<u>Solutions to the Extra Problems for Chapter 13</u>

1. The lead <u>gained 97,000 J</u>. We can use Equations (13.1) and (13.2) to solve this, but the units need to be compatible. Mass is in kg, but the mass unit in specific heat capacity is in grams. Thus, we need to convert 75.0 kg into 7.50×10^4 g.

$$\Delta T = T_{final} - T_{initial} = 24.1\ °C - 16.0\ °C = 8.1\ °C$$

Now we can just plug the numbers into Equation (13.1):

$$q = m \cdot c \cdot \Delta T = (7.50 \times 10^4\ \cancel{g}) \cdot (0.160\ \frac{J}{\cancel{g} \cdot \cancel{°C}}) \cdot (8.1\ \cancel{°C}) = 97{,}000\ J$$

The fact that the heat is positive tells you heat was gained, which makes sense, since the temperature increased.

2. <u>The final temperature is 27.5 °C</u>. We can relate the heat gained to the change in temperature using Equation (13.1).

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

Dividing both sides by m and c gives us:

$$\Delta T = \frac{q}{m \cdot c}$$

Before we can use the equation, however, we need to remember that the heat was gained. That means q is positive. Also, you are supposed to have the specific heat of water (4.184 J/g·°C) memorized.

$$\Delta T = \frac{q}{m \cdot c} = \frac{1{,}600\ \cancel{J}}{(150.0\ \cancel{g}) \cdot (4.184\ \frac{\cancel{J}}{\cancel{g} \cdot °C})} = 2.5\ °C$$

We have the initial temperature, so now that we have ΔT, we can find the final temperature.

$$\Delta T = T_{final} - T_{initial}$$

Adding $T_{initial}$ to both sides gives us:

$$T_{final} = \Delta T + T_{initial} = 2.5\ °C + 25.0\ °C = 27.5\ °C$$

Since we are adding here, we have to look at precision. The least precise measurement has its last significant figure in the tenths place, so our answer must have its last significant figure in the tenths place.

3. <u>The specific heat is 0.97 J/g·°C</u>. We can relate the heat lost to the specific heat capacity using Equation (13.1).

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

Dividing both sides by m and ΔT gives us:

$$c = \frac{q}{m \cdot \Delta T}$$

We will need to get ΔT first:

$$\Delta T = T_{final} - T_{initial} = 24.6\ °C - 94.3\ °C = -69.7\ °C$$

Since the problem wants the answer in J/g·°C, we also have to convert 1.2 kJ into 1,200 J. However, remember that this heat is lost, so it is negative:

$$c = \frac{-1{,}200\ J}{17.8\ g \cdot (-69.7\ °C)} = 0.97\ \frac{J}{g \cdot °C}$$

4. <u>The specific heat is 0.36 J/g·°C</u>. In all calorimetry experiments, we need to find two of the three q's in Equation (13.3). We know that $q_{calorimeter} = 0$, because we are told to ignore the calorimeter. We also have everything we need to determine q_{liquid}, since we have the specific heat capacity of water (4.184 J/g·°C) memorized.

$$\Delta T_{liquid} = T_{final} - T_{initial} = 28.1\ °C - 24.4\ °C = 3.7\ °C$$

Now get q for the water:

$$q_{liquid} = m \cdot c \cdot \Delta T = (250.0\ \cancel{g}) \cdot (4.184\ \frac{J}{\cancel{g} \cdot \cancel{°C}}) \cdot (3.7\ \cancel{°C}) = 3{,}900\ J$$

We now have q_{liquid} (3,900 J) and $q_{calorimeter}$ (0). We can put them into Equation (13.3) to figure out q_{object}.

$$-q_{object} = q_{liquid} + q_{calorimeter} = 3{,}900\ J + 0 = 3{,}900\ J$$

$$q_{object} = -3{,}900\ J$$

Now that we have q_{object}, we can figure out its specific heat capacity using Equation (13.1). However, we need to know ΔT to use that equation. The problem tells us the initial temperature of the metal (100.0 °C). The metal and liquid always have the same final temperature, so the final temperature of the metal is 28.1 °C, as it was for the liquid. That allows us to get ΔT for the object:

$$\Delta T_{object} = T_{final} - T_{initial} = 28.1\ °C - 100.0\ °C = -71.9\ °C$$

The problem gave us the mass of the object (150.0 g), and we figured out q for the object (-3,900 J), so now we can use Equation (13.1) to determine the specific heat capacity:

