

CHAPTER 13: TEMPERATURE, KINETIC THEORY, AND THE GAS LAWS

13.1 TEMPERATURE

1. *What is the Fahrenheit temperature of a person with a 39.0°C fever?*

Solution We can convert from Celsius to Fahrenheit:

$$T_{\text{F}} = \frac{9}{5}(T_{\text{C}}) + 32.0^\circ$$

$$T_{\text{F}} = \frac{9}{5}(39.0^\circ\text{C}) + 32.0^\circ\text{C} = \underline{102^\circ\text{F}}$$

So 39.0°C is equivalent to 102°F .

7. *(a) Suppose a cold front blows into your locale and drops the temperature by 40.0 Fahrenheit degrees. How many degrees Celsius does the temperature decrease when there is a 40.0°F decrease in temperature? (b) Show that any change in temperature in Fahrenheit degrees is nine-fifths the change in Celsius degrees.*

Solution (a) We can use the equation $T_{\text{C}} = \frac{5}{9}(T_{\text{F}} - 32)$ to determine the change in temperature.

$$\Delta T_{\text{C}} = T_{\text{C}2} - T_{\text{C}1} = \frac{5}{9}(T_{\text{F}2} - 32) - \frac{5}{9}(T_{\text{F}1} - 32)$$

$$= \frac{5}{9}(T_{\text{F}2} - T_{\text{F}1}) = \frac{5}{9}\Delta T_{\text{F}} = \frac{5}{9}(40^\circ) = \underline{22.2^\circ\text{C}}$$

(b) We know that $\Delta T_{\text{F}} = T_{\text{F}2} - T_{\text{F}1}$. We also know that $T_{\text{F}2} = \frac{9}{5}T_{\text{C}2} + 32$ and

$$T_{\text{F}1} = \frac{9}{5}T_{\text{C}1} + 32. \text{ So, substituting, we have } \Delta T_{\text{F}} = \left(\frac{9}{5}T_{\text{C}2} + 32\right) - \left(\frac{9}{5}T_{\text{C}1} + 32\right).$$

Partially solving and rearranging the equation, we have $\Delta T_{\text{°F}} = \frac{9}{5}(T_{\text{°C}2} - T_{\text{°C}1})$.

Therefore, $\Delta T_{\text{°F}} = \frac{9}{5} \Delta T_{\text{°C}}$.

13.2 THERMAL EXPANSION OF SOLIDS AND LIQUIDS

15. *Show that 60.0 L of gasoline originally at 15.0°C will expand to 61.1 L when it warms to 35.0°C, as claimed in Example 13.4.*

Solution We can get an expression for the change in volume using the equation $\Delta V = \beta V_0 \Delta T$, so the final volume is $V = V_0 + \Delta V = V_0(1 + \beta \Delta T)$, where $\beta = 9.50 \times 10^{-4} / ^\circ\text{C}$ for gasoline (see Table 13.2), so that

$$V' = V_0 + \beta V_0 \Delta T = 60.0 \text{ L} + (9.50 \times 10^{-4} / ^\circ\text{C})(60.0 \text{ L})(20.0^\circ\text{C}) = \underline{61.1 \text{ L}}$$

As the temperature is increased, the volume also increases.

21. *Show that $\beta \approx 3\alpha$, by calculating the change in volume ΔV of a cube with sides of length L .*

Solution From the equation $\Delta L = \alpha L_0 \Delta T$ we know that length changes with temperature. We also know that the volume of a cube is related to its length by $V = L^3$. Using the equation $V = V_0 + \Delta V$ and substituting for the sides we get $V = (L_0 + \Delta L)^3$. Then we replace ΔL with $\Delta L = \alpha L_0 \Delta T$ to get $V = (L_0 + \alpha L_0 \Delta T)^3 = L_0^3(1 + \alpha \Delta T)^3$. Since $\alpha \Delta T$ is small, we can use the binomial expansion to get $V = L_0^3(1 + 3\alpha \Delta T) = L_0^3 + 3\alpha L_0^3 \Delta T$. Rewriting the length terms in terms of volume gives $V = V_0 + \Delta V = V_0 + 3\alpha V_0 \Delta T$. By comparing forms we get $\Delta V = \beta V_0 \Delta T = 3\alpha V_0 \Delta T$. Thus, $\beta = \underline{3\alpha}$.

