Real Gases, the Van der Waals Equation

The ideal gas law is accurate when:

- The volume of gas particles is small compared to the space between them
- The attractive forces between gas molecules are not significant

These assumptions are valid around STP, but they are not valid:

- **when the pressure is much higher than 1 atm.**

  With higher pressure (several hundred atm), the molecules themselves take up a significant amount of the sample's volume. Ideal gas law predicts the molecules have **no** volume, so the actual volume will be: higher than prediction

  $PV = nRT$

  To correct for this, *V* becomes: $V - nb$

  b = constant = molecular volume

- **when the temperature is much lower than 298 K.**

  With low temperatures, the molecules begin to stick together more, reducing the number of collisions with the walls of the container. The actual pressure will be: lower than prediction < 50 K maybe?

  To correct for this, *P* becomes: $P + \frac{n^2 a}{V^2}$

  a = constant for intermolecular forces

Van der Waals Equation:

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