$\boxed{14.3}$

The formula for aspartame is $C_{14}H_{18}N_2O_5$. From the periodic table in the back of the book, we obtain the atomic weights:

C	12.011 a.m.u.
H	1.00794 a.m.u.
N	14.0067 a.m.u.
O	15.9994 a.m.u.

a. Therefore the formula weight (in atomic mass units) for aspartame is

$$14\,(12.011\text{ a.m.u.}) + 18\,(1.00794\text{ a.m.u.}) + 2\,(14.0067\text{ a.m.u.}) + 5\,(15.9994\text{ a.m.u.}) =$$

$$\boxed{294.307\text{ a.m.u.}}$$

b. So 294.307 g of aspartame equals one mole. And one mole of anything is comprised of 6.022×10^{23} molecules (Avagadro's number). This means

$$294.307\text{ g of aspartame} = 6.022 \times 10^{23}\text{ molecules}$$

Therefore the mass m of a single molecule of aspartame is

$$m = \frac{294.307\text{g}}{6.022 \times 10^{23}} = 4.887 \times 10^{-22}\text{ g} = \boxed{4.887 \times 10^{-25}\text{ kg}}$$

$\boxed{14.12}$

$P_1 = 65.0$ atm.
$T_1 = 288$ K
$P_2 = 1.00$ atm.
$T_2 = 297$ K
$V_1 = 1.00\ m^3$
$V_2 = ?$

In room # 1 $\quad P_1 V_1 = n_1 R T_1 \quad$ # 1
In room # 2 $\quad P_2 V_2 = n_2 R T_2 \quad$ # 2

Dividing eq # 2 by eq # 1 and assuming $n_1 = n_2$ (i.e. no gas escapes)

$$\frac{P_2 V_2}{P_1 V_1} = \frac{n_2 T_2}{n_1 T_1} = \frac{T_2}{T_1} \quad \# 3$$

Solving for V_2 in eq # 3, we obtain

$$V_2 = V_1 \left(\frac{P_1}{P_2}\right)\left(\frac{T_2}{T_1}\right) = 1.00 m^3 \left(\frac{65.0 atm}{1.00 atm}\right)\left(\frac{297 k}{288 k}\right) = \boxed{67.0 m^3}$$

$\boxed{14.13}$

For an ideal gas, $PV = nRT$

$$\therefore \; n = \frac{PV}{RT} = \frac{(1.1 \times 10^5 \, N/m^2)(5400 \, m^3)}{(8.314 \, J/mole\,k)(280 \, k)} = 255{,}163 \text{ moles.}$$

The formula weight for monatomic helium is 4.00260 (from the periodic table of the elements in the back of our textbook), which means that

1 mole He $= 4.00260$ g

$$\therefore \; m = (255{,}163 \text{ moles})\left(\frac{4.00260}{1} \frac{g}{mole}\right) = 1{,}021{,}317 \, g = 1{,}021.317 \, kg$$

Rounded to two significant figures, $m = \boxed{1000 \text{ kg}}$

$\boxed{14.15}$

$T_1 = 305\ k$
$V_2 = \frac{V_1}{16}$
$P_2 = 48.5\ P_1$
$T_2 = ?$

Before compression $\quad P_1 V_1 = n_1 R T_1 \quad \#1$

After compression $\quad P_2 V_2 = n_2 R T_2 \quad \#2$

Dividing eq # 2 by eq # 1, we have

$$\frac{P_2 V_2}{P_1 V_1} = \frac{n_2 R T_2}{n_1 R T_1} = \frac{n_2 T_2}{n_1 T_1} \quad \#3$$

Assuming that no molecules are gained or lost (i.e. $n_1 = n_2$), eq # 3 simplifies to

$$\frac{P_2 V_2}{P_1 V_1} = \frac{T_2}{T_1} \quad \#4$$

and we can solve for T_2:

$$T_2 = T_1 \left(\frac{P_2 V_2}{P_1 V_1}\right) = 305\,k \left[\frac{(48.5 P_1)(V_1/16)}{P_1\ V_1}\right] = 305\,k \left[(48.5)(\tfrac{1}{16})\right] = \boxed{925\ k}$$

$\boxed{14.28}$

$v_{RMS} = 650\ m/s.$
$T = ?$

From the kinetic theory of gases

$$\frac{1}{2}m\overline{v^2} = \frac{3}{2}kT \rightarrow T = \frac{m\overline{v^2}}{3k} \quad \#1.$$