13.3 THE IDEAL GAS LAW

27. In the text, it was shown that $N/V = 2.68 \times 10^{25} \text{ m}^{-3}$ for gas at STP. (a) Show that this quantity is equivalent to $N/V = 2.68 \times 10^{19} \text{ cm}^{-3}$, as stated. (b) About how many atoms are there in one μm^3 (a cubic micrometer) at STP? (c) What does your answer to part (b) imply about the separation of atoms and molecules?

Solution (a) This is a units conversion problem, so

$$\frac{N}{V} = \left(\frac{2.68 \times 10^{25}}{\text{m}^3} \right) \left(\frac{1 \text{ m}}{100 \text{ cm}} \right)^3 = \underline{2.68 \times 10^{19} \text{ cm}^{-3}}$$

(b) Again, we need to convert the units:

$$\frac{N}{V} = \left(\frac{2.68 \times 10^{25}}{\text{m}^3} \right) \left(\frac{1 \text{ m}}{1.00 \times 10^6 \mu\text{m}} \right)^3 = \underline{2.68 \times 10^7 \mu\text{m}^{-3}}$$

(c) This says that atoms and molecules must be on the order of (if they were tightly

$$\text{packed}) V = \frac{N}{2.68 \times 10^7 \mu\text{m}^{-3}} = \frac{1}{2.68 \times 10^7 \mu\text{m}^{-3}} = \underline{3.73 \times 10^{-8} \mu\text{m}^3}$$

Or the average length of an atom is less than approximately
 $(3.73 \times 10^{-8} \mu\text{m}^3)^{1/3} = 3.34 \times 10^{-3} \mu\text{m} = \underline{3 \text{ nm}}$.

Since atoms are widely spaced, the average length is probably more on the order of 0.3 nm.

33. A bicycle tire has a pressure of $7.00 \times 10^5 \text{ N/m}^2$ at a temperature of 18.0°C and contains 2.00 L of gas. What will its pressure be if you let out an amount of air that has a volume of 100 cm^3 at atmospheric pressure? Assume tire temperature and volume remain constant.

Solution First, we need to convert the temperature and volume:

$$T(\text{K}) = T(^{\circ}\text{C}) + 273.15 = 18.0 + 273.15 = 291.2 \text{ K, and}$$

$$V = 2.00 \text{ L} = 2.00 \times 10^{-3} \text{ m}^3.$$

Next, use the ideal gas law to determine the initial number of molecules in the tire:

$$P_1 V = N_1 k T \Rightarrow N_1 = \frac{P_1 V}{k T} = \frac{(7.00 \times 10^5 \text{ N/m}^2)(2.00 \times 10^{-3} \text{ m}^3)}{(1.38 \times 10^{-23} \text{ J/K})(291.15 \text{ K})} = 3.484 \times 10^{23}$$

Then, we need to determine how many molecules were removed from the tire:

$$PV = \Delta N k T \Rightarrow \Delta N = \frac{PV}{kT} = \frac{(1.013 \times 10^5 \text{ N/m}^2) \left(100 \text{ cm}^3 \times \frac{10^{-6} \text{ m}^3}{\text{cm}^3} \right)}{(1.38 \times 10^{-23} \text{ J/K})(291.15 \text{ K})} = 2.521 \times 10^{21}$$

We can now determine how many molecules remain after the gas is released:

$$N_2 = N_1 - \Delta N = 3.484 \times 10^{23} - 2.521 \times 10^{21} = 3.459 \times 10^{23}$$

