

Thermochemistry HOMEWORK

Lesson 1 - Homework

1. Imagine a bowl of soup in the kitchen. The soup is the system. Describe the surroundings.

The surroundings include the bowl, the air, and the entire room around it.

2. Complete the sentence with the words endothermic/exothermic; and positive/negative. When heat is absorbed by the system from the surroundings, that process is said to be endothermic and the sign of q is positive.

3. Which is not true for an endothermic reaction?

- The temperature of the surroundings decreases.
- The heat for the reaction is positive.
- Heat flows from the surroundings into the system.

d. The products have a lower heat energy than the reactants.

e. All of the above are true.

4. If the same amount of heat is added to 5.00 g samples of each of the metals below, which metal will experience the greatest temperature change?

Metal	Specific Heat Capacity (J/g $\cdot$ °C)
Al	0.907
<u>Sn</u>	<u>0.205</u>
Cu	0.385
Fe	0.449
Pb	0.129

The metal with the LOWEST specific heat capacity will require the LEAST amount of energy to increase its temperature.

If the same amount of heat is added to each metal, the one with the LOWEST specific heat capacity will reach the highest temp.

5. Sometimes, specific heat capacities are given in molar instead of grams, with the units J/mol $\cdot$ °C. If the molar heat capacity of a compound with the formula C_2H_6O is 88 J/mol $\cdot$ °C, what is the specific heat (in J/g $\cdot$ °C) of this substance?

$$88 \text{ J/mol}\cdot^\circ\text{C} \times (1 \text{ mol} / 46.1 \text{ g}) = 1.9109978 \text{ J/g}\cdot^\circ\text{C} \rightarrow 1.91 \text{ J/g}\cdot^\circ\text{C}$$

6. A sample of aluminum (0.90 J/g $\cdot$ °C) absorbs 4.32 J of energy while being heated from 23.2 °C to 38.5 °C. Determine the mass of the sample.

$$\begin{aligned} q &= m \cdot C_p \cdot \Delta T \\ 4.32 \text{ J} &= (x \text{ g}) (0.90 \text{ J/g}\cdot^\circ\text{C}) (38.5 - 23.2) \\ (x \text{ g}) (0.771 \text{ J/g}\cdot^\circ\text{C}) &= 4.32 \text{ J} \end{aligned}$$

This is necessary to convert the temperatures to Delta T. This is because the CHANGE in Celsius is the same as the CHANGE in Kelvin. Since you're given both temperatures, you can find the CHANGE in temperature. If you chose using just temperature, you would need to convert the single temperature into Delta T.

7. An iron sphere with a volume of 3.3cm³ and density of 7.86g/cm³ at 25.0 °C is dropped into a perfectly insulated beaker containing 100 g of H₂O at 100.0 °C. The specific heats of Fe and H₂O are 0.450 and 4.184 J/g $\cdot$ °C, respectively. What is the final temperature (in °C) of the iron sphere?

$$\begin{aligned} \text{The volume and density is first given:} \\ 3.3 \text{ cm}^3 \times (7.86 \text{ g/cm}^3) &= 26.038 \text{ g} \end{aligned}$$

$$\begin{aligned} \text{Heat gained} &= \text{Heat lost} \\ -q_{\text{Fe}} &= q_{\text{H}_2\text{O}} \\ -m_{\text{Fe}}(C_p) &= -m_{\text{H}_2\text{O}}(C_p) \\ \text{Fe} &= \text{Water} \end{aligned}$$

$$\begin{aligned} -m_{\text{Fe}}(C_p) &= -m_{\text{H}_2\text{O}}(C_p) \\ -(26.038 \text{ g})(0.450 \text{ J/g}\cdot^\circ\text{C}) &= -(100 \text{ g})(4.184 \text{ J/g}\cdot^\circ\text{C}) - (100 \text{ g})(4.184 \text{ J/g}\cdot^\circ\text{C}) \\ 0.117717 &= 1.682 \text{ J/g} = 1.78 \text{ J/g} \\ 104.09117 &= 1.6298 \text{ J/g} \\ T_f &= 1.71 \text{ J/g}\cdot^\circ\text{C} \rightarrow 1.71 \text{ J/g}\cdot^\circ\text{C} - 273 = -271.2888 \text{ }^\circ\text{C} \end{aligned}$$

