HEAT TRANSFER

Name ____ Section_______

Problem Statement: How is heat transferred between substances?

I. Data Collection

A. Go to https://media.pearsoncmg.com/bc/bc_0media_chem/chem_sim/calorimetry/Calor.php. The simulation will open to an image of the calorimeter setup, which is quickly replaced with a new screen with an Overview page. You are welcome to read the Overview Page, and by clicking on the Learning Outcomes tab near the top of the display, you may read the Learning Outcomes Page. After reviewing these two pages click on the Experiment tab. When the screen changes the page will show two buttons: Run Demonstration button and Run Experiment button. You are welcome to click on the Run Demonstration button, but the instructions below are for the Run Experiment button. After clicking on the Run Experimental button the screen will look like Figure I.

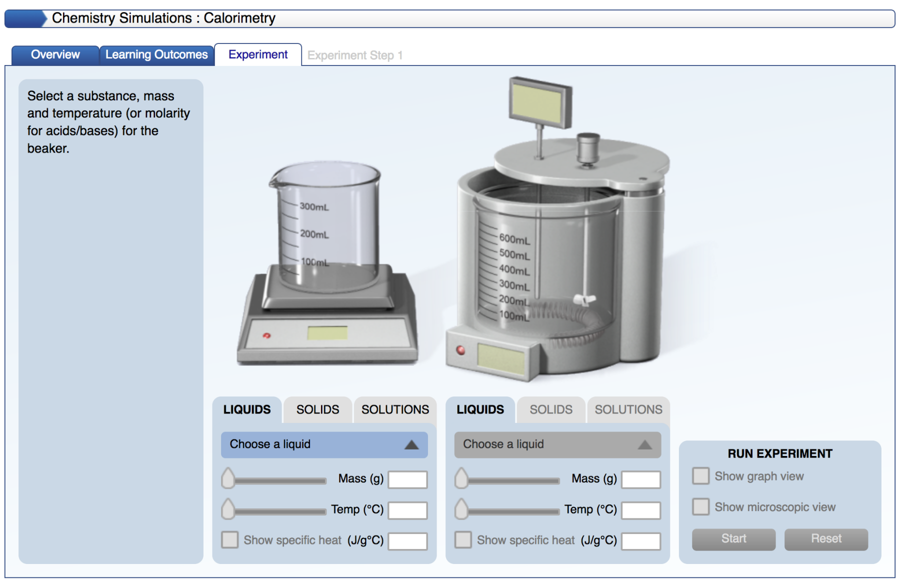


Figure I. Experimental Setup for the Calorimetry Simulation

The Experimental setup shows a beaker on a hot plate to the left, and a calorimeter on the right. Below the beaker and hot plate are three tabs (Liquids, Solids and Solutions). In this activity you will be using the Solids tab beneath the beaker and hotplate, and the Liquids tab beneath the calorimeter.

B. Beneath the beaker and hot plate click on the Solids tab and select Ag. Adjust mass to 20.0 g and adjust the temperature to 200. °C. Click the Next button in the left frame near the bottom of the screen. Now click on the Liquids tab beneath the calorimeter and add 50.00 g of water and adjust the water temperature to 20.00 °C. Record the beginning conditions in the Table I below.

Table I.

	Ag	water
Mass	**20.0 g**	**50.0 g**
Initial Temp	**200. ˚C**	**20.0 ˚C**
Final Temp	**23.96 ˚C**	**23.96 ˚C**
Change in Temp	**176.04 ˚C**	**3.96 ˚C**
Specific Heat ($J\ g^{-1}\ ^{\circ}C^{-1}$)	**0.235**	**4.184**

C. In the Run Experiment section click on the Start button. What do you observe happening? Record the final conditions of Ag and the water in the table above.

After clicking on the start button the lid on the calorimeter closes and the temperature of the water that contains the piece of metal begins to increase. This makes sense because the hot piece of metal is added to the cool water.

II. Data Analysis and Interpretation

A. Which substance, Ag or water, loses heat when they are combined? Which substance, Ag or water, gains heat when they are combined? Which process is endothermic and which is exothermic?

The silver is losing heat, since it is at the higher temperature initially. The water gains heat as it is at the lower temperature initially. The heat lost by the metal is exothermic, and the heat gained by the water is endothermic.

B. Calculate the heat (q) transferred to or from Ag. Use the equation $q = m\ C_S\ \Delta t$ (q is heat in Joules, m is mass, C_S is the heat content, and Δt is the change in temperature).

$$\mathbf{q = mass * C_S * \Delta T}$$

$$\mathbf{q = 20.0\ grams * 0.235 \frac{J}{g\ ^{\circ}C} * (23.96\ ^{\circ}C - 200\ ^{\circ}C) = -8.27 \times 10^2\ Joules}$$

C. Calculate the heat (q) transferred to or from water.

$$\mathbf{q = mass * C_S * \Delta T}$$

$$\mathbf{q = 50.0\ grams * 4.184 \frac{J}{g\ ^{\circ}C} * (23.96\ ^{\circ}C - 20.0\ ^{\circ}C) = 8.28 \times 10^2\ Joule}$$

D. Compare the heats associated with the Ag and water. Make a generalization concerning these heats.

While the heat lost by the metal is 1 J lower compared to the heat gained by the water the amount of heat lost and gained should be the same value.

