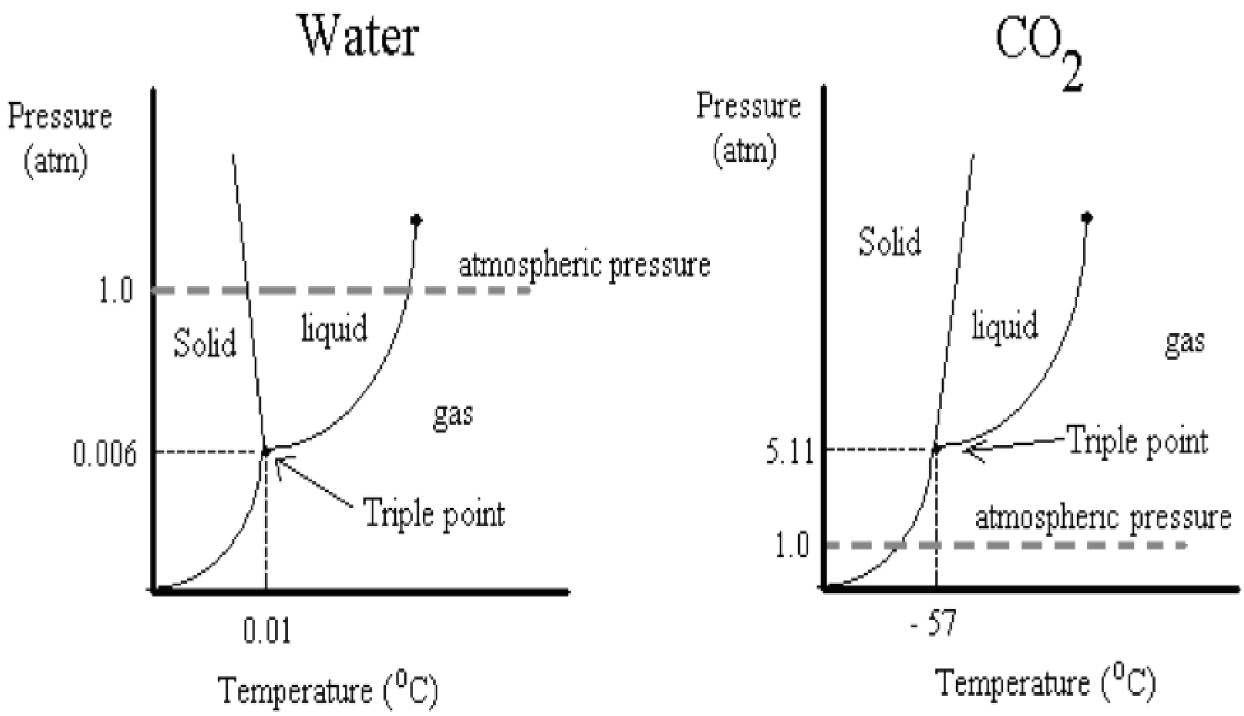
<u>Background Information:</u>

The oxides of carbon could not be more chemically different. Carbon monoxide (CO) is a polar basic gas that binds strongly to metals while carbon dioxide (CO_2) is a non-polar gas that reacts with water. Carbon monoxide is produced when hydrocarbons are burned in a limited amount of oxygen while carbon dioxide is produced when hydrocarbons are burned in excess oxygen. Carbon monoxide is a deadly poison while carbon dioxide is essential to the metabolism of plants and animals.

Carbon dioxide is the fourth most common gas in our atmosphere and its concentration is increasing. Atmospheric carbon dioxide plays an important role in maintaining surface temperatures. It is one of the “greenhouse” gases responsible for global warming. The biological processes of metabolism, which produces carbon dioxide, and photosynthesis, which consumes carbon dioxide have functioned to maintain the levels of carbon dioxide in our atmosphere. The industrial revolution has shifted this balance. The concentration of carbon dioxide has increased more than 30% to the present day value of 400 ppm. Global temperatures are rising.

