Chemistry Lecture #71: Ideal Gas Law

If you had a balloon filled with gas, how could you increase its volume? There are three ways to do this:

- Increase the temperature of the gas
- Decrease the atmospheric pressure surrounding the balloon
- Add more gas atoms/molecules (e.g., blow more air into the balloon). We measure the number of atoms/molecules of gas in moles.

Thus, the volume of the balloon is affected by the temperature and pressure of the gas, and the moles of gas in the balloon. There is a mathematical relationship between volume, temperature, pressure, and moles of gas. The relationship is called the ideal gas law.

$$PV = nRT$$

P = pressure of gas, usually in kPa
V = volume of gas in liters
n = moles of gas (units listed as mol)
T = temperature of gas, always in Kelvins
R = ideal gas constant = $\frac{8.31 \text{ L kPa}}{\text{mol K}}$

R, the ideal gas constant, is just a number we use to make the formula work correctly. 8.31 is used when the pressure units are in kPa. If the pressure is given in atm, then R = 0.0821. If the pressure is given in mm Hg or torr, then R = 62.4. I myself would just convert torr or atm into kPa and use R = 8.31 to avoid confusion. Remember that 760 torr = 101.325 kPa = 1 atm.

Find the volume of 2.40 moles of argon gas at a temperature of 30.0 °C and a pressure of 95.0 kPa.

V = ? n = 2.40 mol T = 30.0 °C + 273 = 303 K

P = 95.0 kPa $R = \dfrac{8.31 \text{ L kPa}}{\text{mol K}}$

Notice that we converted 30.0 °C into Kelvins.

PV = nRT

(95.0) V = (2.40)(8.31)(303)

$$V = \frac{(2.40)(8.31)(303)}{(95.0)}$$

V = 63.61 or 63.6 L

We can modify the formula $PV = nRT$ to find the volume if we are given grams of gas instead of moles.

For example, If I have 167.6 g of Kr gas, and 1 mole of Kr = 83.8 g, how many moles of Kr do I have?

Let m = mass of gas (g)
Let M = molar mass of gas g/mole

$$n = \frac{m}{M} = \frac{\text{mass of gas}}{\text{molar mass}} = \frac{167.6}{83.8} = 2.00 \text{ mol Kr}$$

If we substitute m/M in place of n in the ideal gas law, we get

$$PV = \frac{m}{M}RT \quad \text{or } PV = (m/M)RT$$

Neon gas is kept at a pressure and temperature of 97.2 kPa and 61.0 °C. Its mass is 0.750 g. What is the volume of the gas?

Solution

If we want to use $PV = (m/M)RT$, we need to know the molar mass of neon. Using the periodic chart, we see that neon has a molar mass of 20.2 g/mole.

We also need to remember to convert the Celsius temperature into Kelvins.

$P = 97.2$ kPa $\qquad\qquad$ $T = 61.0\ ^\circ C + 273 = 334\ K$

$V = ?$

$m = 0.750$ g

$M = 20.2$ g/mole

$R = \dfrac{8.31\ L\ kPa}{mol\ K}$

$$PV = \frac{m}{M}RT$$

$$\frac{97.2\ V}{1} = \frac{(0.750)(8.31)(334)}{20.2}$$

$$(97.2)\ V\ (20.2) = (1)(0.750)(8.31)(334)$$

$$V = \frac{(1)(0.750)(8.31)(334)}{(97.2)(20.2)}$$

$$V = 1.06\ L$$