Solutions to the Extra Problems for Chapter 10

1. The temperature is 298.2 K. To convert to Kelvin, we use Equation (10.4):

$$K = {}^{\circ}C + 273.15 = 25.0 + 273.15 = 298.2$$

Since we are adding, the significant figures are determined by precision. 25.0 is the least precise number, because its last significant figure is in the tenths place. Thus, the answer must be to the tenths place as well.

2. The temperature is -78 °C. To convert between Kelvin and Celsius, we use Equation (10.4):

$$K = {}^{\circ}C + 273.15$$

$$195 = {}^{\circ}C + 273.15$$

$${}^{\circ}C = 195 - 273.15 = -78\ {}^{\circ}C$$

Since 195 has its last significant figure in the ones place, the answer must have its last significant digit in the ones place.

3. The volume is 86.5 mL. Since the temperature doesn't change, T_1 and T_2 are the same. That means they cancel in the combined gas law. Since 567 torr and 149.6 mL go together, we will call them P_1 and V_1. That means 981 torr is P_2.

$$\frac{P_1 \cdot V_1}{\cancel{T_1}} = \frac{P_2 \cdot V_2}{\cancel{T_2}}$$

$$(567\ \text{torr}) \cdot (149.6\ \text{mL}) = (981\ \text{torr}) \cdot V_2$$

Dividing by 981 torr solves for V_2.

$$V_2 = \frac{(567\ \cancel{\text{torr}}) \cdot (149.6\ \text{mL})}{981\ \cancel{\text{torr}}} = 86.5\ \text{mL}$$

4. The volume is 23.7 L. Since pressure doesn't change, P_1 and P_2 cancel out in the combined gas law. The volume of 16.7 liters goes with 25.0 °C, so those are V_1 and T_1. This makes 150.0 °C T_2. However, we need to convert both those temperatures to Kelvin first.

$$K = {}^{\circ}C + 273.15 = 25.0 + 273.15 = 298.2$$

$$K = {}^{\circ}C + 273.15 = 150.0 + 273.15 = 423.2$$

$$\frac{\cancel{P_1} \cdot V_1}{T_1} = \frac{\cancel{P_2} \cdot V_2}{T_2}$$

$$\frac{16.7\text{ L}}{298.2\text{ K}} = \frac{V_2}{423.2\text{ K}}$$

Multiplying by 423.2 K solves for V_2:

$$V_2 = \frac{16.7\text{ L}}{298.2\cancel{\text{K}}} \cdot (423.2\cancel{\text{K}}) = 23.7\text{ L}$$

5. <u>The volume is 140 liters</u>. In this problem, nothing is held constant. The volume of 15.6 liters goes with 25.0 °C and 1.01 atm, so those are V_1, T_1, and P_1. That means -12.1 °C is T_2 and 0.10 atm is P_2. Before we use the combined gas law, however, the temperatures must be in Kelvin.

$$K = {}^\circ C + 273.15 = 25.0 + 273.15 = 298.2$$

$$K = {}^\circ C + 273.15 = -12.1 + 273.15 = 261.1$$

$$\frac{P_1 \cdot V_1}{T_1} = \frac{P_2 \cdot V_2}{T_2}$$

$$\frac{(1.01\text{ atm}) \cdot (15.6\text{ L})}{298.2\text{ K}} = \frac{(0.10\text{ atm}) \cdot V_2}{261.1\text{ K}}$$

Multiplying by 261.1 K and dividing by 0.10 atm solves for V_2:

$$V_2 = \frac{(1.01\cancel{\text{atm}}) \cdot (15.6\text{ L}) \cdot (261.1\cancel{\text{K}})}{298.2\cancel{\text{K}} \cdot (0.10\cancel{\text{atm}})} = 140\text{ L}$$

6. <u>There are 0.0249 moles of water vapor</u>. The Ideal Gas Law relates moles to pressure, volume, and temperature, but the units need to be atm, liters, and K:

$$\frac{745\cancel{\text{torr}}}{1} \times \frac{1\text{ atm}}{760\cancel{\text{torr}}} = 0.980\text{ atm}$$

$$\frac{615\cancel{\text{mL}}}{1} \times \frac{0.001\text{ L}}{1\cancel{\text{mL}}} = 0.615\text{ L}$$

$$K = {}^\circ C + 273.15 = 21.5 + 273.15 = 294.7$$

Now we can use the Ideal Gas Law:

$$PV = nRT$$

$$(0.980\text{ atm}) \cdot (0.615\text{ L}) = n \cdot \left(0.0821\frac{\text{L} \cdot \text{atm}}{\text{mole} \cdot \text{K}}\right) \cdot (294.7\text{ K})$$

