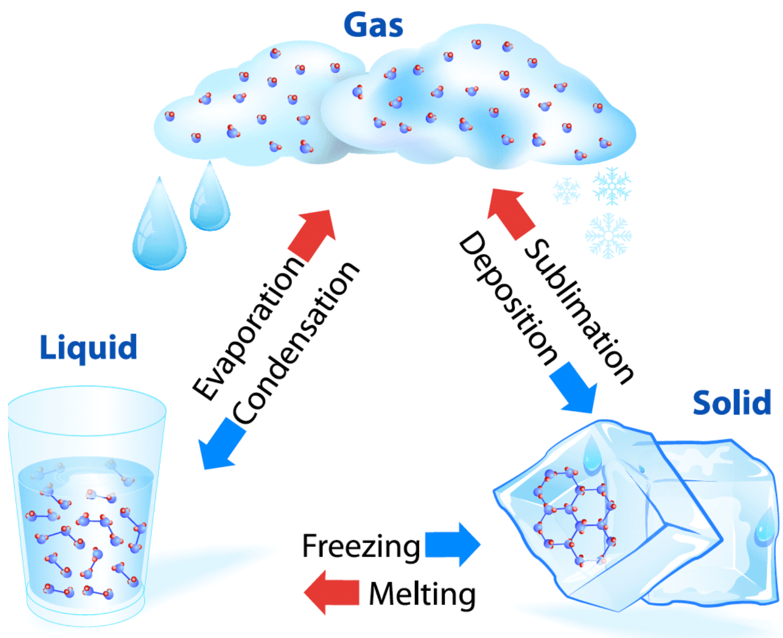
Changes of State

A **change of state**, also called a phase change, is the conversion of a substance from one of the physical states of matter to another. One example of a change of state is when a liquid evaporates from an open container.

Possible Changes of State



Energy and Change of State

The different changes of state can be categorized based on whether they *require* energy or *release* energy.

In order to melt ice, for example, it is necessary to overcome the relatively strong intermolecular forces between the water molecules. Overcoming these forces *requires* energy, which we typically supply in the form of heat. Melting, evaporation, and sublimation are all examples of changes of state that require energy.

The reverse processes (freezing, condensation, and deposition) *release* energy.

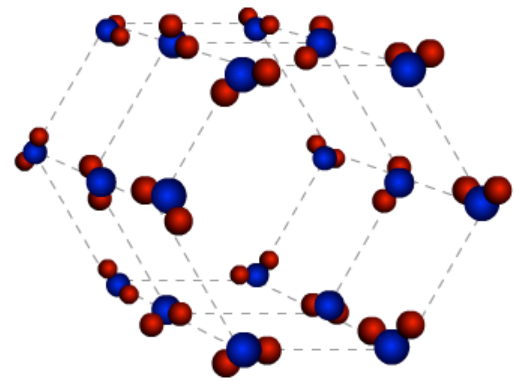
Changes of state that require energy are called **endothermic**, while those that release energy are called **exothermic**.

Endothermic Changes of State

The changes of state that require energy (heat) are melting, vaporization, and sublimation. In the next few sections we will examine how each of these processes works.

Melting

The change of state from solid to liquid is called **melting**. This change is accomplished by adding heat energy to a solid.



In a solid, such as ice, the particles are held together in a rigid, crystalline structure by strong **intermolecular forces**. In order for a solid to become a liquid, the particles must absorb enough heat energy to overcome the intermolecular forces and move apart.

When heat energy is added to a solid, the particles of the solid will absorb the energy and begin to vibrate faster. As they absorb energy, the temperature of the solid will increase. Once they have absorbed enough energy to overcome the bonds holding them together, the particles will break loose from their lattice positions and enter the liquid phase.

It is extremely important to note that **during the time when the solid is changing to a liquid, the temperature will not change!** The reason for this is that any energy added during the actual phase change is being used to break the bonds holding the particles of the solid together, thus increasing the *potential* energy of the particles – not increasing their *kinetic* energy (temperature).

During the phase change, the material will exist in both states at the same time: solid and liquid. As the material melts, the amount of solid will decrease, and the amount of liquid will increase until there is only liquid. The temperature at which the liquid phase and solid phase of a given substance can coexist is called the **melting point**.

