Worksheet #17

Calculating Heat

1. How much heat is needed to bring 12.0 g of water from 28.3 °C to 43.87 °C, if the specific heat capacity of water is 4.184 J/(g•°C)?

2. How much heat is released when 143 g of ice is cooled from 14 °C to – 75 °C, if the specific heat capacity of ice is 2.087 J/(g•°C).

3. When 137 mL of water at 25 °C is mixed with 82 mL of water at 76 °C, what is the final temperature of the water? Assume that no heat is lost to the surroundings and that the density of water is 1.00 g/mL.

4. An unknown volume of water at 14.3 °C was added to 35.5 mL of water at 37.8 °C. If the final temperature was 28.8 °C, what was the unknown volume? Assume that no heat is lost to the surroundings and that the density of water is 1.00 g/mL

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Determining Specific Heat Capacity

5. An alloy of unknown composition is heated to 137 °C and placed into 100.0 g of water at 25.0 °C. If the final temperature of the water was 36.4 °C, and the alloy weighed 2.71 g, what is the specific heat capacity of the alloy? The specific heat of water is 4.184 J/gºC.

6. A 45.0 g rock is heated to 97.2 °C and placed into 75.3 g of water originally at 32.0 °C. If the final temperature of the water was 46.2 °C, what is the specific heat capacity of the rock?

7. Given that the specific heat of gold is 0.129 J/gºC, calculate the final system temperature if a 200.0 g block of gold at 100.0 ºC is placed in a coffee-cup calorimeter containing 50.0 g of water at an initial temperature of 25.0 ºC.

1. How many calories are needed to raise the temperature of exactly 500 g of water from 22.15 ºC to 24.70 ºC?

2. How many joules are released when 30.0 mL of chloroform cool 18.0 ºC? The s pecific heat of chloroform is 0.971 J/(g)(ºC) and its density is 1.50 g/mL.

3. The specific heat of benzene is 1.83 J/(g)(ºC). Suppose that you have a 200 g sample of benzene at 45.0 ºC and remove 1.463 kJ. What does its temperature become?

4. Determine the specific heat of Cu from the fact that 64.0 J are needed to raise the temperature of 15.0 g of Cu metal from 22.0 ºC to 33.0 ºC.

6.33 Calculate *q* when 12.0 g of water is heated from 20.ºC to 100.ºC.

6.35 A 295-g aluminum engine part at an initial temperature of 3.00 ºC absorbs 85.0 kJ of heat. What is the final temperature of the part (*c* of Al = 0.900 J/g·K)?

6.37 Two iron bolts of equal mass—one at 100.ºC, the other at 50 ºC—are placed in an insulated container. Assuming the heat capacity of the container is negligible, what is the final temperature inside the container (*c* of iron = 0.450 J/g·K)?

6.39 When 165 mL of water at 22 ºC is mixed with 85 mL of water at 82 ºC, what is the final temperature? (Assume that no heat is lost to the surroundings; *d* of water is 1.00 g/mL.)

6.41 A 505-g piece of copper tubing is heated to 99.9 ºC and placed in an insulated vessel containing 59.8 g of water at 24.8 ºC. Assuming no loss of water and a heat capacity for the vessel of 10.0 J/K, what is the final temperature of the system (*c* of copper = 0.387 J/g·K)?

Answers to Worksheet # 17

Calculating Heat

The specific heat capacity (c) of a substance is the amount of heat required to raise the temperature of 1 gram of a substance by 1 K. Units are in J/g•K or J/g•°C. The molar heat capacity (C) of a substance is the amount of heat required to raise the temperature of 1 mole of a substance by 1 K. Units are in J/mol•K or J/mol•°C. To calculate heat, remember that q = mcΔT, where $\Delta T = T_{final} - T_{initial}$.

1. ΔT = 43.87 °C – 28.3 °C = 15.6 °C

$$q = mc\Delta T = (12.0g)\left(4.184\frac{J}{g\cdot°C}\right)(15.6°C) = 783J$$ Heat is absorbed by the system.

2. ΔT = -75 °C – 14 °C = -89 °C

$$q = mc\Delta T = (143g)\left(2.087\frac{J}{g\cdot°C}\right)(-89°C) = -27,000J = -27kJ$$ Heat is released by the system.