To solve for n, we need to divide both sides by 0.0821 (L·atm)/(mole·K) and 294.7 K:

$$n = \frac{(0.980\ \cancel{\text{atm}})\cdot(0.615\ \cancel{\text{L}})}{\left(0.0821\ \frac{\cancel{\text{L}}\cdot\cancel{\text{atm}}}{\text{mole}\cdot\cancel{\text{K}}}\right)\cdot(294.7\ \cancel{\text{K}})} = 0.0249 \text{ moles}$$

7. <u>The volume is 151 L</u>. There is no relationship among grams, volume, pressure, and temperature for a gas, but the Ideal Gas Law relates moles, volume, pressure, and temperature, as long as pressure is in atm and temperature is in K. Thus, we must get from grams to moles and convert the pressure and temperature to the right units.

$$\frac{100.0\ \cancel{\text{g } NH_3}}{1} \times \frac{1 \text{ mole } NH_3}{17.04\ \cancel{\text{g } NH_3}} = 5.869 \text{ moles } NH_3$$

$$\frac{869\ \cancel{\text{torr}}}{1} \times \frac{1 \text{ atm}}{760\ \cancel{\text{torr}}} = 1.14 \text{ atm}$$

$$K = {}^{\circ}C \ + \ 273.15 = \ 84.5 \ + \ 273.15 \ = \ 357.7$$

$$PV = nRT$$

$$(1.14 \text{ atm})\cdot(V) = (5.869 \text{ moles})\cdot\left(0.0821\ \frac{\text{L}\cdot\text{atm}}{\text{mole}\cdot\text{K}}\right)\cdot(357.7 \text{ K})$$

Dividing by 1.14 atm solves for V:

$$V = \frac{(5.869\ \cancel{\text{moles}})\cdot\left(0.0821\ \frac{\text{L}\cdot\cancel{\text{atm}}}{\cancel{\text{mole}}\cdot\cancel{\text{K}}}\right)\cdot(357.7\ \cancel{\text{K}})}{1.14\ \cancel{\text{atm}}} = 151 \text{ L}$$

8. <u>There are 22.5 L of hydrogen gas produced</u>. If we want to get the volume of hydrogen, we will need to figure out the moles of hydrogen. We can get that by doing stoichiometry.

$$\frac{25.0\ \cancel{\text{g Mg}}}{1} \times \frac{1 \text{ mole Mg}}{24.31\ \cancel{\text{g Mg}}} = 1.03 \text{ moles Mg}$$

Now we can use the chemical equation to convert to O_2:

$$\frac{1.03\ \cancel{\text{moles Mg}}}{1} \times \frac{1 \text{ mole } H_2}{1\ \cancel{\text{mole Mg}}} = 1.03 \text{ moles } H_2$$

We now have moles for the Ideal Gas Law, but we need to convert the temperature to Kelvin:

$$K = {}^{\circ}C + 273.15 = 25.0 + 273.15 = 298.2$$

$$PV = nRT$$

$$(1.12\text{ atm})\cdot(V) = (1.03\text{ moles})\cdot\left(0.0821\,\frac{L\cdot atm}{mole\cdot K}\right)\cdot(298.2\text{ K})$$

Dividing by 1.12 atm solves for V:

$$V = \frac{(1.03\,\cancel{\text{moles}})\cdot\left(0.0821\,\frac{L\cdot\cancel{atm}}{\cancel{mole}\cdot\cancel{K}}\right)\cdot(298.2\,\cancel{K})}{1.12\,\cancel{\text{atm}}} = 22.5\text{ L}$$

9. <u>She will need 9 L</u>. We could use the Ideal Gas Law, but since the things we are interested in are both gases, we can just use Avogadro's Law and allow the chemical equation to give us the relationship between the volumes.

