

2.2: The States of Matter

As described in Section 2.1, a molecule of water is composed of two atoms of hydrogen bonded to one atom of oxygen (H_2O). All water molecules are exactly the same (same ratio of elements, same geometric bonding pattern), but we encounter water in three different forms in the world around us. At low temperature, water exists as a solid (ice). As the temperature increases, water exists as a liquid, and at high temperature, as water vapor, a gas. These three forms of water represent the three **states of matter**: solids, liquids and gases. States of matter are examples of **physical properties** of a substance. Other physical properties include appearance (shiny, dull, smooth, rough), odor, electrical conductivity, thermal conductivity, hardness and density, to name just a few. We will discuss density in more detail in the next section, but first let's examine the states of matter and how they differ on an atomic level.

If ice, liquid water and water vapor all consist of identical molecules, then what accounts for the difference in their properties? So far, we have talked about molecules as if they were standing still, but in fact, they are always moving. In chemistry, we often explain the states of matter in terms of the **kinetic molecular theory (KMT)**. The word *kinetic* refers to motion and the kinetic molecular theory suggests that atoms and molecules are *always* in motion. The energy associated with this motion is termed *kinetic energy*. The amount of kinetic energy that a particle has is a direct function of temperature, and it is the kinetic energy of the water molecules under different conditions that determines the different properties of the three states of water.

Atoms and molecules move in different ways under different conditions because of the forces attracting them to each other, called **intermolecular forces**. *Intermolecular forces* is a general term describing the fact that all atoms, and molecules share a certain inherent *attraction* for each other. These attractive forces are much weaker than the bonds that hold molecules together, but in a large cluster of atoms or molecules the sum of all of these attractive forces can be quite significant.

Now, consider a group of molecules or atoms clustered together and held in place by these attractive forces. At low temperature, the molecules or atoms will remain stuck together in a *lump* of defined shape and structure, like water in the form of an ice cube. This is referred to as the **solid phase**. At the atomic level, the molecules or atoms in a solid are closely packed, and although they are still all rapidly moving, their movements are so small that they can be thought of as vibrating about a fixed position. As an analogy here, think of a handful of small magnets stuck together in a solid mass. Solids and liquids are the most tightly packed states of matter. Because of the intermolecular forces, solids have a defined shape, which is independent of the container in which they are placed. As energy is added to the system, usually in the form of heat, the individual molecules or atoms acquire enough energy to overcome some of the attractive intermolecular forces between them so that neighboring particles are free to move past or slide over one another. This state of matter is called the **liquid phase**. As in a solid, in a liquid, the attractive forces are strong enough to hold the molecules or atoms close together so they are not easily compressed and have a definite volume. Unlike in a solid, however, the particles will flow (slide over each other) so that they can assume the shape of their container.

Finally, if enough energy is put into the system, the individual molecules or atoms acquire enough energy to totally break all of the attractive forces between them and they are free to separate and rapidly move throughout the entire volume of their container. This is called the **gas phase** and atoms or molecules in the gas phase will totally fill whatever container they occupy, taking on the shape and volume of their container. Because there is so much space between the particles in a gas, a gas is *highly compressible*, which means that the molecules can be forced closer together to fit in a much smaller space. We are all familiar with cylinders of compressed gas, where the compressibility of gasses is exploited to allow a large amount of gas to be transported in a very small space.

Returning to our example of water, at low temperature, water exists as the solid, ice. As the solid is warmed, the water molecules acquire enough energy to overcome the strongest of the attractive forces between them and the ice *melts* to form liquid water. This transition from the solid phase to the liquid phase happens at a fixed temperature for each substance called the **melting point**. The melting point of a solid is one of the physical properties of that solid. If we remove energy from the liquid molecules they will slow down enough for the attractive forces to take hold again and a solid will form. The temperature that this happens is called the **freezing point** and is the same temperature as the melting point.

As more energy is put into the system, the water heats up, the molecules begin moving faster and faster until there is finally enough energy in the system to totally overcome the attractive forces. When this happens, the water molecules are free to fly away from each other, fill whatever container they are occupying and become a gas. The transition from the liquid phase to the gas phase happens at a fixed temperature for each substance and is called the **boiling point**. Like the melting point, the boiling point is another physical property of a liquid.

Phase transitions for a typical substance can be shown using simple diagram showing the physical states, separated by transitions for melting and boiling points. For example, if you are told that a pure substance is 15° C above its boiling point, you can use the diagram to plot the temperature relative to the boiling point. Because you are above the boiling point, the substance will exist in the **gas phase**.

There are, however, some exceptions to the rules for changes of state that we have just established,. For example, ice is a solid and the molecules in the interior are held together tightly by intermolecular forces. Surface molecules, however, are exposed and they have the opportunity to absorb energy from the environment (think of a patch of snow on a bright sunny day). If some of these surface molecules absorb enough energy, they can break the attractive forces that are holding them and escape as a gas (water vapor) without ever going through the liquid phase. The transition from a solid directly into a gas is called **sublimation**. The reverse process, a direct transition from a gas to a solid, is called **deposition**. Perhaps the most common example of a solid that does not melt, but *only* sublimes, is dry ice (solid carbon dioxide;CO₂). This property of dry ice is what makes it a good refrigerant for shipping perishables. It is quite cold, keeping things well frozen, but does not melt into a messy liquid as it warms during shipment.

Just like surface molecules in solids can move directly into the gas phase, surface molecules in liquids also absorb energy from the environment and move into the gas phase, even though the liquid itself is below the boiling point. This is the process of **vaporization** (evaporation). The reverse process, a transition from a gas to a liquid, is called **condensation**. Liquid substances undergo vaporization and the space above any liquid has molecules of that substance in the gas state. This is called the **vapor pressure** of the liquid, and vapor pressure (at a given temperature) is another of the physical properties of liquid substances.

Summarizing what we know about the different states of matter:

In a **gas**:

- the molecules or atoms are highly separated, making a gas highly compressible,
- attractive forces between the particles are *minimal*, allowing the gas to take on the shape and volume of its container.

In a **liquid**:

- the molecules or atoms are closely spaced, making a liquid much less compressible than a gas,
- attractive forces between the particles are *intermediate*, allowing the molecules or atoms to move past, or slide over one another,
- liquids have a definite volume, but will take on the shape of their container.

In a **solid**:

- the attractive forces are strong, keeping the atoms or molecules in relatively fixed positions,
- the neighboring atoms or molecules are close together, making the solid not compressible and giving it a definite shape that is independent of the shape and size of its container.

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