

The kinetic molecular theory of gases

We have a gas problem around my house. You see, something I did recently caused the house to stink of gas in a most terrible way. My family was upset, my cats were upset, and even I was disturbed by the noxious smells filling my home. Fortunately, the gas company was able to come and figure out that I'd accidentally left the stove burner on "low" after cooking dinner and when it blew out the natural gas in the pipes started to fill the house. Man, that was a load off of my mind!



Wait, what did you think I was talking about?

In any case, for the purposes of this post, I'll pretend that this got me thinking about the gas laws we all know and love in chemistry. Because that makes for a better segue.

What's a gas?

This the question that textbooks often ask, even though we already know the answer. Seriously, do those books think that we're so stupid that we're going to go out and inhale a glass of water if they don't give us a definition?



Pictured: Warning label for people with very stupid children.

So, how do gases behave? Intuitively, you know that all gases behave in different ways – and this is true. However, it turns out that gases are similar enough to one another that we can simplify our lives by treating them as if they were the same. This hypothetical gas that we use as a reference for

all other gases is called an *ideal gas*, and pretending that it's real turns out to be a surprisingly good assumption.

So, if we're going to pretend that we have a wondrous ideal gas, we need to figure out how it behaves. And to figure out how it behaves, we need to know more about what the molecules are up to. And fortunately, we've got a great way of doing this:

The kinetic molecular theory of gases

The basic idea of the kinetic molecular theory of gases is this:

The way that gas molecules behave on a molecular scale can tell us how they behave on an observable scale.

Since, all gases have slightly different properties, we'll use the properties of our hypothetical ideal gas to figure out how gas particles behave. Let's take a look at the postulates (i.e. assumptions) that the kinetic molecular theory of gases uses:

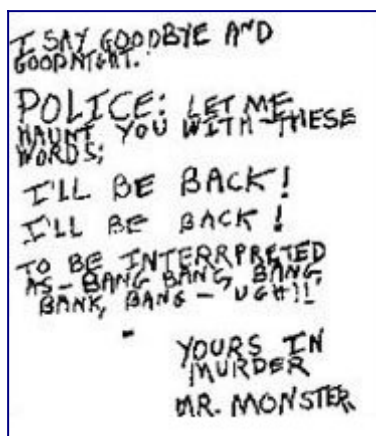
Postulate 1: The particles of an ideal gas are infinitely small.

You may think this is ridiculous because gas molecules are *not* infinitely small (some molecules such as sulfur hexafluoride are quite large), but when you compare the volume of a gas molecule to the overall volume occupied by the gas, it turns out that even big gas particles are so small that their volume can be ignored.¹

Postulate 2: The particles of an ideal gas experience no intermolecular forces.

The idea behind this postulate hinges on the fact that gas molecules move really really really fast. Like, really super fast. In fact, they move so quickly past one another that their intermolecular forces don't really have time to affect the gases behavior.

Think of it like this: We all knew this kid in our elementary school class who absolutely stank like garbage. When you walked by this kid, you found that the smell was so bad that you'd get that feeling you get when you're going to puke but don't quite puke. As a result, everybody learned to run past that kid as fast as they could, because the effect of this stench wouldn't really bother you if you moved past him quickly enough. Intermolecular forces are like this: If they don't have time to exert their influence, they don't really affect anything.



Whereas the stinky kid tends to have issues later in life.

Postulate 3: The particles of an ideal gas are in random, constant motion, and only change direction when they hit something else.

This is something so obvious that you don't even think about it. Let's consider each point individually:

Ideal gas particles are in random motion, which means that they don't preferentially go in one direction or the other. As a result, you don't really have to worry about all of the air molecules in the room travelling to one side, leaving you in a vacuum in the other half of the room.



Wind is case where this isn't true, as the gas molecules preferentially travel in the direction the wind blows. Even in a stiff wind, however, most of the gas molecules are acting pretty randomly.

As for the collision part of this postulate, it's the force of the gas particles hitting things that is responsible for gas pressure. As more gas molecules hit an object, the pressure increases.

Postulate 4: The kinetic energy of particles in an ideal gas is directly proportional to their temperature in Kelvin.

To put it simply, if you heat up a gas, the particles will gain kinetic energy. You've probably heard that kinetic energy is a measure of the energy something has because of its motion – this essentially means that gas molecules will move more quickly as the temperature increases.

Interestingly, in an ideal gas, every type of gas has the same average kinetic energy at the same temperature. As a result, very light helium molecules will have the same average kinetic energy as much heavier carbon dioxide molecules. However, because helium molecules are so much lighter, they have to travel much faster than carbon dioxide molecules at the same temperature to achieve this energy. In the same way, a .22 caliber rifle round travelling at 720 mph has the same kinetic energy as a 16 pound bowling ball travelling at about 13 mph.



And, by extension, this child might as well be playing with a loaded gun.

OK... now what?

It turns out that understanding the postulates of the kinetic molecular theory of gases really does a nice job of explaining the behavior of gases. Let's see:

- Gases have no fixed volume, shape, or size: Because gas molecules experience no intermolecular forces, there's nothing holding them together. As a result, they're free to move as far apart as they want and fill the nooks and crannies of whatever container you put them in. Solids and liquids, on the other hand, have strong forces holding the particles together, which means that the particles have much less freedom to do these things.
- Gases are less dense than liquids or solids: Liquids and solids experience strong forces that keep the particles crammed together. In gases, there are no such forces, which means that they tend to have a lot of space between their molecules. This empty space equals low density.
- Gases mix well: Essentially, the kinetic molecular theory says that all gas molecules pretty much ignore the other gas molecules that are around it. As a result, if you add one gas to another, the gas you add will mix freely with the other.



A point that Uncle Lou makes every chili night.

And the most important thing about the kinetic molecular theory is...

When we make the assumptions above, we can use the tools of statistics to figure out how gases behave on a large scale. Yes, when you start learning about the gas laws in the next amazing tutorial, you'll find that all of the equations that arise can be traced back to the assumptions of the kinetic molecular theory. This use of statistics to study collections of things is called [statistical mechanics](#), and it's one of the most useful tools that chemists and physicists use to figure out how various things (not just atoms or molecules) behave.

Footnotes:

1. For real gases, this volume does have a very small effect on the behavior of that gas. Generally speaking, small molecules behave in a more ideal fashion than large ones. Furthermore, gases at low pressure (i.e. they're not crammed together very tightly) behave in a more ideal fashion than gases at high pressure. The good news is that, except in extreme cases, this isn't really something that makes much effect on the properties of the gas.
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