<u>Purpose</u>: This experiment focuses on a physical property of CO_2 – its energy content, and a chemical property of CO_2 – its ability to dissolve in water to make the solution acidic. Like all molecules CO_2 stores potential energy in the molecular bonds that hold the oxygen to the carbon and in the intermolecular forces between molecules. **We will measure the heat of sublimation of dry ice to investigate the intermolecular forces between CO_2 molecules.**

The phase diagrams below indicate the stable forms of compounds at different pressures P and temperatures T. The lines describe when two phases are in equilibrium and the triple point is the place where all three phases can coexist in equilibrium. For example, the triple point of water occurs at 0.01 ^oC and 0.006 atm. In order to convert ice directly into steam the partial pressures of water would have to be lower than 0.006 atm. Because we live at higher pressures (i.e., at 1 atm), we always see ice melt before it vaporizes to steam. The process of converting a solid to its liquid is called melting (aka, fusion). The triple point of carbon dioxide on the other hand occurs at a much higher pressure and lower temperature, -57 ^oC and 5.11 atm. We call solid carbon dioxide “dry ice” because at normal atmospheric pressures it converts directly from the solid into the gas. The process of converting a solid into a gas is called sublimation. Both melting and sublimation require the input of heat.



Phase diagrams

Calorimetry

When a solid substance is heated, it absorbs thermal energy and its temperature increases. When the melting point of the substance is reached, addition of further thermal energy breaks up the forces holding the solid together, and a liquid begins to form. A balance is established between solid and liquid, and as more thermal energy is added the temperature remains constant while the amount of solid decreases and the amount of liquid increases. The heat absorbed when one mole of solid is melted at constant pressure is called the molar enthalpy of fusion (ΔH_{fus}). When heat is absorbed, the sign of ΔH_{fus} is, by convention, positive. Reactions that absorb heat are said to be endothermic reactions. Those which produce heat are said to be exothermic.

On the molecular scale, many processes occur as the ice absorbs heat: the ordered array of molecules in the crystal lattice of the solid is broken down into a collection of mobile, disordered liquid-phase molecules. The water molecules in ice are also losing potential energy associated with intermolecular forces between molecules and gaining kinetic energy of motion. Meanwhile, the temperature of the water, associated with the average kinetic energy of motion, rises as the molecules move about and vibrate more rapidly.

The change in internal energy for a reaction ΔE_{rxn} can be measured by running the reaction in a constant volume bomb calorimeter. By designing the calorimeter so that no heat leaks out to the surroundings, the heat absorbed by the reaction should equal the heat lost by the calorimeter. The heat absorbed by the calorimeter can be measured by the heat capacity of the calorimeter ($C_{calorimeter}$) times the change in temperature.

$$\Delta E_{rxn} = -C_{calorimeter}\,\Delta T$$

The melting of ice or the sublimation of carbon dioxide is not a constant volume process but rather a constant pressure process. To investigate the heat flows in a constant pressure process we will have to define a new form of internal energy called enthalpy, H. At constant pressure the change in enthalpy is related to the change in internal energy minus the work done on the system due to volume changes (ΔV).

$$\Delta H = \Delta E + \Delta(PV) = \Delta E + P\Delta V$$

At constant pressure, enthalpy changes are a direct measure of the heat absorbed by the system. Enthalpy changes are easier to measure than internal energy changes because it is easier to maintain the constant pressure than constant volume.

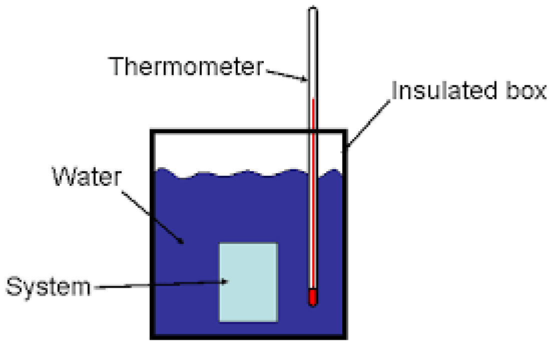


Diagram of a constant pressure calorimeter

Δ H is measured with a calorimeter where the amount of heat flowing is reflected in a temperature change of a known mass of water. In fact, the unit of energy, the calorie, has been defined to be that amount of heat that will raise one gram of pure water one degree Celsius in temperature. Although the calorie is convenient energy unit, the scientific community now uses the Joule as its standard unit of energy. (1 cal = 4.184 J) Assuming no heat loss during the transfer, the heat of a reaction will be equivalent to the heat absorbed by g grams of water when the temperature raises Δ T degrees Celsius, or

