key

Energy and Chemical Reactions

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

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

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? 41.84 J
   b) How much energy is needed to heat 30.0 g H_2O from 10.0 °C to 50.0 °C? 5020 J
   $q = (4.184\ J/g°C)(30.0g)(50-10)$

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.) = 1000 g/L
   $2.5L = 2500g$
   $q = (2500g)(4.184)(100-25)$
   $q = 784500\ J \rightarrow 7.8 \times 10^5\ J$
   What would this amount of heat be in kJ? 7.8×10^2 kJ
   $\frac{7.8 \times 10^5 J}{} \times \frac{1\ kJ}{1000\ J} = 7.8 \times 10^2\ kJ$

3. What amount of heat is *released* when 175 g of water *cools* from 100.°C to room temperature, 20.0 °C?
   $q = (175g)(4.184)(20-100)$
   $q = -58600\ J$

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 (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?
   $q = (5.20g)(0.39\ J/g°C)(100-20)$
   $q = 162\ J$

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?
   copper
   $300.J = (5.00)(.39)(x-10)$
   $x = 163°C$
   water
   $300.J = (5.00)(4.184)(x-10)$
   $x = 24.3°C$

Signs of DT and q:

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

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

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

If something is getting **cooler** (75° ® 25°) the DT 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 ($Dq_{hot} = m_{hot}CDT_{hot}$)

$q = (90.0g)(4.184)(x - 90)$

b. Set up an expression for the energy ***absorbed*** (q) by the cold water ($Dq_{cold} = m_{cold}CDT_{cold}$)

$q = (10.0g)(4.184)(x - 10)$

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

$-(90.0g \cdot 4.184 \cdot (x-90)) = (10.0g)(4.184)(x-10)$

$-376.56x + 33890.4 = 41.84x - 418.4$

$-418.4x = -34308.8$

$x = 82°C$

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)

$q = (175g)(.900 J/g°C)(x - 100)$

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

$q = (40.0g)(4.184 J/g°C)(x - 0)$

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

$-(175g \cdot .900 \cdot (x-100)) = 40.0g \cdot 4.184 (x \cdot 0)$

$-157.5x + 15750 = 167.36x$

$x = 48.5°C$

8. **Somewhat Confusing Definitions:** *look at the units to figure out calculation*
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.

$\frac{q}{\Delta T}$ $\quad \frac{74.8\,J}{27.0-10.0} = 4.40\ J/°C$

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.**

$q = mC\Delta T$

$74.8\,J = (18.69\,g)(C)(27.0-10.0°C)$

$C = 0.235\ J/g°C$