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

Dividing both sides by m and ΔT gives us:

$$c = \frac{q}{m \cdot \Delta T} = \frac{-3{,}900 \text{ J}}{(150.0 \text{ g}) \cdot (-71.9\ °C)} = 0.36 \frac{\text{J}}{\text{g} \cdot °C}$$

5. <u>The starting temperature was 61 °C</u>. Once again, we have to figure out two of the three q's in Equation (13.3). In this case, we know $q_{calorimeter} = 0$, because we are told to ignore the calorimeter. We can calculate q_{liquid}, because know the mass of the water, the specific heat, and the temperature change.

$$\Delta T_{liquid} = T_{final} - T_{initial} = 29.0\ °C - 24.0\ °C = 5.0\ °C$$

We can now get q for the water:

$$q_{liquid} = m \cdot c \cdot \Delta T = (150.0\ \cancel{\text{g}}) \cdot (4.184 \frac{\text{J}}{\cancel{\text{g}} \cdot \cancel{°C}}) \cdot (5.0\ \cancel{°C}) = 3{,}100 \text{ J}$$

We can use these two q's to figure out the q of the object:

$$-q_{object} = q_{liquid} + q_{calorimeter} = 3{,}100 \text{ J} + 0 \text{ J} = 3{,}100 \text{ J}$$

$$q_{object} = -3{,}100 \text{ J}$$

Now we can get the ΔT of the object:

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

$$\Delta T = \frac{q}{m \cdot c}$$

$$\Delta T = \frac{q}{m \cdot c} = \frac{-3{,}100\ \cancel{\text{J}}}{(115.0\ \cancel{\text{g}}) \cdot (0.840 \frac{\cancel{\text{J}}}{\cancel{\text{g}} \cdot °C})} = -32\ °C$$

We have the initial temperature, so now that we have ΔT, we can find the final temperature.

$$\Delta T = T_{final} - T_{initial}$$

Now remember, the final temperature of the object is the same as that of the liquid. Adding $T_{initial}$ to both sides gives us:

$$T_{initial} = T_{final} - \Delta T = 29.0°C - -32.0 = 61°C$$

6. <u>The object gained 229 J</u>. In this case, we are dealing with heat capacity, not specific heat capacity. Thus, mass is already taken into account. Other than that, this is just like any other problem relating heat and temperature. That means we need to get ΔT:

$$\Delta T = T_{final} - T_{initial} = 45.7 - 25.1 = 20.6$$

Now we can figure out q:

$$q = C \cdot \Delta T = 11.1 \frac{J}{\cancel{°C}} \cdot 20.6 \cancel{°C} = 229 \text{ J}$$

Since q is positive, the object gained the heat, which makes sense, since its temperature increased.

7. <u>The specific heat is 0.42 J/g·°C</u>. In this problem, we can't ignore the calorimeter. However, its q is easy to calculate, because it has the same initial and final temperatures as the liquid.

$$\Delta T = T_{final} - T_{initial} = 28.5 \text{ °C} - 24.4 \text{ °C} = 4.1 \text{ °C}$$

$$q_{calorimeter} = C \cdot \Delta T = 1{,}170 \frac{J}{\cancel{°C}} \cdot 4.1 \cancel{°C} = 4{,}800 \text{ J}$$

We also have all we need to calculate q_{liquid}, and its ΔT is the same as that of the calorimeter.

$$q_{liquid} = m \cdot c \cdot \Delta T = (250.0 \cancel{g}) \cdot (4.184 \frac{J}{\cancel{g} \cdot \cancel{°C}}) \cdot (4.1 \cancel{°C}) = 4{,}300 \text{ J}$$

We can now figure out the q of the object:

$$-q_{object} = q_{liquid} + q_{calorimeter} = 4{,}300 \text{ J} + 4{,}800 \text{ J}$$

$$q_{object} = -9{,}100 \text{ J}$$

The metal starts out at 125.0 °C and ends up at the same final temperature as the water, which is 28.5 °C.

$$\Delta T_{object} = T_{final} - T_{initial} = 28.5 \text{ °C} - 125.0 \text{ °C} = -96.5 \text{ °C}$$

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

Dividing both sides by m and ΔT gives us:

$$c = \frac{q}{m \cdot \Delta T} = \frac{-9{,}100 \text{ J}}{(225.0 \text{ g}) \cdot (-96.5 \text{ °C})} = 0.42 \frac{J}{g \cdot °C}$$