$\Delta H_{rxn} = - g\ H_2O\ (4.184\ J/g\ H_2O\ ^oC\ )\ \Delta T$

This is the basic equation describing a solution calorimeter that is intended to measure the change in enthalpy during a constant pressure process.

Procedure: Part I. Enthalpy of Sublimation

1. Weigh the calorimeter.

2. Fill a 100 mL volumetric flask to the mark with distilled water adjusted to room temperature and weigh it. Record the stable temperature; this is the starting temperature of the water.

3. Once you have weighed the calorimeter and measured the stable temperature reading, take a piece of CO_2 (between 2 and 10 grams) with tweezers or crucible tongs and put it into the calorimeter. The mass will keep rolling down as it sublimes, therefore after a few seconds record the mass, not bothering with the last digit. Take the calorimeter off the balance and put the lid on it with the thermometer through one of the slits. Pour the 100 mL of water into the cup, replace the cover, and gently push down, trying not to spill too much water as it seeps into the lower cup. Reweigh volumetric flask to determine the mass of water transferred.

4. Even though you see your system “smoking”, the “smoke” shouldn’t be cold to the touch as the CO_2 gas has to absorb the heat of the water as it travels through it. Swirl the calorimeter gently, again making sure to keep the cover on. Watch the temperature drop. When it stabilizes and you don’t feel the bubbles in solution, swirl again and look in the calorimeter. If it’s cloudy, return to swirling. If it’s clear, record the stabilized temperature.

5. Empty your calorimeter into the sink with lots of water running.

Data Table/Calculations Table: $\Delta H_{sub} = -$ (g of water)(4.184J J/g ° C)(T_2-T_1)

g of water	
T_1	
T_2	
ΔH_{sub}	
g CO_2	
ΔH_{sub} per g CO_2	

Calculation: The heat gained by the CO_2 (ΔH_{sub}) is equal to the heat released by the water. Knowing that the specific heat of water is 4.184 J g^{-1} C^o, calculate the average ΔH_{sub}/g CO_2 and the ΔH_{sub}/mol CO_2. Compare versus the literature value.

Ingredients (from video):

2 cups half and half or heavy cream

1 teaspoon vanilla extract + desired flavoring (strawberry, lemon, coconut, etc.)

1 cup powdered sugar

2-3 cups dry ice

Part II. Instructions: Get in groups of 4

1. Mix 2 cups half and half and 1 cup powdered sugar in a large mixing bowl; add teaspoon vanilla extract. Beat until smooth with large spoon.
2. Add desired flavoring (strawberries, chocolate chips, different extracts, etc)
3. Place chunk of dry ice in plastic bag. Crush the dry ice with another bowl or spoon so that small chunks remain. Leave the bag partially open to allow the gas to escape. Slowly add dry ice to mixture from step 1, 1 cup at a time, stirring constantly. The ice cream base will start to thicken and chill. Continue mixing and adding dry ice until desired texture has been met.
4. ENSURE ALL CHUNKS OF DRY ICE HAVE COMPLETELY DISSIPATED BEFORE INGESTING ANY OF THE ICE CREAM. DRY ICE IS VERY DANGEROUS AND MAY CAUSE SEVERE DAMAGE TO YOUR MOUTH/THROAT IF SWALLOWED.
5. Once all of the dry ice has dissipated transfer to small bowl, then serve and enjoy:)
6. Clean up by washing and drying all equipment and bowls thoroughly. Return all items where they belong.