E. How would your results have been different if you had used different amounts of Ag and water starting at different temperatures? Try this out by doing new experiments with Ag and water and report your data and your conclusions below.

	Ag	Water
Mass	**40.0 grams**	**50.0 grams**
Initial Temperature	**200 ˚C**	**20 ˚C**
Final Temperature	**27.74 ˚C**	**27.74 ˚C**
Change in Temperature	**172.26 ˚C**	**7.74 ˚C**
Heat (q = mass * SH * ΔT)	**-1.62×10^3 Joules**	**1.62×10^3 Joules**

In this experiment the heat transferred is different (more than the heat transferred in the earlier experiment), but as before the heat lost by the metal is equal and opposite to the heat gained by the water. Also notice that doubling the amount of metal almost doubled the amount of heat released by the metal.

Heat released by the hot metal:

$q_{hot} = \text{mass} * C_s * \Delta T$

$q_{hot} = 40.0 \text{ grams} * 0.235 \frac{J}{g\ ^\circ C} * (27.74\ ^\circ C - 200\ ^\circ C) = -1.62 \times 10^3 \text{ Joules}$

Heat absorbed by the cold water:

$q_{cold} = \text{mass} * C_s * \Delta T$

$q_{cold} = 50.0 \text{ grams} * 4.184 \frac{J}{g\ ^\circ C} * (27.74\ ^\circ C - 20\ ^\circ C) = 1.62 \times 10^3 \text{ Joules}$

	Ag	Water
Mass	**20.0 grams**	**100.0 grams**
Initial Temperature	**200 ˚C**	**20.00 ˚C**
Final Temperature	**22.00 ˚C**	**22.00 ˚C**
Change in Temperature	**178 ˚C**	**2.00 ˚C**
Heat (q = mass * SH * ΔT)	**-8.37×10^2 Joules**	**8.37×10^3 Joules**

In this experiment we used five times as much water, but kept the amount of Ag constant and observe that the heat lost by the metal is equal and opposite to the heat gained by the water.

Heat released by the hot metal:

$q_{hot} = \text{mass} * C_S * \Delta T$

$q_{hot} = 20.0 \text{ grams} * 0.235 \frac{J}{g\ {}^{\circ}C} * (22.00\ {}^{\circ}C - 200\ {}^{\circ}C) = -8.37 \times 10^2 \text{ Joules}$

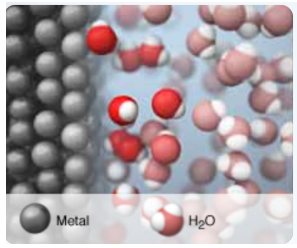
Heat absorbed by the cold water:

$q_{cold} = \text{mass} * C_S * \Delta T$

$q_{cold} = 100.0 \text{ grams} * 4.184 \frac{J}{g\ {}^{\circ}C} * (22.00\ {}^{\circ}C - 20\ {}^{\circ}C) = 8.37 \times 10^2 \text{ Joules}$

III. Particulate Level View

A. Repeat the experiment you set up in I.B. above. This time be sure to check the box to show the microscopic view. You may want to repeat the experiment with the box selected to see all of the different behaviors. In the space below draw a picture depicting a microscopic view of a piece of solid silver metal in liquid water.



B. In II. A. you indicated which substance lost heat and which substance gained heat when a hot piece of silver metal is added to water at room temperature. Using words, and if you like, pictures, describe how heat was transferred between the two substances. Be sure to include what you observed happening to the atoms of silver and molecules of water, during the period just after they were added to each other and until the final temperature was reached, that would help explain how heat is transferred.

Here the metal atoms from the hot metal transfer energy to the water molecules whenever a water molecule collides with the metal atom. The water molecule with the additional energy collides with other water molecules to transfer energy. This happens repeatedly as more collisions between water molecules and metal atoms, and collisions between water molecules continue to occur. As water molecules collide with the hot metal atoms the metal atoms lose energy and the water molecules gaain energy, until everything in the system is at the same energy.

IV. Data Collection
Repeat the experiment for Al, Cu, and Fe. Record the data you collect in the following table.

	Al	water	Cu	water	Fe	water
Mass	**20.0**	**50.0**	**20.0**	**50.0**	**20.0**	**50.0**
T_i	**200**	**20**	**200**	**20**	**200**	**20**
T_f	**34.30**	**34.30**	**26.39**	**26.39**	**27.41**	**27.41**
ΔT	**165.70**	**14.30**	**173.61**	**6.39**	**172.59**	**7.41**
Specific heat (J g^{-1} $^{\circ}C^{-1}$)	**0.90**	**4.184**	**0.385**	**4.184**	**0.452**	**4.184**
q	**-2.98 x 10^3**	**2.99 x 10^3**	**-1.34 x 10^3**	**1.34 x 10^3**	**-1.56 x 10^3**	**1.55 x 10^3**
molar heat capacity (J mol^{-1} $^{\circ}C^{-1}$)	**24.3**		**24.5**		**25.2**	

V. Data Analysis and Interpretation

A. Calculate the heat lost or gained by each metal. Show your work for one of the calculations below.

In all cases the metal lost/released heat to the water. For Al

$q_{hot\ metal}$ = mass * C_s * ΔT = 20.0 g · 0.90 J g^{-1} $^{\circ}C^{-1}$ · (34.30 $^{\circ}$C – 200. $^{\circ}$C)

$q_{hot\ metal}$ = - 2.98 x 10^3 J

B. Compare the results for all four metals. How are these metals different from each other?

The metals differ from each other by their specific heats. The higher the specific heat the more heat that is transferred to the water, and the higher the final temperature.

