Real Gases

> ## Skills to Develop
>
> - Describe how real gases differ from ideal gases
> - Derive the Van der Waals equation from the ideal gas equation

By the kinetic-molecular theory of gases, we imagine a gas as many small particles that bump off each other perfectly elastically (with conservation of momentum). They behave like hard little balls, and don't attract each other at all. Their kinetic energy depends on the temperature. In the derivation of the ideal gas law, we assume that there are no attractive forces between the particles and that the particles don't take up any space. These two assumptions are obviously incorrect: if there are no attractions between particles, there would be no liquids or solids. Also, the particles do take up a little space. Since we know that attractive forces become important at low temperatures, and that the volume of the particles will be important when the volume is relatively low (meaning pressure is high) we can predict that the ideal gas equation works best at high temperatures and low pressures.

If we want to make another equation that is closer to the real behavior of gases, we can make a few changes in the Ideal Gas Equation. First, we will assume that the particles have some volume. Instead of V we will use (V - nb) where n is the number of molecules or moles, and b is a constant for each different gas that means roughly how big it is.

Second, we need to include the effect of attractions between particles. If the particles attract each other, they will stay closer together and pump the walls a little less, so the observed pressure will be lower than we would expect. The higher the concentration of the gas (bigger n/V), the more important the attractive forces are. Actually the attractive forces depend on $(n/V)^2$, because this tells us how many other particles each particle can interact with. So we replace P with $(P + a(n/V)^2)$, where a is a constant that depends on the gas, and tells approximately how big the attractive forces are. This makes sense because bigger attractive forces happen when the particles are very close together (V is small), and they cause the pressure to seem lower.

Putting this together, we have the Van der Waals equation:

$$\left(P + a \left(\frac{n}{V} \right)^2 \right) (V - nb) = nRT \tag{1}$$

This equation describes real gases pretty well, although there are other equations used also. The constants a and b are found by fitting the real data for each gas to this equation. You can look up values of a and b in tables.

Outside Link

- CrashCourse Chemistry: Real Gases (12 min)

Contributors and Attributions

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