

Intermolecular forces

Now that you're a professional polarity-learning person, it's time to become a pro at the whole intermolecular forces thing. As I'm sure your teacher has told you, intermolecular forces are as exciting as a monkey knife fight and as useful as pants. So, without further ado...



Now, if I only had a switchblade and a pair of Dockers...

Forces between stuff

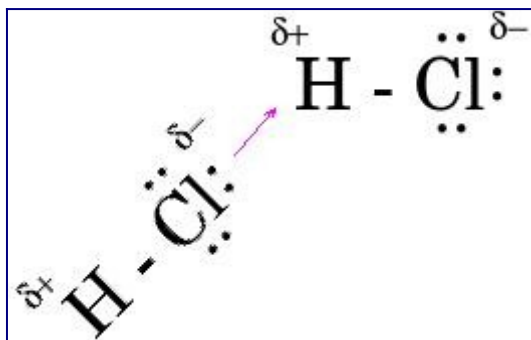
In chemistry we talk about two main types of force that can hold different things together. These include the following:

- **Intramolecular forces:** “Intra” means “within”, so these forces hold the things within molecules together. Or, to put it more clearly, these are chemical bonds. Though the electrostatic interactions in ionic bonds aren't real bonds (no shared electrons and whatnot), these are also counted as intramolecular forces. Intramolecular forces are really really strong, because they need to be strong to hold the various atoms and ions in chemical compounds together.
- **Intermolecular forces:** These are forces between two different covalent molecules. In other words, if you have one water molecules kind of hanging around somewhere, other water molecules will be attracted to it via various intermolecular forces. Intermolecular forces have varying strengths, but are generally nowhere near as strong as intramolecular forces. As a result, this force cannot be thought of as locking things in place, but rather keeping things sort of loosely attached.¹

The types of intermolecular forces

There are three big types of intermolecular force that you'll need to know. These are:

Dipole-dipole forces: As you learned in the tutorial about polarity, one side of polar molecules has partially-positive charge while the other side has partially-negative charge. As a result, if you put two polar molecules near each other, the partial positive side of one will be attracted to the partial negative side of the other:



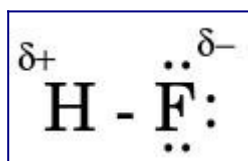
The partial positive charge on the Cl on the HCl on the left is attracted to the partial positive charge of the H on the HCl on the right.

Because + and – attract, this force helps to keep the HCl molecules stuck near each other.

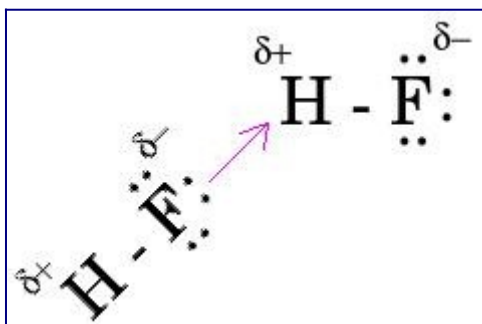
Hydrogen bonds: Very special, very strong intermolecular forces that take place when a hydrogen atom is stuck to fluorine, nitrogen, or oxygen.

Here's the deal: Fluorine, nitrogen, and oxygen are all highly electronegative and very small. As a result, they contain a very concentrated negative charge. This charge is a *lot* more concentrated than in other atoms. As a result, when you have molecules containing these atoms, these atoms kind of go nuts and pull electrons away from whatever they're bonded to.