$$\frac{6\,\cancel{\text{liters }NH_3}}{1}\times\frac{3\text{ liters }H_2}{2\,\cancel{\text{liters }NH_3}} = 9\text{ liters }H_2$$

10. <u>The mole fraction of hydrogen is 0.675, and the mole fraction for nitrogen is 0.325</u>. The mole fraction is related to the partial pressure and total pressure. We have the partial pressures, so we just have to get the total pressure:

$$P_{total} = 715\text{ torr} + 345\text{ torr} = 1.060\times10^3\text{ torr}$$

$$X_a = \frac{P_a}{P_{total}}$$

$$X_{hydrogen} = \frac{715\,\cancel{\text{torr}}}{1.060\times10^3\,\cancel{\text{torr}}} = 0.675$$

$$X_{nitrogen} = \frac{345\,\cancel{\text{torr}}}{1.060\times10^3\,\cancel{\text{torr}}} = 0.325$$

11. <u>The partial pressures are 1.4 atm for argon, 2.4 atm for hydrogen, and 1.2 atm for nitrogen</u>. Equation (10.9) allows us to figure out the partial pressure of each gas if we know the mole fraction. That's our first step.

$$X_a = \frac{\text{moles of component a}}{\text{total moles in the mixture}}$$

$$\text{Total moles in the mixture} = 0.54\text{ moles} + 0.92\text{ moles} + 0.45\text{ moles} = 1.91\text{ moles}$$

$$X_{argon} = \frac{0.54\ \cancel{moles}}{1.91\ \cancel{moles}} = 0.28$$

$$X_{hydrogen} = \frac{0.92\ \cancel{moles}}{1.91\ \cancel{moles}} = 0.48$$

$$X_{nitrogen} = \frac{0.45\ \cancel{moles}}{1.91\ \cancel{moles}} = 0.24$$

We can now get the partial pressures from the total pressure:

$$X_a = \frac{P_a}{P_{total}}$$

$$P_a = X_a \cdot P_{total}$$

$$P_{argon} = (0.28)\cdot(5.1 \text{ atm}) = 1.4 \text{ atm}$$

$$P_{hydrogen} = (0.48)\cdot(5.1 \text{ atm}) = 2.4 \text{ atm}$$

$$P_{nitrogen} = (0.24)\cdot(5.1 \text{ atm}) = 1.2 \text{ atm}$$

12. <u>The chemist made 0.0512 moles</u>. We can't use the Ideal Gas Law immediately, because the gas is collected over water. Thus, it is a mixture of hydrogen and water vapor. To subtract out the water vapor, we need the vapor pressure of water at 30.0 °C. The table on page 311 says it is 31.8 torr. However, to use the Ideal Gas Law, that has to be in atm.

$$\frac{31.8\ \cancel{torr}}{1} \times \frac{1 \text{ atm}}{760\ \cancel{torr}} = 0.0418 \text{ atm}$$

The pressure given is the total pressure. According to Dalton's Law of Partial Pressure:

$$P_{total} = P_a + P_b$$

$$0.921 \text{ atm} = 0.0418 \text{ atm} + P_{oxygen}$$

$$P_{oxygen} = 0.921 \text{ atm} - 0.0418 \text{ atm} = 0.879 \text{ atm}$$

To use the Ideal Gas Law, temperature must be in Kelvin.

$$K = {}^{\circ}C + 273.15 = 30.0\ {}^{\circ}C + 273.15 = 303.2$$

$$PV = nRT$$

$$(0.879\ \text{atm})\cdot(1.45\ \text{L}) = \text{n}\cdot\left(0.0821\,\frac{\text{L}\cdot\text{atm}}{\text{mole}\cdot\text{K}}\right)\cdot(303.2\ \text{K})$$

To solve for n, we need to divide both sides by 0.0821 (L·atm)/(mole·K) and 303.2 K:

$$\text{n} = \frac{(0.879\ \cancel{\text{atm}})\cdot(1.45\ \cancel{\text{L}})}{\left(0.0821\,\frac{\cancel{\text{L}}\cdot\cancel{\text{atm}}}{\text{mole}\cdot\cancel{\text{K}}}\right)\cdot(303.2\ \cancel{\text{K}})} = 0.0512\ \text{moles}$$