To use eq #1, we will need to first determine the mass m of a single molecule of carbon dioxide (CO_2). The formula weight of CO_2 is

$$12 + 2(16) = 44$$

Therefore 1 mole CO_2 = 44 g = 6.022×10^{23} molecules.
and therefore the mass of one molecule of CO_2 is

$$m = \frac{44\,g}{6.022 \times 10^{23}} = 7.307 \times 10^{-23}\,g = 7.307 \times 10^{-26}\,kg$$

Notice that the mass is reexpressed in kilograms to be consistent with the units of the Boltzmann constant.

Now substituting into eq #1,

$$T = \frac{(7.307 \times 10^{-26}\,kg)(650\,m/s)^2}{3(1.38 \times 10^{-23}\,J/k)} = \boxed{746\ k}$$

$\boxed{14.30}$

From the kinetic theory of gases,

$$\frac{1}{2} m \overline{v^2} = \frac{3}{2} kT \quad \#1$$

Because the water and the carbon dioxide molecules are at the same temperatures, the must have the same average kinetic energies. So,

$$\frac{1}{2} m_{H_2O} \overline{v^2_{H_2O}} = \frac{1}{2} m_{CO_2} \overline{v^2_{CO_2}} \quad \#2$$

from which we obtain

$$\overline{v^2_{CO_2}} = \overline{v^2_{H_2O}} \left(\frac{m_{H_2O}}{m_{CO_2}}\right) \quad \#3$$

Taking the square root of this equation yields

$$v_{RMS_{CO_2}} = v_{RMS_{H_2O}} \sqrt{\frac{m_{H_2O}}{m_{CO_2}}} \quad \#4$$

The "m" in the equations above represent the mass of a single molecule, found by dividing the formula weight (f.w.) by Avogadro's number, N_A. So,

$$\frac{m_{H_2O}}{m_{CO_2}} = \frac{f.w.\,H_2O / \cancel{N_A}}{f.w.\,CO_2 / N_A} = \frac{f.w.\,H_2O}{f.w.\,CO_2}$$

Since $f.w._{H_2O} = 2(1) + 16 = 18$ and $f.w._{CO_2} = 12 + 2(16) = 44$

$$\frac{m_{H_2O}}{m_{CO_2}} = \frac{18}{44} \quad \#5$$

Substituting eq 5 into eq 4 along with $v_{RMS_{H_2O}} = 648\ m/s$

$$v_{RMS_{CO_2}} = 648\ m/s \sqrt{\frac{18}{44}} = 414.462\ m/s$$

Rounded to two significant figures, $v_{RMS_{CO_2}} = \boxed{410\ m/s}$

$\boxed{14.54}$

First, we can determine the temperature of the SO_2 by using the equation of state for an ideal gas

$$PV = nRT \quad \#1$$

$$T = \frac{PV}{nR} = \frac{(2.12\times10^{4}\ N/m^2)(50.0\ m^3)}{(421\ moles)(8.314\ J/mole\ k)} = 302.84\ k$$

Next, from the kinetic theory of gases equation, $\frac{1}{2}m\overline{v^2} = \frac{3}{2}kT$, we can solve for the rms speed

$$v_{RMS} = \sqrt{\overline{v^2}} = \sqrt{\frac{3kT}{m}} \quad \#2$$

The "m" in eq 2 represents the mass of a SO_2 molecule and is found by dividing the formula weight (f.w.) expressed in grams) by Avagadro's number (N_A). Since the f.w. for SO_2 equals the atomic weight of one sulfur and two oxygen atoms

$$f.w._{SO_2} = 32 + 2(16) = 64$$

So, 64g of SO_2 = 6.02×10^{23} molecules and therefore

$$m = \frac{64\ g/mole}{6.02\times10^{23}\ molecules/mole} = 1.0631\times10^{-22}\ g. = 1.0631\times10^{-25}\ kg.$$

Substituting this value into eq 2 along with the temperature and the Boltzmann constant $k = 1.38\times10^{-23}\ J/K$

$$v_{RMS} = \sqrt{\frac{3(1.38\times10^{-23}\ J/k)(302.84\ k)}{1.0631\times10^{-25}\ kg.}} = \boxed{343\ m/s}$$