8. <u>The heat capacity is 200 J/°C</u>. We can't figure out the q of the calorimeter, because we don't have its heat capacity. However, we have enough to calculate the q of the liquid and the object:

$$\Delta T_{object} = T_{final} - T_{initial} = 28.5 \text{ °C} - 100.0 \text{ °C} = -71.5 \text{ °C}$$

$$q_{object} = m \cdot c \cdot \Delta T = (275.0 \cancel{g}) \cdot (0.240 \frac{J}{\cancel{g} \cdot \cancel{°C}}) \cdot (-71.5 \cancel{°C}) = -4{,}720 \text{ J}$$

We can also figure out q_{liquid}:

$$\Delta T_{liquid} = T_{final} - T_{initial} = 28.5\ °C - 25.3\ °C = 3.2\ °C$$

$$q_{liquid} = m \cdot c \cdot \Delta T = (300.0\ \cancel{g}) \cdot (4.184\ \frac{J}{\cancel{g} \cdot \cancel{°C}}) \cdot (3.2\ \cancel{°C}) = 4.0 \times 10^{3}\ J$$

We can use these two q's to calculate the q of the calorimeter.

$$-q_{object} = q_{liquid} + q_{calorimeter}$$

$$q_{calorimeter} = -q_{object} - q_{liquid} = -(-4{,}720\ J) - 4.0 \times 10^{3}\ J = 700\ J$$

The significant figures are worth discussing here. We are subtracting, so we need to look at decimal place. 4,720 has its last significant figure in the tens place. What about 4.0×10^3? It is the same as 4,000 with the first zero significant. That puts its last significant figure in the hundreds place. Thus, it is the least precise number, so our answer must be rounded to the hundreds place.

We already calculated the ΔT of the liquid (3.2 °C), and it is the same for the calorimeter so we can now solve for the heat capacity:

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

Dividing both sides by ΔT gives us:

$$C = \frac{q}{\Delta T} = \frac{700\ J}{3.2\ °C} = 200\ \frac{J}{°C}$$

9. <u>It takes 453,000 J</u>. We are using Equation (13.5) here, because we are dealing with latent heat. However, we need to convert the mass to grams, because that's the mass unit in the latent heat. That means the mass is 2,210 g. In addition, we have to use the right latent heat. We are dealing with melting here, so we use L_f, the latent heat of fusion.

$$q = m \cdot L = (2{,}210\ \cancel{g}) \cdot 205\ \frac{J}{\cancel{g}} = 453{,}000\ J$$

10. <u>It absorbed 349,000 J</u>. This is a multiple-step process. First, we have to heat the ice cube to 0 °C (which is exact) to allow it to melt. That requires using Equation (13.1) with the specific heat capacity of ice:

$$\Delta T_{ice} = T_{final} - T_{initial} = 0\ °C - (-11.0\ °C) = 11.0\ °C$$

$$q = m \cdot c \cdot \Delta T = (115.0\ \cancel{g}) \cdot (2.093\ \frac{J}{\cancel{g} \cdot \cancel{°C}}) \cdot (11.0\ \cancel{°C}) = 2{,}650\ J$$

Now that it is at 0 °C, it can melt. For that, we need to use the latent heat of fusion:

$$q = m \cdot L = (115.0\ \cancel{g}) \cdot 334\ \frac{J}{\cancel{g}} = 38{,}400\ J$$

Now we can heat it up to 100 °C (which is exact) so it can boil:

$$\Delta T_{water} = T_{final} - T_{initial} = 100\ °C - 0\ °C = 100\ °C$$

Since both numbers are exact, the ΔT is as well.

$$q = m \cdot c \cdot \Delta T = (115.0\ \cancel{g}) \cdot (4.184\ \frac{J}{\cancel{g} \cdot \cancel{°C}}) \cdot (100\ \cancel{°C}) = 48{,}120\ J$$

Because the 100 is exact, the mass and specific heat determine the significant figures. Now that the water is at 100 °C, we can vaporize it using the latent heat of vaporization:

$$q = m \cdot L = (115.0\ \cancel{g}) \cdot 2{,}260\ \frac{J}{\cancel{g}} = 2.60 \times 10^5\ J$$

The total heat is the sum of the individual heats:

$$q_{total} = 2{,}650\ J + 38{,}400\ J + 48{,}120\ J + 2.60 \times 10^5 = 349{,}000\ J$$

Once again, the significant figures are worth discussing. 2.60×10^5 is 260,000 with the first zero significant. Thus, its last significant figure is in the thousands place. That makes it the least precise number, which limits the answer to the thousands place.