Finally, the final pressure is:

$$\begin{aligned} P_2 &= \frac{N_2 k T}{V} = \frac{(3.459 \times 10^{23})(1.38 \times 10^{-23} \text{ J/K})(291.15 \text{ K})}{2.00 \times 10^{-3} \text{ m}^3} \\ &= 6.95 \times 10^5 \text{ N/m}^2 = \underline{6.95 \times 10^5 \text{ Pa}} \end{aligned}$$

38. (a) *In the deep space between galaxies, the density of atoms is as low as 10^6 atoms/m³, and the temperature is a frigid 2.7 K. What is the pressure? (b) What volume (in m³) is occupied by 1 mol of gas? (c) If this volume is a cube, what is the length of its sides in kilometers?*

Solution (a) Use the ideal gas law, where

$$\begin{aligned} PV &= NkT \\ P &= \frac{N}{V} kT = \frac{10^6}{1 \text{ m}^3} (1.38 \times 10^{-23} \text{ J/K})(2.7 \text{ K}) \\ &= 3.73 \times 10^{-17} \text{ N/m}^2 = 3.7 \times 10^{-17} \text{ N/m}^2 = \underline{3.7 \times 10^{-17} \text{ Pa}} \end{aligned}$$

(b) Now, using the pressure found in part (a), let $n = 1.00$ mol. Use the ideal gas law:

$$\begin{aligned} PV &= nRT \\ V &= \frac{nRT}{P} = \frac{(1.00 \text{ mol})(8.31 \text{ J/mol} \cdot \text{K})(2.7 \text{ K})}{3.73 \times 10^{-17} \text{ N/m}^2} = 6.02 \times 10^{17} \text{ m}^3 = \underline{6.0 \times 10^{17} \text{ m}^3} \end{aligned}$$

(c) Since the volume of a cube is its length cubed:

$$L = V^{1/3} = (6.02 \times 10^{17} \text{ m}^3)^{1/3} = 8.45 \times 10^5 \text{ m} = \underline{8.4 \times 10^2 \text{ km}}$$

13.4 KINETIC THEORY: ATOMIC AND MOLECULAR EXPLANATION OF PRESSURE AND TEMPERATURE

44. *Nuclear fusion, the energy source of the Sun, hydrogen bombs, and fusion reactors, occurs much more readily when the average kinetic energy of the atoms is high—that is, at high temperatures. Suppose you want the atoms in your fusion experiment to have average kinetic energies of 6.40×10^{-14} J. What temperature is needed?*

Solution Use the equation $\overline{KE} = \frac{3}{2}kT$ to find the temperature:

$$\overline{KE} = \frac{3}{2}kT \Rightarrow T = \frac{2\overline{KE}}{3k} = \frac{2(6.40 \times 10^{-14} \text{ J})}{3(1.38 \times 10^{-23} \text{ J/K})} = \underline{3.09 \times 10^9 \text{ K}}$$

13.6 HUMIDITY, EVAPORATION, AND BOILING

50. *(a) What is the vapor pressure of water at 20.0°C ? (b) What percentage of atmospheric pressure does this correspond to? (c) What percent of 20.0°C air is water vapor if it has 100% relative humidity? (The density of dry air at 20.0°C is 1.20 kg/m^3 .)*

Solution (a) Vapor Pressure for $\text{H}_2\text{O}(20^\circ\text{C}) = 2.33 \times 10^3 \text{ N/m}^2 = \underline{2.33 \times 10^3 \text{ Pa}}$

(b) Divide the vapor pressure by atmospheric pressure:

$$\frac{2.33 \times 10^3 \text{ N/m}^2}{1.01 \times 10^5 \text{ N/m}^2} \times 100\% = \underline{2.30\%}$$

(c) The density of water in this air is equal to the saturation vapor density of water at this temperature, taken from [Table 13.5](#). Dividing by the density of dry air, we can

get the percentage of water in the air: $\frac{1.72 \times 10^{-2} \text{ kg/m}^3}{1.20 \text{ kg/m}^3} \times 100\% = \underline{1.43\%}$

56. *What is the density of water vapor in g/m^3 on a hot dry day in the desert when the temperature is 40.0°C and the relative humidity is 6.00%?*