3. $q_{sys} + q_{surr} = 0$,

$$mass(g) = 137mLH_2O\left|\frac{1.00g}{1mL}\right| = 137gH_2O \qquad mass(g) = 82mLH_2O\left|\frac{1.00g}{1mL}\right| = 82gH_2O$$

$$(137g)\left(4.184\frac{J}{g°C}\right)(T_f - 25°C) + (82g)\left(4.184\frac{J}{g°C}\right)(T_f - 76°C) = 0$$

$$\left(573\frac{J}{°C}\right)T_f - 14330J + \left(343\frac{J}{°C}\right)T_f - 26075J = 0$$

$$\left(916\frac{J}{°C}\right)T_f = 40405J \qquad T_f = \frac{40405J}{916J} = 44.1°C$$

4. $mass(g) = 35.5mLH_2O\left|\frac{1.00g}{1mL}\right| = 35.5gH_2O$

$$(X)\left(4.184\frac{J}{g°C}\right)(28.8°C - 14.3°C) + (35.5g)\left(4.184\frac{J}{g°C}\right)(28.8 - 37.8°C) = 0$$

$$\left(60.7\frac{J}{g}\right)X - 1337J = 0 \qquad X = \frac{1337J}{60.7J} = 22.0g \quad vol(mL) = 22.0g\left|\frac{1mL}{1.00g}\right| = 22.0mL$$

Determining Specific Heat

These problems are based on the fact that $q_{sys} + q_{surr} = 0$. Note that the final temperature of the system and the surroundings has to be the same.

5. The system is the alloy, and the surroundings are water.

$\Delta T_{sys} = 36.4\ °C - 137\ °C = -101\ °C$ $\Delta T_{surr} = 36.4\ °C - 25.0\ °C = 11.4\ °C$

$$(2.71g)c_{sys}(-101°C) + (100.0g)\left(4.184\frac{J}{g°C}\right)(11.4°C) = 0$$

$$-274g°C\,c_{sys} + 4770J = 0 \qquad c_{sys} = \frac{-4770J}{-274g°C} = 17.4\frac{J}{g°C}$$

6. The system is the rock and the surroundings are water.

$\Delta T_{sys} = 46.2\ °C - 97.2\ °C = -51.0\ °C$ $\Delta T_{surr} = 46.2\ °C - 32.0\ °C = 14.2\ °C$

$$(45.0)c_{sys}(-51.0°C) + (75.3g)\left(4.184\frac{J}{g°C}\right)(14.2°C) = 0$$

$$-2295g°C\,c_{sys} + 4474J = 0 \qquad c_{sys} = \frac{-4474J}{-2295g°C} = 1.95\frac{J}{g°C}$$

7. The system is gold and the surroundings are water.

$$(200.0g)\left(0.129\frac{J}{g°C}\right)(T_f - 100.0°C) + (50.0g)\left(4.184\frac{J}{g°C}\right)(T_f - 25.0°C) = 0$$

$$\left(25.8\frac{J}{°C}\right)(T_f - 100.0°C) + \left(209\frac{J}{°C}\right)(T_f - 25.0°C) = 0$$

$$25.8\frac{J}{°C}T_f - 2580J + 209\frac{J}{°C}T_f + 5230J = 0$$

$$235\frac{J}{°C}T_f = 7810J \qquad T_f = \frac{7810J}{235\frac{J}{°C}} = 33.2°C$$

1. $$? \text{ cal} = 500 \cancel{\text{ g water}} \cdot \frac{1 \text{ cal}}{(\cancel{\text{ g water}})(\cancel{^\circ\text{C}})} \cdot 2.55 \cancel{^\circ\text{C}} = 1275 \text{ cal} = 1.28 \cdot 10^3 \text{ cal}$$

[ΔT = 24.70 ºC – 22.15 ºC = 2.55 ºC]

2. $$? \text{ J} = 30.0 \cancel{\text{ mL CHCl}_3} \cdot \frac{1.50 \cancel{\text{ g CHCl}_3}}{\cancel{\text{mL CHCl}_3}} \cdot \frac{0.971 \text{ J}}{(\cancel{\text{ g CHCl}_3})(^\circ\text{C})} \cdot (-18.0 \,^\circ\text{C}) = -786 \text{ J}$$

3. $$? \; \Delta T = ? \,^\circ\text{C} = -1463 \text{ J} \cdot \frac{(\cancel{\text{ g benzene}})(^\circ\text{C})}{1.83 \text{ J}} \cdot \frac{1}{20.0 \cancel{\text{ g benzene}}} \cdot -40.0 \,^\circ\text{C}$$