Once all of the material has changed to the liquid phase, the temperature will once again start to increase (assuming we continue to add heat energy).

Vaporization

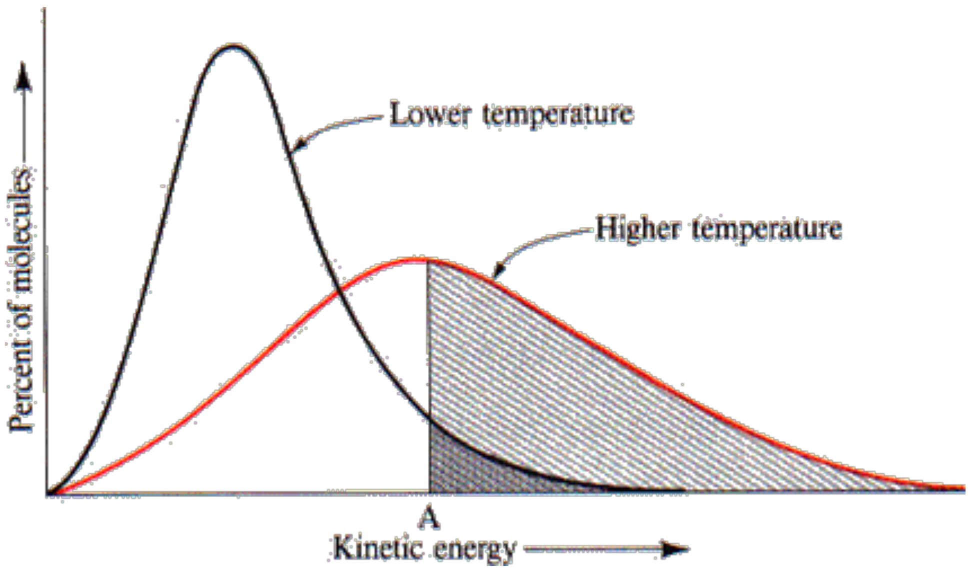
The change of state from liquid to gas is called **vaporization**. This change is accomplished by adding heat energy to a liquid.

The mechanism here is very similar to that of melting. The particles of a liquid are held together somewhat loosely by intermolecular forces. In order to enter the gas phase, the particles must absorb enough energy to overcome these forces and move far enough apart that the forces drop to nearly zero.

There are actually two different processes by which a liquid can become a gas: evaporation and boiling. We will discuss both.

Evaporation

In the liquid state, some molecules will have more kinetic energy than others. The diagram below shows how energy is distributed among the molecules in a liquid at room temperature.



The shaded portion indicates the molecules that have the energy required to overcome the intermolecular forces holding the molecules together in the liquid. If any of these molecules happen to be near the surface of the liquid, they will escape from the surface of the liquid and enter the gas phase. When vaporization occurs only at the surface of a liquid, the process is called **evaporation**.

Notice that, even at low temperatures, there are always some molecules that have enough energy to evaporate. This means that evaporation can occur at any temperature (even temperatures below freezing). As the temperature increases, more molecules will have the required energy which will result in a *faster rate* of evaporation.

Note: For a substance that is ordinarily a liquid at room temperature, the gas phase is called a **vapor**.

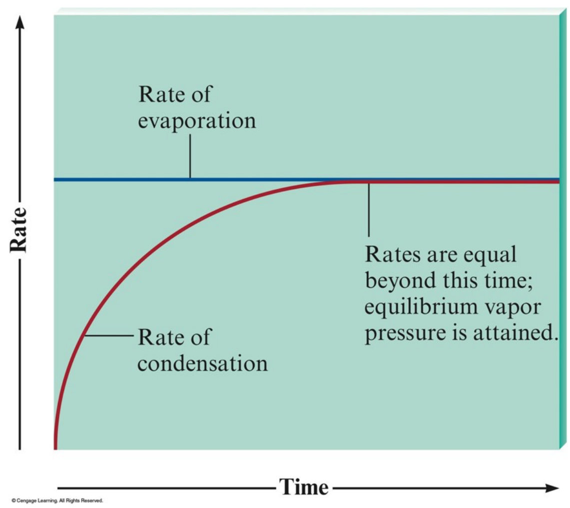
Vapor Pressure

At the surface of any liquid, there are two processes occurring. Evaporation (particles escaping the liquid) is one, and condensation (particles being recaptured by the liquid) is the other.