C. Which of these metals would make the best cookware? Explain your answer.

Of the metals tested in this experiment aluminum would be the best cookware. It will undergo the smallest temperature changed when heated. Cooking in an aluminum cookware will not acquire as high temperatures, which will be better for the metal.

D. Calculate the molar heat capacity for Al, Cu, and Fe in units of $\frac{J}{mol\ ^{\circ}C}$. Record the value in the table on the previous page. Compare the molar heat capacity, C_s for each of the metals.

the value in the table on the previous page. Compare the molar heat capacity for each of the metals.

For aluminum:

$$0.90\ \frac{J}{g\ ^\circ C}\ 26.98\ \frac{g}{mol} = 24.3\ \frac{J}{mol\ ^\circ C}$$

For copper:

$$0.385\ \frac{J}{g\ ^\circ C} * 63.55\ \frac{g}{mol} = 24.5\ \frac{J}{mol\ ^\circ C}$$

For iron:

$$0.452\ \frac{J}{g\ ^\circ C} * 55.85\ \frac{g}{mol} = 25.2\ \frac{J}{mol\ ^\circ C}$$

VI. Conclusions

A. Repeat the experiment for the unknown solid metals. Record the data you collect in the following table.

	Unknown I	water	Unknown II	water
m	**20.0 g**	**50.0 g**	**20.0 g**	**50.0 g**
T_i	**200. ˚C**	**20.0 ˚C**	**200. ˚C**	**20.0 ˚C**
T_f	**26.44 ˚C**	**26.44 ˚C**	**22.18 ˚C**	**22.18 ˚C**
Δt	**173.56 ˚C**	**6.44 ˚C**	**178.82 ˚C**	**2.18 ˚C**
Specific heat (J g^{-1} $°C^{-1}$)				
q				

B. Calculate values for the specific heats for the two unknown metals.

For Unknown I the heat gained by water :

$$q_{hot\ metal} = -\ q_{cold\ water}$$

$$(mass * C_S * \Delta T)_{hot\ metal} = -\ (50.0\ g * 4.184\ \frac{J}{g\ ^\circ C} * (35.03\ ^\circ C - 20.0\ ^\circ C))_{cold\ water}$$

so the metal must have lost that much heat and the specific heat would be

$$(mass * C_S * \Delta T)_{hot\ metal} = -1.35 \times 10^3\ Joules$$

$$C_S = \frac{q}{\text{mass } \Delta T} = \frac{-1.35 \times 10^3 \text{ Joules}}{20.0 \cdot 173.56 \text{ °C}} = 0.388 \frac{\text{J}}{\text{g °C}}$$

For Unknown II the heat gained by water :

$$q_{\text{hot metal}} = - q_{\text{cold water}}$$

$$(\text{mass} * C_S * \Delta T)_{\text{hot metal}} = - (50.0 \text{ g} * 4.184 \frac{\text{J}}{\text{g °C}} * (22.18 \text{ °C} - 20.0 \text{ °C}))_{\text{cold water}}$$

so the metal must have lost that much heat and the specific heat would be

$$(\text{mass} * C_S * \Delta T)_{\text{hot metal}} = -4.56 \times 10^2 \text{ Joules}$$

$$C_S = \frac{q}{\text{mass } \Delta T} = \frac{-4.56 \times 10^2 \text{ Joules}}{20.0 \cdot 178.82 \text{ °C}} = 0.128 \frac{\text{J}}{\text{g °C}}$$

C. Based on the comparison you made in IV.D., estimate the molar heat capacity for each unknown metal.

Assume the molar heat capacity is an average of the molar heat capacities of aluminum, copper and iron: $25.0 \frac{\text{J}}{\text{mol °C}}$

D. Calculate the molar mass of each unknown metal.

For Unknown I

$$0.388 \frac{\text{J}}{\text{g °C}} \cdot \text{MM} = 25.0 \frac{\text{J}}{\text{mol °C}}$$

$$\text{MM} = 64.4 \frac{\text{g}}{\text{mol}}$$

For Unknown II

$$0.128 \frac{\text{J}}{\text{g °C}} \cdot \text{MM} = 25.0 \frac{\text{J}}{\text{mol °C}}$$

$$\text{MM} = 195 \frac{\text{g}}{\text{mol}}$$

E. Assuming the unknown metals are pure substances, identify them. Show how you arrived at your answers below.

Unknown I could be copper or zinc, while Unknown II could be platinum or gold.