$T_{final} = T_{init} + \Delta T = 45 \;^\circ\text{C} \; - \; 40.0 \;^\circ\text{C} \; = 5.0 \;^\circ\text{C}$

4. $$\text{Specific heat} = \frac{? \text{ cal}}{(\text{gram Cu})(^\circ\text{C})} = \frac{64.0 \text{ J}}{(15.0 \text{ g Cu})(11.0 \,^\circ\text{C})} = \frac{0.388 \text{ J}}{(\text{gram Cu})(^\circ\text{C})}$$

[ΔT = 33.0 ºC – 22.0 ºC = 11.0 ºC]

5. $$? \text{ kcal} = 85.0 \cancel{\text{ g water}} \cdot \frac{\cancel{\text{mole water}}}{18.0 \cancel{\text{ g water}}} \cdot \frac{9.71 \text{ kcal}}{\cancel{\text{mole water}}} = 45.9 \text{ kcal}$$

6.33 Plan: The heat required to raise the temperature of water by 80.°C is found by using equation 6.7, or

q = c x mass x ΔT. The specific heat capacity, c_{water}, is found in Table 6.4. Because the Celsius degree is the same size as the kelvin degree, ΔT = 80.°C = 80. K.

$$q(\text{J}) = (\text{mass}) \; C \; (\Delta T) = (12.0 \text{ g})\left(4.184 \frac{\text{J}}{\text{g } ^\circ\text{C}}\right)((100. - 20.)^\circ\text{C}) = 4016.64 = 4.0 \text{ x } 10^3 \text{ J}$$

6.35 q(J) = (mass) C (ΔT)

85.0 kJ(10^3 J/kJ) = (295 g) (0.900 J/g°C) (T_f - 3.00)°C

$$(T_f - 3.00)^\circ\text{C} = \frac{(85.0 \text{ kJ})\left(\frac{10^3 \text{ J}}{1 \text{ kJ}}\right)}{(295 \text{ g})\left(\frac{0.900 \text{ J}}{\text{g } ^\circ\text{C}}\right)} = 320.15^\circ\text{C (unrounded)}$$

T_f = **323°C**

6.37 Since the bolts have the same mass, and one must cool as the other heats, the intuitive answer is [(T_1 + T_2) / 2] = **77.5°C**.

6.39 Use the same procedure in 6.37, but mass is now a factor because the two water samples are not identical. Both volumes are converted to mass using the density (e.g., 85 mL x 1.00 g/mL = 85 g). Convert °C temperatures (22°C, 82°C) to Kelvin temperatures (295 K, 355 K).

$-q_{lost} = q_{gained}$

- [85 mL (1.00 g/mL)] (4.184 J/g°C) (T_f - 82)°C = [165 mL (1.00 g/mL)] (4.184 J/g°C) (T_f - 22)°C

$- [85](T_f - 82) = [165] (T_f - 22)$

$6970 - 85\ T_f = 165\ T_f - 3630$

$6970 + 3630 = 165\ T_f + 85\ T_f$

$10600. = 250.\ T_f$

$T_f = (10600. / 250.) = 42.4 =$ **42°C**

6.41 Plan: Heat gained by water plus heat lost by copper tubing must equal zero, so $q_{water} = -q_{copper}$. However, some heat will be lost to the insulated container.

Solution:

$- q_{lost} = q_{gained} = q_{water} + q_{calorimeter}$

$-$ (505 g Cu) (0.387 J/g°C) (T_f - 99.9)°C = (59.8 g H_2O) (4.184 J/g°C) (T_f - 24.8)°C + (10.0 J/°C) (T_f - 24.8)°C

$- (195.435) (T_f - 99.9) = (250.2032) (T_f - 24.8) + (10.0) (T_f - 24.8)$

$19523.9565 - 195.435\ T_f = 250.2032\ T_f - 6205.03936 + 10.0\ T_f - 248.$

$19523.9565 + 6205.03936 + 248 = 250.2032\ T_f + 195.435\ T_f + 10.0\ T_f$

$19523.9565 + 6205.03936 + 248 = 250.2032\ T_f + 195.435\ T_f + 10.0\ T_f$

$25976.99586 = (455.6382)\ T_f$

$T_f = 25976.99586 / (455.6382) = 57.0123 =$ **57.0°C**