

**Question:**

- A. What is the pressure of 1.50 mol Cl<sub>2</sub> in a 22.4 L container at 0.00 °C, assuming it is an ideal gas?
- B. If Cl<sub>2</sub> does not behave ideally, what will the pressure be assuming the van der Waals equation can be used to describe the state of the gas? The van der Waals parameters for Cl<sub>2</sub> are:  $a = 6.49 \text{ L}^2\text{-atm/mol}^2$ ,  $b = 0.0562 \text{ L/mol}$ .

**Concepts:** To solve this question we need to be able to understand and utilize both the ideal gas law and the van der Waals equation (for non-ideal gases).

**Connections:**

*What is given?* We are given the volume, number of moles and temperature of a sample of chlorine gas. We are also given the two van der Waals parameters for chlorine.

*What do I want to know?* We want to know the volume of the sample when it is behaving ideally and when it is not.

There are two equations of state for gases that we are utilizing: the ideal gas law and the van der Waals equation. These are described in Lessons 12-4 and 12-5. Most of the time we can assume that a gas behaves ideally and use the ideal gas equation to describe its state:

$$PV=nRT$$

However, this relationship fails when we are dealing with “real” gases. That is to say when a gas fails the basic assumptions of kinetic molecular theory and does not behave ideally. When dealing with real gases we must use the Van der Waals equation of state:

$$(P + an^2/V^2)(V - nb) = nRT$$

This equation compensates for the effects of intermolecular forces and the molecular volume of the gas molecules (parameters  $a$  and  $b$  in the equation).

Be sure you understand everything above before moving on to the solution below.



### Solution:

When a gas behaves ideally we can use the ideal gas law. Summarize the given information to determine if we need anything else and to determine which  $R$  value to use.

$$V = 22.4 \text{ L}$$

$$n = 1.50 \text{ moles}$$

$$T = 0.00^\circ\text{C}$$

Since we want to solve for pressure we use  $R = 0.0821 \text{ L-atm/K-mol}$

This value for  $R$  tells us we must convert  $T$  to Kelvin:

$$T = 0.00^\circ\text{C} + 273.15 = 273.15 \text{ K}$$

Now we use the ideal gas law, making sure to track our units:

$$P = nRT/V = (1.50 \text{ mol})(0.0821 \text{ L-atm/K-mol})(273\text{K})/22.4 \text{ L} = 1.50 \text{ atm}$$

We also need to calculate the pressure if the gas is not behaving ideally. We can summarize our information again:

$$a = 6.49 \text{ L}^2\text{-atm/mol}^2$$

$$b = 0.0562 \text{ L/mol}$$

$$V = 22.4 \text{ L}$$

$$n = 1.50 \text{ moles}$$

$$T = 273.15 \text{ K}$$

$$(P + an^2/V^2)(V-nb) = nRT$$

$$P + an^2/V^2 = nRT/(V-nb)$$

$$P = (nRT/(V-nb)) - an^2/V^2$$

$$P = (1.50 \text{ moles} \times 0.0821 \text{ L-atm/mol-K} \times 273.15 \text{ K}) / (22.4 - (1.50 \text{ moles} \times 0.0562 \text{ L/mol})) - ((6.49 \text{ L}^2\text{-atm/mol}^2 \times 1.50 \text{ mol} \times 1.50 \text{ mol}) / (22.4 \text{ L} \times 22.4 \text{ L}))$$

$$P = (33.64 \text{ L-atm}) / (22.3157 \text{ L}) - (0.029 \text{ atm}) = 1.48 \text{ atm}$$