Physics 1401
Homework Solutions - Walker, Chapter 18

Problems

1. First law of thermodynamics:

$$\Delta U = Q - W \ .$$

If swimmer does 6.7×10^5 J of work (on the swimmer's surroundings), then $W = 6.7\times10^5$ J. If the swimmer *gives off* 4.1×10^5 J of heat, then $Q = -4.1\times10^5$ J. (Remember, Q is the heat *into* the system. Considering the swimmer as the thermodynamic system, if the swimmer *gives off* heat, then Q is *negative*.) So the change in the swimmer's internal energy is:

$$\Delta U = -4.1\times10^5 \text{ J} - 6.7\times10^5 \text{ J} = -10.8\times10^5 \text{ J} \ .$$

3. (a.) If 42 J of work are done *on* the system, then the work the *system* does (on its *surroundings*) is:

$$W = -42 \text{ J}$$

If 77 J of heat are added *to* the system, then:

$$Q = 77 \text{ J}$$

So:

$$\Delta U = Q - W = 77 \text{ J} - \left(-42 \text{ J}\right) = 119 \text{ J}$$

(b.) If the system does 42 J of work on its surroundings, then:

$$W = 42 \text{ J}$$

If 77 J of heat are added to the system, then:

$$Q = 77 \text{ J}$$

So:

$$\Delta U = Q - W = 35 \text{ J}$$

(c.) If the system's internal energy *decreases* by 120 J, then:

$$\Delta U = -120 \text{ J}$$

If the system performs 120 J of work on its surroundings, then:

$$W = 120 \text{ J}$$

So the head added to the system is:

$$Q = \Delta U + W = 0 \text{ J}$$

5. (a.) The amount of heat required to evaporate 0.110 kg of water is:

$$Q = m_w \left(L_v\right)_w = \left(0.110 \text{ kg}\right)\left(2.26\times10^6 \text{ J/kg}\right) = 2.486\times10^5 \text{ J}$$

This amount of heat *leaves* the basketball player and goes *into* her surroundings. Since the "*Q*" in the first law of thermodynamics is defined to be the heat that goes *into the system*, then, from the point of view of the basketball player as the system:

$$Q = -2.486\times10^5 \text{ J}$$

So if she does 2.13×10^5 J of work on her surroundings, then the change in her internal energy must be:

$$\Delta U = Q - W = -2.486\times10^5 \text{ J} - 2.13\times10^5 \text{ J}$$

$$\Delta U = -4.616\times10^5 \text{ J} = -4.62\times10^5 \text{ J} \ ,$$

keeping 3 sig figs.

(b.) The number of *joules* the player has converted to work is 2.13×10^5 J. And one "nutritional calorie" (i.e., a *dietician's* "Calorie") is equivalent to 4186 J:

$$1 \text{ Calorie} = 4186 \text{ J}$$

(Recall Sec. 16-4.) So:

$$W = \frac{\left(2.13\times10^{5}\ \text{J}\right)}{\left(4186\,\dfrac{\text{J}}{\text{Cal}}\right)} = 50.9\ \text{Cal}$$

The number of Calories the player has converted to heat is:

$$Q = \frac{2.486\times10^{5}\ \text{J}}{4186\,\dfrac{\text{J}}{\text{Cal}}} = 59.4\ \text{Cal}$$

So the *total* number of Calories that go into *both* W *and* Q is:

$$W + Q = 110\ \text{Cal}$$