Now, consider hydrogen. Hydrogen isn't particularly interested in either gaining or losing electrons, so when you have F, N, or O nearby, most of the shared electrons in the covalent bonds gets pulled away from hydrogen, giving hydrogen a very strong partial positive charge and the other side a very strong partial negative charge:



Let's see what happens when one HF molecules comes near another:



You may be asking yourself how this is any different from the HCl example above. And that's a good question, because they're pretty much the same thing. However, there *are* differences, and they're what make hydrogen bonding so special:

- HF (or HO and HN bonds) are very polar. This means that there's a lot of partial positive and negative charge, which leads to a very strong dipole-dipole force.
- F and H are both very small atoms (as are N and O). As a result, they can get very close to one another, causing the already strong interaction between them to get even stronger. This strength isn't quite as strong as it is in a bond, but it's strong enough that people used to think that it actually was a weak bond. But it's not. It's just called a hydrogen bond, but it's not a real bond.

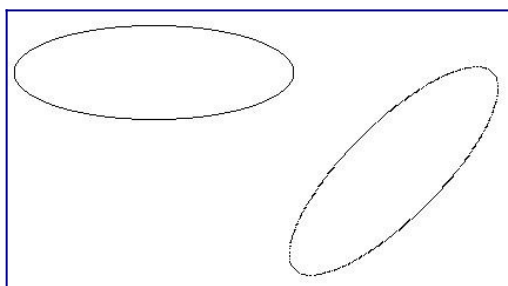
London dispersion forces (sometimes erroneously called Van der Waals forces): These are forces that take place when a partial charge on one nonpolar molecule is attracted to the partial charge on another nonpolar molecule. If you find that confusing, then you're on the right track to understanding this stuff.



Dr. London was so well-dressed that cameras couldn't properly focus on him.

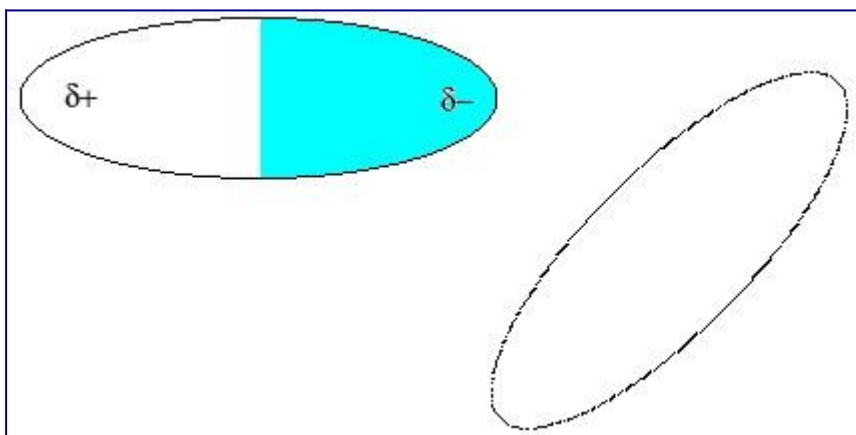
The reason you might find this confusing is that nonpolar molecules don't have partial charges.² Saying that one partial charge on a nonpolar charge does anything at all sounds like nonsense, because it doesn't even exist. Just as you don't have to worry about the horns on your cat damaging your furniture³, it would seem weird to worry about the partial charge on a nonpolar molecule.

Let's see how this works:



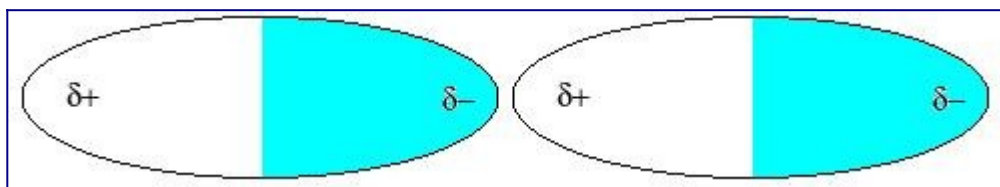
As our story starts, two nonpolar molecules are just chilling. They aren't paying any attention to each other at all.

But then, something crazy happens:



And then for no particular reason at all, random movements of electrons within the molecule on the left cause the electrons to shift, making the side with more electrons partially-negative and the side without them partially-positive.

