PQ 11 Heat

Q and A

Q1

When 3.0 kg of water is cooled from 80.0°C to 10.0°C, how much heat energy is lost

$$C_{p,w} = Q / m \cdot \Delta T$$

$$Q = C_{p,w} (m) \Delta T$$

$$= (4186)(3)(10-80)$$

$$= -879{,}060 \text{ J}$$

Q2

How much heat is needed to raise a 0.30 kg piece of aluminum from 30.°C to 150°C?

$$Q = c_{p,a}(m_A)(\Delta T_A)$$

$$= 899(.3)(150-30)$$

$$= +32{,}364\ J$$

Q3

Calculate the temperature change when:

a) 10.0 kg of water loses 232 kJ of heat.

$$232\,kJ \times \frac{1000\,J}{1\,kJ} = 232{,}000\,J$$

$$-232{,}000\,J = (4186)(10.0\,Kg)\,\Delta T$$

$$\frac{-232{,}000}{41860} = \frac{41860\,\Delta T}{41860}$$

$$-5.5^{\circ}C = \Delta T$$

Q3 continued

b) 1.96 kJ of heat are added to 500. g of copper.

$$1.96\ KJ \cdot \frac{1000\ J}{1\ KJ} = 1960\ J \qquad 500g \cdot \frac{1\ Kg}{1000\ g} = .5\ Kg$$

$$1960\ J = (387)(.5\ Kg)\ \Delta T$$

$$\frac{1960\ J}{193.5} = \frac{193.5\ \Delta T}{193.5}$$

$$10.1°C = \Delta T$$

Q4

A piece of metal weighing 59.047 g was heated to 100.0 °C and then put it into 100.0 mL of water (initially at 23.7 °C). The metal and water were allowed to come to an equilibrium temperature, determined to be 27.8 °C. Assuming no heat lost to the environment, calculate the specific heat of the metal.

$q_{metal} = q_{water}$ (mass) (Δt) (C_p) = (mass) (Δt) (C_p)

(59.047 g) (72.2 °C) (x) = (100.0 g) (4.1 °C) (4.184 J g^{-1} $°C^{-1}$)

x = 0.402 J g^{-1} $°C^{-1}$

Q5

- A piece of iron with a mass of 21.5 g at a temp of 100.0 °C is dropped into an insulated container of water. The mass of the water is 132.0 g and its temperature before adding the iron is 20.0 °C. What will be the final temp of the system? Specific heat of iron is 0.449 kJ/kg K.

Q5 continued

- qlost, metal = qgained, water we write
- (mass) (Δt) (Cp, metal) = (mass) (Δt) (Cp, water) 2) Substituting:
- (21.5) (100 - x) (0.449) = (132.0) (x - 20) (4.184) Some explanation:
- a) 100 - x is the Δt for the metal; it starts at 100.0 °C and drops to some unknown, final value.
  b) x - 20 is the Δt for the water; it starts at 20.0 °C and rises to some unknown, final value.
  c) Since both metal and water wind up at the same ending value, we need to use only one unknown for the two Δt expressions. 3) A wee bit of algebra:
- (2150 - 21.5x) (0.449) = (132x - 2640) (4.184) 965.35 - 9.6535x = 552.288x - 11045.76
- 561.9415x = 12011.11
- To 3 sig figs, the answer is 21.4 °C.

Q6

- A 12.48 g sample of an unknown metal, heated to 99.0 °C was then plunged into 50.0 mL of 25.0 °C water. The temperature of the water rose to 28.1 °C. Assuming no loss of energy to the surroundings:
- 1. How many joules of energy did the water absorb?
  2. How many joules of energy did the metal lose?
  3. What is the heat capacity of the metal?
  4. What is the specific heat of metal?