The lower the number with J/g°C the heat increase it is meaning it is higher temperature. $T_f = 1.71 \text{ J/g}\cdot^\circ\text{C} \rightarrow 1.71 \text{ J/g}\cdot^\circ\text{C} - 273 = -271.2888 \text{ }^\circ\text{C}$

8. Given that silver has a specific heat of 0.235 J/g $\cdot$ °C.

a. If a 75.0g sample of silver at 100. °C is dropped into a calorimeter (perfectly insulated container) containing 100. g of water at 22.0 °C, what will be the final temperature of the system at equilibrium? (The specific heat capacity of water is 4.184 J/g $\cdot$ °C).

$$\begin{aligned} \text{Heat gained} &= \text{Heat lost} \\ -q_{\text{Ag}} &= q_{\text{H}_2\text{O}} \\ -m_{\text{Ag}}(C_p) &= -m_{\text{H}_2\text{O}}(C_p) \\ \text{Silver} &= \text{Water} \end{aligned}$$

$$\begin{aligned} -m_{\text{Ag}}(C_p) &= -m_{\text{H}_2\text{O}}(C_p) \\ -(75.0 \text{ g})(0.235 \text{ J/g}\cdot^\circ\text{C}) &= -(100 \text{ g})(4.184 \text{ J/g}\cdot^\circ\text{C}) - (100 \text{ g})(4.184 \text{ J/g}\cdot^\circ\text{C}) \\ -17.625 &= 1.7828 \text{ J/g} = 1.78 \text{ J/g} \\ 406.0917 &= 1.6298 \text{ J/g} \\ T_f &= 246.10917 \text{ }^\circ\text{C} \rightarrow 246.11 \text{ }^\circ\text{C} - 273 = -26.8908 \text{ }^\circ\text{C} \end{aligned}$$

b. How much heat energy is transferred to the water in part a?

$$\begin{aligned} q &= m \cdot C_p \cdot \Delta T \\ q &= (100 \text{ g})(4.184 \text{ J/g}\cdot^\circ\text{C})(22.0 - 273) \\ q &= 111.1 \text{ J/g} \end{aligned}$$

9. A 75.0g sample of water at 80.0 °C is added to an unknown amount of water at 22.0 °C in a perfectly insulated container. The final temperature is 30.0 °C. How many grams of water were there initially at 22.0 °C? (The specific heat capacity of water is 4.184 J/g $\cdot$ °C).

$$\begin{aligned} \text{Heat gained} &= \text{Heat lost} \\ -q_{\text{H}_2\text{O}} &= q_{\text{H}_2\text{O}} \\ \text{Cold Water} &= \text{Hot Water} \end{aligned}$$

$$\begin{aligned} -(75.0 \text{ g})(4.184 \text{ J/g}\cdot^\circ\text{C}) &= -(x \text{ g})(4.184 \text{ J/g}\cdot^\circ\text{C}) - (x \text{ g})(4.184 \text{ J/g}\cdot^\circ\text{C}) \\ -313.8 &= 1.7828 \text{ J/g} \\ x &= 176.11117 \text{ g} \rightarrow 176 \text{ g} \end{aligned}$$

10. The specific heat capacity of copper is 0.385 J/g $\cdot$ °C. What quantity of heat is required to heat 200 g of copper from -6.0 °C to +22.0 °C? (The specific heat capacity of water is 4.184 J/g $\cdot$ °C).

$$\begin{aligned} q &= m \cdot C_p \cdot \Delta T \\ q &= (200 \text{ g})(0.385 \text{ J/g}\cdot^\circ\text{C})(-6.0 - 22.0) \\ q &= 111.1 \text{ J/g} \end{aligned}$$

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