In turn:



This imbalance in electrons causes the electrons in the second molecule to compensate by shifting its electrons, too. As a result, the electrons nearest the partial negative side of the first molecule move away (they have the same charges as the electrons on the other molecule), making the second molecule polar, too. This causes the molecules to attract in the same method as molecules experiencing dipole-dipole forces.

However, this force doesn't last for long. Remember, it was the random movement of electrons that caused this charge imbalance in the first place, so it's the random movement of electrons that eventually undoes it and changes everything back to the initial state. Furthermore, because nonpolar molecules can't be made very polar using this method, London dispersion forces are very weak.

How can you tell which force is the important one for some particular molecule?

Draw the Lewis structure!

- If the molecule is nonpolar, then the primary intermolecular force is a London dispersion force.
 - If the molecule has an H-F, H-O, or H-N bond anywhere on it, the primary intermolecular force is hydrogen bonding.
 - If the molecule is polar but doesn't have the bonds mentioned above, its primary intermolecular force is a dipole-dipole force.
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Strengths of intermolecular forces

As mentioned above, intermolecular forces are much weaker than chemical bonds. However, when you compare intermolecular forces to each other, you can see based on what I mentioned above that London dispersion forces are the weakest, followed by dipole-dipole forces, and then hydrogen bonds are the strongest.

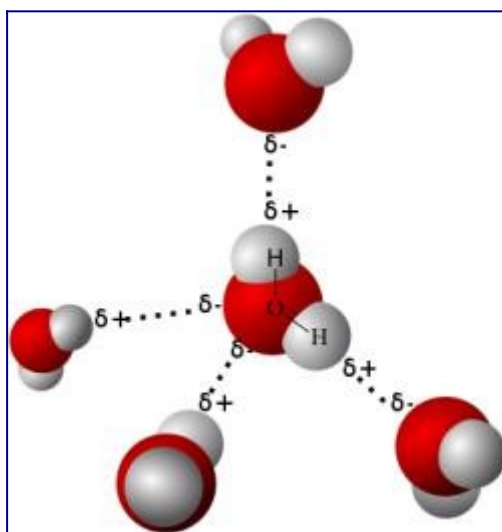
That said, not all dipole-dipole forces have identical strengths with one another, nor is this the case in any other molecule. A closer look is needed.

London dispersion forces: These are the weakest, but are stronger in large molecules than they are in small molecules. The reason is simple: If the random movement of electrons causes London dispersion forces to occur, then larger molecules (which have more electrons) will have stronger London dispersion forces than smaller ones (which have fewer). Examples:

- Boiling point of methane (CH_4): -161 degrees Celsius
- Boiling point of octane ($\text{CH}_3(\text{CH}_2)_6\text{CH}_3$): 126 degrees Celsius

