

The ideal gas law

In this tutorial:

- Recap: What's an ideal gas?
 - Discussion of the ideal gas law.
 - Learn how to use the gas law, with plentiful examples.
 - Nonideal gases and the Van der Waals equation.
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Introduction

Now that you've mastered the world of Boyle's law, Charles's law, Gay-Lussac's law, and the combined gas law, it's time to move to the ideal gas law. If you're scared of math (and i know that some of you are), don't be – the math is pretty simple, and I'll walk you through it anyway. Enough talking. Let's get started.

Recap: What's an ideal gas?

Whenever we do calculations using gases, we assume that we're working with something called an ideal gas. Ideal gases are gases that conform to the postulates of the kinetic molecular theory, among them:

- Gas molecules are infinitely small.
- Gas molecules don't experience intermolecular forces.
- Gas molecules move in random directions unless they hit something.
- The kinetic energy of gas molecules is proportional to their temperature in Kelvin.

To the layman, what this all means is that gas molecules behave like troubled loners, in that they don't really interact with any other gas molecules – as a result, we can treat all of them as if they were alone in the world.



Close up view of an ideal gas molecule.

For the whole tutorial about the kinetic molecular theory and ideal gases, click [here](#).

The ideal gas law:

Unlike the [other gas laws](#) we talked about, the ideal gas law doesn't describe what happens to a gas when you manipulate it (i.e. when you change the pressure, volume, temperature). Instead, the ideal gas law describes how a gas will behave under some unchanging set of conditions referred to as an *equation of state*.



*Another equation of state:
Florida = home of best “Cops” episodes*

The ideal gas law looks like this:

$$PV = nRT$$

The terms in this equation should be mostly familiar to you if you've already learned the combined gas law (and the other ones like it). However, if it's not, let's review:

- P = the pressure of the gas. In ideal gas equations, this is typically given either in atmospheres or kilopascals.
- V = the volume of the gas. Always use liters or cubic decimeters (dm³) – they're the same thing.
- n = number of moles of gas. Yep, you've got to go back and remember how to do mole calculations if you've forgotten. Fortunately, you can do that [here](#).
- R = the ideal gas constant. There are two values for this: 8.314 L·kPa/mol·K (if the pressure is given in kPa) or 0.08206 L·atm/mol·K (if pressure is given in atm).
- T = the temperature in Kelvin (K). Don't use degrees Celsius – remember that the temperature in Kelvin is equal to the temperature in degrees Celsius + 273.

So, let's see what this actually means. Let's say that we want to understand how the other variables affect the pressure of a gas: If you have a lot of gas at a high temperature, the pressure will be high. If the gas is crammed into a small volume, the pressure will be even higher than if it's in a large container.

I'll let you figure out the others for yourself, but you get the idea: This equation isn't a weird mathematical mutant – it's just a relationship that tells you how gases seem to behave.¹



It seems that pressure (left) and volume (right) enjoy a close relationship.

Using the ideal gas law

Instead of doing a lot of writing, let's just get into an example:

Question: If I have 3.9 moles of methane gas at a pressure of 1.7 atm and a temperature of 25 degrees Celsius, what will the volume of this gas be?

Answer: To solve this question, we need to figure out what numbers go into the equation:

- P: 1.7 atm (the problem gave that to us)
- V: X (we're solving for this)
- n: 3.9 moles (the "methane" part doesn't make any difference in this problem, as all gases behave the same way.)
- R: 0.08206 L·atm/mol·K (we use this value and not the other because this problem gives pressure in "atm" and this constant has "atm" in the units. The other constant has "kPa" in the units, so it doesn't work out.)
- T: 298 K (remember to convert 25 degrees Celsius to Kelvin by adding 273!)

Plugging this into the equation:

$$PV = nRT$$

$$(1.7 \text{ atm})(X) = (3.9 \text{ mol})(0.08206 \text{ L} \cdot \text{atm} / \text{mol} \cdot \text{K})(298 \text{ K})$$

$$X = 56 \text{ L}$$

It's that easy!

Another example: If I had 74 grams of methane gas in a container that has a volume of 50.0 L and a pressure of 7500 kPa, what is the temperature of the gas?

Answer: Again, find the variables:

- P = 7500 kPa
- V = 50.0 L
- n = 4.6 moles (remember, if you're given grams, you have to convert to moles to put it into this equation. In this example, you divide 74 grams by the molar mass of methane – 16.0 g/mol – to get the answer. For more mole help, visit [this tutorial.](#))

- $R = 8.314 \text{ L}\cdot\text{kPa/mol}\cdot\text{K}$ (since the pressure is given to you in kPa in the problem, you have to use the constant that uses kPa to solve this question)
- $T = X$ (we're solving for this)

Let's plug them into the equation:

$$PV = nRT$$

$$(7500 \text{ kPa})(50.0 \text{ L}) = (4.6 \text{ mol})(8.314 \text{ L}\cdot\text{kPa/mol}\cdot\text{K}) X$$

$$X = 9,800 \text{ K}$$

You can leave the answer in Kelvin, or subtract 273 to find that it's 9,500 degrees Celsius (yep, I'm using significant figures!) Either way, it's hot.

When gases aren't ideal: The Van der Waals equation

On a number of occasions, I've said that all gases are the same. However, you've probably realized with almost no effort that this isn't true – all gases *are* different. It's not a *big* difference (maybe a 5-10% deviation from ideal), but sometimes it's handy to squeeze that extra precision out of our calculations in order to make things even more awesome. When we want to do that, we break out the Van der Waals equation:

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

This is, as you can imagine, not a fun equation to work with. Though if you squint really hard, you can see that this equation does seem to take the form $PV = nRT$, which means that it's at least based on the ideal gas law.

Let's talk about the differences:

- The term “a” compensates for the attractive force between the gas molecules. As seen in this equation, the stronger the attractive force, the lower the pressure will be. Note again that this is usually pretty small, so the term involving “a” is usually ignored. This term, incidentally, differs depending on which gas you're working with – not surprisingly, helium (which has almost nothing in the way of intermolecular forces) has an “a” value that's nearly zero.
- The term “b” compensates for the volume occupied by the gas molecules themselves. In this equation, the V term we're used to is replaced with a term that only refers to the volume between the particles and not the volume of the gas particles themselves. Again, this term is usually small, so we ignore it. Again, the values for “b” differ for every gas, and are smallest for very small gas molecules.

And, if we ignore both of the terms with a and b, we can see that we end up with $PV = nRT$.

So, when will you need the Van der Waals equation? Probably never, because it's pretty obscure. In fact, it's not even the best equation out there for figuring out what happens to real gases: If you want that, take a look at the [Redlich-Kwong equation](#).

Footnotes:

1. Actually, all equations are just relationships between different things. Of course, you may have noticed that in the world of math this is harder to figure out because you're not talking about any actual things that exist in the universe. However, remember that the job of mathematicians is to figure out ways of doing computations so that those of us who do science can better understand how different things are related to each other.
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