11.44 $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

$$\boldsymbol{V_2} = \frac{P_1V_1T_2}{P_2T_1} = \frac{(572\text{ mmHg})(6.15\text{ L})(273\text{ K})}{(760\text{ mmHg})(308\text{ K})} = \mathbf{4.10\ L}$$

11.50 The density is given by:

$$\textbf{density} = \frac{\text{mass}}{\text{volume}} = \frac{4.65\text{ g}}{2.10\text{ L}} = \mathbf{2.21\ g/L}$$

Solving for the molar mass:

$$\textbf{molar mass} = \frac{dRT}{P} = \frac{(2.21\text{ g/L})\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(300\text{ K})}{(1.00\text{ atm})} = \mathbf{54.4\ g/mol}$$

11.70 If we can calculate the moles of H_2 gas collected, we can determine the amount of Na that must have reacted. We can calculate the moles of H_2 gas using the ideal gas equation.

$$P_{H_2} = P_{\text{Total}} - P_{H_2O} = 1.00\text{ atm} - 0.0313\text{ atm} = 0.97\text{ atm}$$

The number of moles of hydrogen gas collected is:

$$n_{H_2} = \frac{P_{H_2}V}{RT} = \frac{(0.97\text{ atm})(0.246\text{ L})}{\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(25 + 273)\text{K}} = 0.0098\text{ mol }H_2$$

The balanced equation shows a 2:1 mole ratio between Na and H_2. The mass of Na consumed in the reaction is:

$$\textbf{? g Na} = 0.0098\text{ mol }H_2 \times \frac{2\text{ mol Na}}{1\text{ mol }H_2} \times \frac{22.99\text{ g Na}}{1\text{ mol Na}} = \mathbf{0.45\ g\ Na}$$

11.86 The separation factor is given by:

$$s = \frac{r_1}{r_2} = \sqrt{\frac{\mathcal{M}_2}{\mathcal{M}_1}}$$

This equation is the same as Graham's Law, Equation 11.15 of the text. For $^{235}UF_6$ and $^{238}UF_6$, we have:

$$s = \sqrt{\frac{238 + (6)(19.00)}{235 + (6)(19.00)}} = \mathbf{1.0043}$$

This is a very small separation factor, which is why many (thousands) stages of effusion are needed to enrich ^{235}U.

12.62 A phase change is when a material changes from one phase (solid, liquid, gas) to another. The possible phase changes are fusion (melting), freezing, vaporization (evaporation), condensation, sublimation and deposition. See summary on page 495.

12.80 The molar heat of vaporization of water is 40.79 kJ/mol. One must find the number of moles of water in the sample:

$$\text{Moles } H_2O = 150.2\text{ g } H_2O \times \frac{1\text{ mol } H_2O}{18.02\text{ g } H_2O} = 8.34\text{ mol } H_2O$$

We can then calculate the amount of heat.

$$\boldsymbol{q} = 8.34\text{ mol } H_2O \times \frac{40.79\text{ kJ}}{1\text{ mol } H_2O} = \mathbf{340.\ kJ}$$

12.81 *Step 1:* Warming ice to the melting point.

$$q_1 = ms\Delta T = (866\text{ g } H_2O)(2.03\text{ J/g°C})[0 - (-15)\text{°C}] = 26.4\text{ kJ}$$

Step 2: Converting ice at the melting point to liquid water at 0°C. (See Table 12.8 of the text for the heat of fusion of water.)

$$q_2 = 866\text{ g } H_2O \times \frac{1\text{ mol}}{18.02\text{ g } H_2O} \times \frac{6.01\text{ kJ}}{1\text{ mol}} = 289\text{ kJ}$$

Step 3: Heating water from 0°C to 100°C.

$$q_3 = ms\Delta T = (866\text{ g } H_2O)(4.184\text{ J/g°C})[(100 - 0)\text{°C}] = 362\text{ kJ}$$

Step 4: Converting water at 100°C to steam at 100°C. (See Table 12.6 of the text for the heat of vaporization of water.)

$$q_4 = 866\text{ g } H_2O \times \frac{1\text{ mol}}{18.02\text{ g } H_2O} \times \frac{40.79\text{ kJ}}{1\text{ mol}} = 1.96 \times 10^3\text{ kJ}$$

Step 5: Heating steam from 100°C to 146°C.

$$q_5 = ms\Delta T = (866\text{ g } H_2O)(1.99\text{ J/g°C})[(146 - 100)\text{°C}] = 79.3\text{ kJ}$$

$$\boldsymbol{q_{total}} = q_1 + q_2 + q_3 + q_4 + q_5 = \mathbf{2.72 \times 10^3\ kJ}$$

1st printing of the textbook has a typo in the answer to 12.81. It says 2.71×10^3 kJ; it should say 2.72×10^3 kJ.

12.97 Region labels: The region containing point A is the solid region. The region containing point B is the liquid region. The region containing point C is the gas region.

(a) Ice would melt. (If heating continues, the liquid water would eventually boil and become a vapor.)

(b) Liquid water would vaporize.

(c) Water vapor would solidify without becoming a liquid.