Q6 continued

- 1) $q = (50.0\ g)\ (3.1\ °C)\ (4.184\ J\ g^{-1}\ °C^{-1}) = 648.52\ J$
- 2) 648.52 J
- 3) 648.52 J / 70.9 °C = 9.147 J/°C
- 4) 9.147 J/°C divided by 12.48 g = 0.733 $J\ g^{-1}\ °C^{-1}$

Q7

- A 43.2 g block of an unknown metal at 89.0 °C was dropped into an insulated vessel containing 43.00 g of ice and 26.00 g of water at 0 °C. After the system had reached equilibrium it was determined that 9.15 g of the ice had melted. What is the specific heat of the metal? (The heat of fusion of ice = 334.166 J g^{-1}.)

Q7 continued

- 1) Determine heat gained by the ice that melted:
- 9.15 g times 334.166 J g^{-1} = 3057.62 J
- 2) Substitue and solve for the specific heat:
- q = (mass) (Δt) (Cp, metal)
- 3057.62 J = (43.2 g) (89.0 °C) (x)
- x = 0.795 J g^{-1} $°C^{-1}$

Q8

- A 35.0 g block of metal at 80.0 °C is added to a mixture of 100.0 g of water and 15.0 g of ice in an isolated container. All the ice melted and the temperature in the container rose to 10.0 °C. What is the specific heat of the metal?

Q8 continued

- 1) Determine heat required to melt the ice:
- q = (15.0 g) (334.166 J g^{-1}) = 5012.49 J Note that the 100 g of water is not mentioned yet.
- 2) Determine heat need to raise 115 g of water from 0 to 10.0 °C:
- q = (115 g) (10.0 °C) (4.184 J g^{-1} $°C^{-1}$) = 4811.6 J Note the inclusion of the melted 15 g of ice. Also, notice that the water was at zero °C. We know this from the presence of the ice.
- 3) Determine the specific heat of the metal:
- (5012.49 J + 4811.6 J) = (35.0 g) (70.0 °C) (x) x = 4.01 J g^{-1} $°C^{-1}$

Q9

- A 12.48 g sample of an unknown metal is heated to 99.0 °C and then was plunged into 50.0 mL of 25.0 °C water. The temperature of the water rose to 28.1°C.
- (a) How many joules of energy did the water absorb?
  (b) How many joules of energy did the metal lose?
  (c) What is the heat capacity of the metal?
  (d) What is the specific heat capacity of the metal?

Q9 continued

- 1) Solution to (a):
- q = (50.0 g) (3.1 °C) (4.181 J g^{-1} $°C^{-1}$) = 648.52 J I used 50.0 g because the density of water is 1.00 g/mL and I had 50.0 mL of water.
- 2) Solution to (b):
- q = 648.52 J We assume all heat absorbed by the water was lost by the metal. We assume no loss of heat energy to the outside during the transfer.
- 3) Solution to (c):
- 648.52 J / 74.0 °C = 8.76 J/°C (or 8.76 J/K) 4)
- Solution to (d):
- (50.0 g) (3.1 °C) (4.181 J g^{-1} $°C^{-1}$) = (12.48 g) (74.0 °C) (x) Solve for x!!!!!!!

Q10 continued

- A 25.6 g piece of metal was taken from a beaker of boiling water at 100.0 °C and placed directly into a calorimeter holding 100.0 mL of water at 25.0 °C. The calorimeter heat capacity is 1.23 J/K. Given that the final temperature at thermal equilibrium is 26.2 °C, determine the specific heat capacity of the metal.

Q10 continued

- 1) We know this:
- $q_{lost, metal} = q_{gained}$
- 2) However, energy is gained by two different entities (the water and the calorimeter itself). Therefore:
- $q_{lost, metal} = q_{gained, water} + q_{gained, calorimeter}$ 3) Substituting, we have:
- (mass) (Δt) ($C_{p, metal}$) = (mass) (Δt) ($C_{p, water}$) + (Δt of water) (calorimeter constant) 4) Putting values into place and solving:
- (25.6 g) (73.8 °C) (x) = (100.0 g) (1.2 °C) (4.184 J/g °C) + (1.2 °C) (1.23 J/K) x = 0.266 J/g °C