Solution

$$\begin{aligned}\text{percent relative humidity} &= \frac{\text{vapor density}}{\text{saturation vapor density}} \times 100\% \\ \text{vapor density} &= \frac{(\text{percent relative humidity})(\text{saturation vapor density})}{100\%} \\ &= \frac{(6.00\%)(51.1 \text{ g/m}^3)}{100\%} = \underline{3.07 \text{ g/m}^3}\end{aligned}$$

62. *Atmospheric pressure atop Mt. Everest is $3.30 \times 10^4 \text{ N/m}^2$. (a) What is the partial pressure of oxygen there if it is 20.9% of the air? (b) What percent oxygen should a mountain climber breathe so that its partial pressure is the same as at sea level, where atmospheric pressure is $1.01 \times 10^5 \text{ N/m}^2$? (c) One of the most severe problems for those climbing very high mountains is the extreme drying of breathing passages. Why does this drying occur?*

Solution (a) The partial pressure is the pressure a gas would create if it alone occupied the total volume, or the partial pressure is the percent the gas occupies times the total pressure:

$$\begin{aligned}\text{partial pressure (O}_2\text{)} &= (\% \text{O}_2)(\text{atmospheric pressure}) \\ &= (0.209)(3.30 \times 10^4 \text{ N/m}^2) = \underline{6.90 \times 10^3 \text{ Pa}}\end{aligned}$$

(b) First calculate the partial pressure at sea level:

$$\begin{aligned}\text{partial pressure (at sea level)} &= (\% \text{O}_2)(\text{atmospheric pressure}) \\ &= (0.209)(1.013 \times 10^5 \text{ N/m}^2) = \underline{2.117 \times 10^4 \text{ Pa}}\end{aligned}$$

Set this equal to the percent oxygen times the pressure at the top of Mt. Everest:

$$\text{partial pressure (at sea level)} = \left(\frac{\% \text{O}_2}{100\%} \right) (3.30 \times 10^4 \text{ N/m}^2) = 2.117 \times 10^4 \text{ Pa}$$

$$\text{Thus, } \% \text{O}_2 = \frac{2.117 \times 10^4 \text{ N/m}^2}{3.30 \times 10^4 \text{ N/m}^2} \times 100\% = \underline{64.2\%}$$

The mountain climber should breathe air containing 64.2% oxygen at the top of Mt. Everest to maintain the same partial pressure as at sea level. Clearly, the air

does not contain that much oxygen. This is why you feel lightheaded at high altitudes: You are partially oxygen deprived.

- (c) This drying process occurs because the partial pressure of water vapor at high altitudes is decreased substantially. The climbers breathe very dry air, which leads to a lot of moisture being lost due to evaporation. The breathing passages are therefore not getting the moisture they require from the air being breathed.

68. **Integrated Concepts** *If you want to cook in water at 150°C , you need a pressure cooker that can withstand the necessary pressure. (a) What pressure is required for the boiling point of water to be this high? (b) If the lid of the pressure cooker is a disk 25.0 cm in diameter, what force must it be able to withstand at this pressure?*

Solution (a) From Table 13.5, we can get the vapor pressure of water at 150°C : Vapor pressure = $\underline{4.76 \times 10^5 \text{ N/m}^2}$

- (b) Using the equation $P = \frac{F}{A}$, we can calculate the force exerted on the pressure cooker lid. Here, we need to use Newton's laws to balance forces. Assuming that we are cooking at sea level, the forces on the lid will stem from the internal pressure, found in part (a), the ambient atmospheric pressure, and the forces holding the lid shut. Thus we have a "balance of pressures":

$$P + 1 \text{ atm} - (4.76 \times 10^5 \text{ Pa}) = 0 \Rightarrow P = 3.75 \times 10^5 \text{ Pa}$$

$$\text{net } F = PA = (3.75 \times 10^5 \text{ Pa})(\pi(0.125 \text{ m})^2) = \underline{1.84 \times 10^4 \text{ N}}$$