Intro to Phases and Intermolecular Forces

Skills to Develop

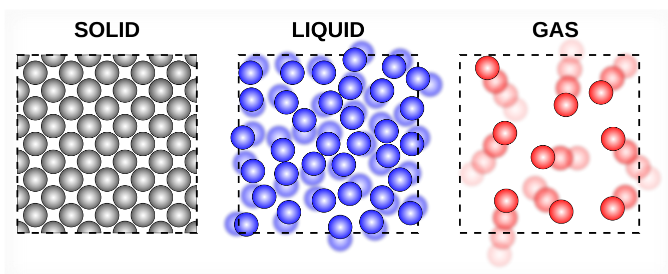
- Describe phases and phase changes on a molecular level

Phases of matter mean the state of a material, like solid, liquid or gas. How can we predict whether a material will be a solid, liquid or gas under certain conditions? We have to know about the forces that hold the material together.

In solids, the separate molecules or ions are held tightly in their positions by some type of force. They might vibrate a little bit in place, but they don't move around. For this reason, solids don't easily change shape. You might be able to bend or break them, but usually they don't change shape by themselves.

In liquids, the separate molecules or ions are held close together by some type of force, but they can move around while staying close together. For this reason, a liquid can change shape to fit whatever is holding it. But like solids, liquids have approximately constant volume. Even if the liquid flows into a new shape, the distance between the molecules doesn't change, so the total volume stays the same.

In a gas, usually the molecules bounce around as though there are no forces between them. (See Kinetic-Molecular Theory of Gases.) At very high pressure, when they are forced to be close together, we might start to notice that there are some forces between the molecules (because the pressure is less than we expect from the Ideal Gas Law) but usually they move around separately. Because they aren't really attracted to each other and have a lot of kinetic energy (at least at normal temperatures), they fill the whole space they have. So gases can take any shape, and also can change volume a lot.



An illustration of the 3 phases of matter.

The higher the temperature, the more kinetic energy the molecules or ions have. With more kinetic energy, it's harder for them to stay in their place in a solid, or not to bounce right out of a liquid and become a gas. So as we increase the temperature, we might see **phase transitions** from solid to liquid to gas. We can think about these transitions using equilibrium, like when we think about reaction and solubility equilibria. Think about a liquid in a closed container, like a bottle half full of water. Some of the molecules have bigger kinetic energy and some have smaller kinetic energy. If a molecule with big kinetic energy is on the surface of the liquid, it might fly off and enter the gas state. At the same time, other molecules in the gas state might bump the surface of the liquid, and if they don't have very much kinetic energy, they might stay there and join the liquid. This is a dynamic equilibrium: the molecules go back and forth between the 2 states. If we increase the temperature, the average kinetic energy increases, and that means the molecules are more likely to have enough kinetic energy to go into or stay in the gas phase. Liquid molecules will become gas molecules more often, and gas molecules will become liquid molecules less often. Then the equilibrium will move, so a bigger % of the total molecules are gas.

In many cases, there is a specific temperature above which all of a material goes from solid to liquid (a melting point) or from liquid to gas (a boiling point). What temperature that is depends on the strength and type of forces between the molecules and ions. If the forces are strong, then more kinetic energy is needed to make the molecules move around or separate and become gas, which means the melt point and boiling point are higher. If the forces between molecules are very weak, then the material may be a gas, and it may be hard to cool it enough to make a liquid.

A graph showing phase changes with respect to changes in temperature.

In the next sections, we will talk about the forces between molecules that determine boiling points and melting points and other important properties. (Thus, they are called **intermolecular forces**, to separate them from the forces inside molecules that hold the molecules together.) These forces are often called Van der Waals forces after Johannes van der Waals, who wrote the equation for real gases. Van der Waals figured out that the reason gas pressures are often lower than we expect (at high pressure) is because there are attractions between the molecules. Actually Van der Waals was the son of a carpenter who wasn't allowed to enter university because he didn't have the right expensive primary education. But he took classes anyway, became a teacher, and eventually they changed the rules for university admission, so he got a doctorate, became famous, became a professor, and won the Nobel Prize.

Outside Links

- Khan Academy: States of Matter (19 min)
- CrashCourse Chemistry: Liquids (11 min)

Contributors and Attributions

- Emily V Eames (City College of San Francisco)
