Ideal Gases I – Answers

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1) Air is a solution of gases largely composed of N_2, O_2, Ar, CO_2, H_2O and other trace gases. Since we live surrounded by air, we live immersed in a gaseous solution.

9) (a) $822 \text{ torr} \times \frac{760.0 \text{ mm Hg}}{760.0 \text{ torr}} = 822 \text{ mm Hg}$

(b) $121.4 \text{ kPa} \times \frac{10^3 \text{ Pa}}{1 \text{ kPa}} \times \frac{1 \text{ atm}}{101{,}325 \text{ Pa}} \times \frac{760.0 \text{ mm Hg}}{1 \text{ atm}} = 910.57 \text{ mm Hg}$

(c) $1.14 \text{ atm} \times \frac{760.0 \text{ mm Hg}}{1 \text{ atm}} = 866.4 \text{ mm Hg}$

(d) $9.73 \text{ psi} \times \frac{1 \text{ atm}}{14.69 \text{ psi}} \times \frac{760.0 \text{ mm Hg}}{1 \text{ atm}} = 504.42 \text{ mm Hg}$

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46) Gases behave most ideally at high temperatures and ~~in~~ containers with large volumes and low pressures.

47) The ideal gas law states that $PV = nRT$.

At a constant temperature for a given amount of gas, $nRT = K$, where K is a constant.

Thus, $PV = K$, where K is a constant for a fixed T and amount of gas.

The above is a statement of Boyle's law relating the inversely proportional relationship between pressure and volume.

49) (a) $P = 782.4 \text{ mm Hg}$; $n = 0.1021 \text{ mol}$; $T = 26.2°C$; $V = ?$

$PV = nRT$

$$V = \frac{nRT}{P} = \frac{(0.1021 \text{ mol})(0.08206 \frac{\text{L atm}}{\text{mol K}})(26.2 + 273 \text{ K})}{\left(782.4 \text{ mm Hg} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}}\right)} = 2.44 \text{ L}$$

(b) $$P = \frac{nRT}{V} = \frac{(0.007812 \text{ mol})(0.08206 \frac{\text{L atm}}{\text{mol K}})(16.6 + 273 \text{ K})}{27.5 \times 10^{-3} \text{ L}} \times \frac{760 \text{ mm Hg}}{1 \text{ atm}} = 5{,}130.66 \text{ mm Hg}$$

(c) $$T = \frac{PV}{nR} = \frac{(1.045 \text{ atm})(45.2 \times 10^{-3} \text{ L})}{(0.002241 \text{ mol})(0.08206 \frac{\text{L atm}}{\text{mol K}})} = 256.85 \text{ K}$$

55) $P = 255$ atm ; $T = 25°C + 273 = 298$ K ; $V = 100.0$ L

$PV = nRT$

$$n = \frac{PV}{RT} = \frac{(255\text{ atm})(100\text{ L})}{(0.08206 \frac{\text{L atm}}{\text{mol K}})(298\text{ K})} = 1042.8 \text{ mol He}$$

$$\text{mol He} \times \frac{4.003\text{ g He}}{1\text{ mol He}} = 4174.3\text{ g He} = 4.17\text{ kg He}$$

$$m = \frac{PV}{RT} \times \mathcal{M} = \frac{(255\text{ atm})(100\text{ L})}{(0.08206 \frac{\text{L atm}}{\text{mol K}})(298\text{ K})} \times \frac{32\text{ g } O_2}{1\text{ mol } O_2} = 33369.6\text{ g } O_2 = 33.7\text{ kg } O_2$$

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77) We assume that the volume of the actual molecules have a neglible volume compared to volume of the gas, since the volume of the gas is equal to the volume of its container, and most containers are relatively so large that most of the space between gas molecules is "empty".

78) Collisions of the molecules in a sample of gas with the walls of the container are responsible for the gas's observed $\boxed{\text{pressure}}$.

79) Temperature is a measure of the average $\boxed{\text{speed}}$ of the molecules in a sample of gas.

80) The kinetic molecular theory of gases suggests that gas particles exert $\boxed{\text{neither}}$ attractive nor repulsive forces on each other.

81) The phenomenon of temperature is explained through the speed (average) at which molecules in the gas phase travel. The microscopic speed of molecules is reflected in their macroscopic temperature.

82) An increase in the temperature of a gas means an increase in the average speed of gas molecules. If confined to a rigid container, in which the volume does not change, the molecules moving at higher speeds will exert a greater force on the walls of the container, manifested in a greater pressure. This, a higher-temperature gas, with molecules moving at higher speeds, produces a greater pressure on the walls of a constant-volume container.