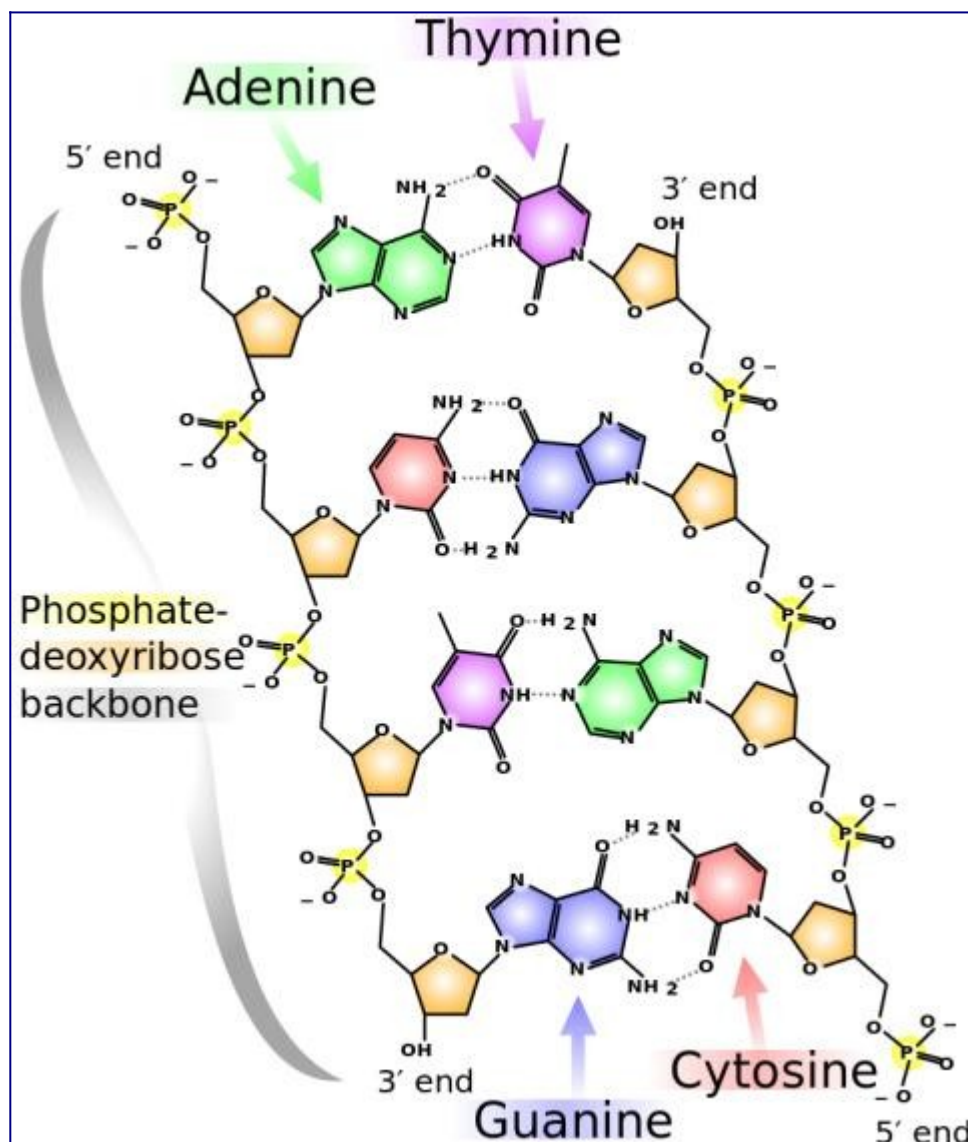
Dipole-dipole forces: Stronger than London dispersion forces, these are weaker than hydrogen bonds. Dipole-dipole forces tend to be strongest in highly polar molecules, which causes similar effects on boiling points. There are lots of examples of this, many of which you can see in [this table](#).

Hydrogen bonds: These are really strong intermolecular forces, and are strongest when more than one occurs. Some very impressive examples of this include water:



All of the atoms in water undergo hydrogen bonding, making the forces between these molecules extremely strong.

and DNA:



Those little dashes are hydrogen bonds.

As you probably learned in biology class, nucleic acids like DNA and RNA can “unzip”⁴, do some coding stuff, and then “rezip.” They can do this because the two strands are connected via big strings of hydrogen bonds, which are strong enough to keep the two strands together under normal circumstances but not strong enough so they can’t unzip for use in cells.

Footnotes:

1. Well, they do hold things in place in solids, but that’s because there’s not much energy around to pull the solid molecules apart. In that case, it’s not so much that the intermolecular forces are strong but that there’s so little energy around to overcome it.
2. Other reasons: You weren’t paying that much attention, you don’t care, you’re surfing the web or updating Facebook, or your cat just mangled your leg.
3. If you get into genetic engineering, I totally want a cat with horns.
4. I was going to put a silly photo here, but it turns out that if you search stock photo databases for “unzip” you end up seeing terrible, terrible things.

Image credits:

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