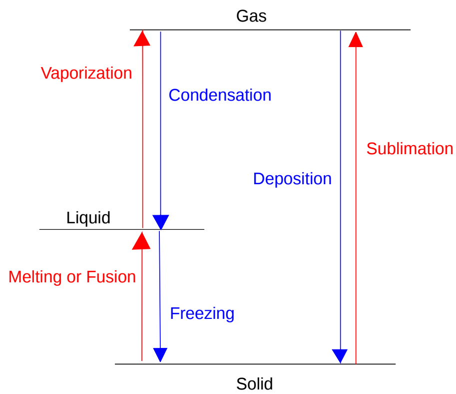
Phase Changes

Skills to Develop

- Describe the relationship between heat (energy), bonding forces, and phase changes

Most phase changes occur at specific temperature-pressure combinations. For instance, at atmospheric pressure, water melts at 0 °C and boils at 100 °C. In this section, we will talk about when and how they happen. The names of the different phase changes are shown below:



Important terms describing phase changes

Predicting Phase-Change Temperatures

We can predict the relative temperature at which phase changes will happen using intermolecular forces. If the intermolecular forces are strong, then the melting point and boiling point will be high. If the intermolecular forces are weak, the melting and boiling point will be low.

London forces vary widely in strength based on the number of electrons present. The number of electrons is related to the molecular or atomic weight. Heavy elements or molecules, like iodine or wax, are solids at room temperature because they have relatively strong London forces, which correlate with big molecular weights. London forces are always present, but in small molecules or atoms, like helium, they are quite weak.

Dipole-dipole forces are present in molecules with a permanent dipole. We can predict this by drawing a Lewis structure, identifying polar bonds using electronegativity, predicting the shape of the molecule, and seeing if the bond dipoles on different molecules can touch. If they can, there will be dipole-dipole forces. The bigger the dipoles (bigger electronegativity difference, etc.) and the closer together they can get, the bigger the dipole-dipole forces are.

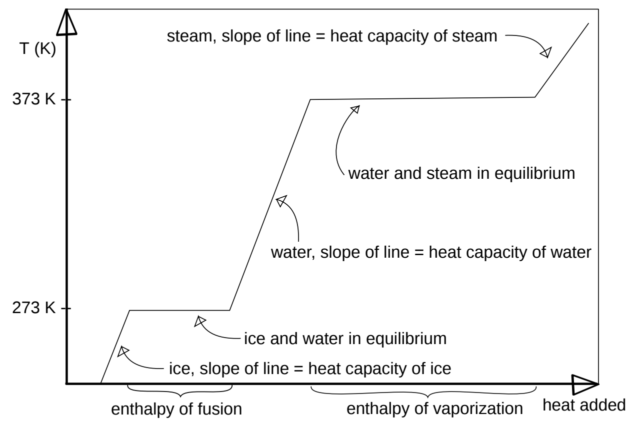
Hydrogen bonds occur only when there are H atoms bonded to N, O, or F and lone pairs on N, O, or F. Look for both of these in molecules to see if they can hydrogen bond.

If you can find which types of intermolecular forces are present in a molecule, you can make some guesses about which molecules have higher or lower melting or boiling points. For instance, let's compare methane (CH_4), silane (SiH_4), hydrogen sulfide (H_2S) and water (H_2O). Methane and silane are non-polar, because of the tetrahedral shape and also the small electronegativity differences. Because these don't have dipole-dipole forces, the boiling point will depend on how strong the London forces are. Silane is heavier, so it has bigger London forces and a higher boiling point. Between water and hydrogen sulfide, both are polar, and have dipole-dipole forces, so they have higher boiling points than methane or silane. But water has hydrogen bonds, which are extra-strong dipole-dipole forces. Water boils much hotter than hydrogen sulfide.

Energy and Phase Changes

We can't really explain phase changes in terms of energy without entropy, which we haven't talked about yet. For now, we can just say that as we add energy to a substance, it usually gets hotter and the particles have more kinetic energy. This will make it easier for them go from solid to liquid, or liquid to gas. Gases have more energy than liquids, which have more energy than solids. As we increase the temperature, the stable form of the substance goes from solid to liquid to gas. The transition temperatures (melting point, boiling point) are the temperatures at which both phases are stable and in equilibrium. Actually, there will be some gas in equilibrium with solid and liquid all the time, because a few molecules can always escape the solid/liquid, but the solid or liquid won't be present above certain temperatures.

For instance, imagine heating a solid. The molecules start moving more, and the temperature increases as predicted by the heat capacity. At some point, they have so much energy that it's hard for them to stay in the orderly solid, so the solid starts to melt. As we add more heat, the temperature doesn't change, because all the heat we add goes into melting the solid. The solid can't get any hotter than it is, and the liquid can't increase it's temperature because its kinetic energy is absorbed to melt the remaining solid. The amount of energy needed to melt the solid is the **enthalpy of fusion**. When all the solid is melted, if we keep adding heat, the temperature will rise again. As the temperature rises, the vapor pressure increases, because more molecules have enough kinetic energy to escape. Still, most of the molecules are in the liquid form, because the total pressure pushes on the liquid and keeps it from expanding into a gas. When the temperature increases to the boiling point, then the vapor pressure will be equal to the outside pressure. Now, because the vapor pressure is equal to the atmospheric pressure, bubbles form in the liquid. It can expand into a gas, because it's pressure is the same as the atmospheric pressure. The temperature will stay constant again as all the liquid become gas, while you add the **enthalpy of vaporization**. Then if you keep heating the temperature of the gas will increase. This is shown in the diagram below:



Heating diagram for water, showing change in temperature and heat is added.

Outside Links

- Khan Academy: States of Matter (19 min)
- Khan Academy: Specific Heat, Heat of Fusion and Vaporization (15 min)

Contributors and Attributions

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