South Pasadena • AP Chemistry

Name______________________________
Period __ Date __/__/__

6 • Energy and Chemical Reactions

CALORIMETRY

Measuring heat (formerly measured in calories) is called **calorimetry**. Now we measure heat energy in Joules (J). The equation we use is:

$$q = m \cdot C \cdot \Delta T$$

q = heat energy
m = mass of water
C = the specific heat capacity
ΔT = the change in temperature (in °C or K)

1. Water has a specific heat capacity of 4.184 J/g·°C.
   This means it takes 4.184 J to heat 1.00 gram of water 1.00°C.
   a) How much energy will it take to heat 10.0 grams of water 1°C? _______________
   b) How much energy is needed to heat 30.0 g H_2O from 10.0 °C to 50.0 °C? _____________

2. Let's try a standard **calorimetry** problem.
   A pot of water (2.5 Liters of water) initially at 25.0°C is heated to boiling (100.°C).
   How much energy (in J) is needed to heat the water? (The density of water is 1 g/mL.)

   What would this amount of heat be in kJ? ____________

3. What amount of heat is *released* when 175 g of water *cools* from 100.°C to room temperature, 20.0 °C?

4. We don't always have to warm up or cool down water. The specific heat capacity of copper metal is 0.39 J/g·°C. It is ______________ (easier/more difficult) to heat up copper than to heat up water.

   How much energy would it take to heat up a 5.20 g sample of copper from 20.0 °C to 100.°C?

5. If 300. J of heat energy were used to heat up a 5.00 gram sample of copper metal and a 5.00 gram sample of water both starting at 10.0°C, calculate the final temperature of each sample?

Signs of ΔT and q:

– q means heat is **released**. + q means heat is **absorbed**.

ΔT is always ***final*** temperature – ***initial*** temperature.

If something is getting **hotter** (10° → 30°) the ΔT is 30 – 10 = + 20°. (heat is **absorbed**)

If something is getting **cooler** (75° → 25°) the ΔT is 25 – 75 = **–** 50°. (heat is **released**)

6. Suppose we mix 90.0 grams of **hot water** (90.0°C) with 10.0 grams of **cold water** (10.0°C).
   Let x = the final temperature. C = 4.184 J/g·°C
   a. Set up an expression for the energy ***released*** (q) by the hot water ($\Delta q_{hot} = m_{hot}C\Delta T_{hot}$)

   b. Set up an expression for the energy ***absorbed*** (q) by the cold water ($\Delta q_{cold} = m_{cold}C\Delta T_{cold}$)

   c. Knowing that the heat released = **–** heat absorbed, combine the two expressions and solve for x.

7. We don't always have to use water. Let's use some **aluminum shot**.

"shot" are these little pellets.

175 grams of hot aluminum (100.°C) is dropped into an insulated cup that contains 40.0 mL of ice cold water (0.0°C). Follow the example above to determine the final temperature, x.

a. Set up an expression for the heat lost by the aluminum (C=0.900 J/g·°C)

b. Set up an expression for the heat gained by the cold water.

c. Put the two expressions together (don't forget to change one of the signs) and solve for x.

8. **Somewhat Confusing Definitions:**
   There are several terms used in this chapter that sound very similar. Use the data provided to calculate each of them to clarify the differences. I've added some "Notes" that I hope will help.
   74.8 J of heat is required to raise the temperature of 18.69 g of silver from 10.0°C to 27.0°C.
   a. What is the **heat capacity** of the silver sample? (J/°C)
   Note: This is a useful value only for this specific sample of silver.

   b. What is the **specific heat capacity** of silver? (J/g·°C)
   Note: This is a useful value for ***any*** sample of silver that is heated or cooled. This is equivalent to the 4.184 $J \cdot g^{-1} \cdot °C^{-1}$ that we use for water. This value is also called the **specific heat**.