When a liquid is placed in an **open container**, it will evaporate until all of the liquid has vaporized. The time it takes for this to occur depends on the amount of liquid, the surface area exposed, and the temperature.

When a liquid is placed in a **closed container**, the amount of liquid will decrease at first but eventually becomes constant. The initial decrease is the result of the liquid evaporating. The rate of evaporation is constant at a given temperature.

As the number of vapor molecules increases, the rate of condensation (vapor returning to the liquid phase) increases. Eventually the rate of evaporation equals the rate of condensation. The system is at **equilibrium** — the amount of liquid and the amount of vapor are now constant.



The pressure of the vapor present at equilibrium is called the **vapor pressure**. The vapor pressure of liquids vary widely. Liquids with high vapor pressures are said to be **volatile** — they evaporate rapidly from an open container.

The vapor pressure of a liquid is primarily determined by:

1. Intermolecular Forces
    - substances with stronger intermolecular forces will have lower vapor pressures

2. Temperature (Average Kinetic Energy)
    - as the temperature of a substance increases, its vapor pressure will also increase

Boiling

At low temperatures, only particles near the surface of a liquid have enough energy to vaporize. As the temperature of the liquid increases, the number of particles that are able to enter the gaseous state also increases. Eventually, particles throughout the liquid will have enough energy to vaporize.

As particles within the liquid vaporize, they will form micro-bubbles. If the vapor pressure of these micro-bubbles is less than the atmospheric pressure, then they will collapse and the particles will return to the liquid state.

If, on the other hand, the vapor pressure of the micro-bubbles equals or exceeds the atmospheric pressure, then the bubbles become larger and rise to the surface. Once at the surface, the vapor inside the bubbles escapes. This is **boiling**.

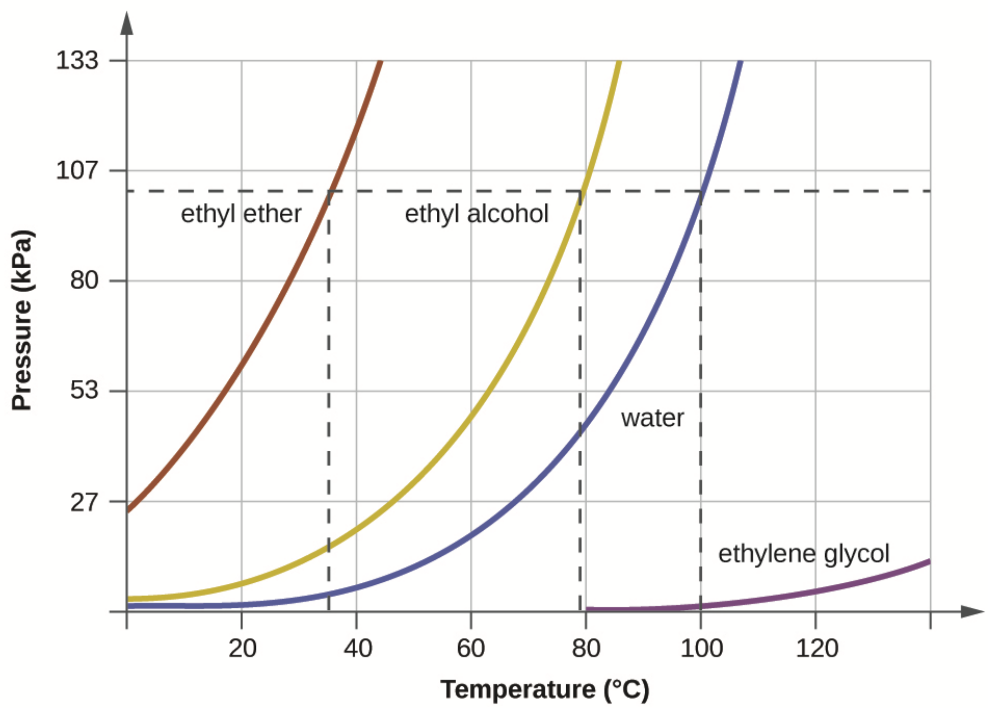
The temperature at which the vapor pressure of a liquid equals the external or atmospheric pressure is called the **boiling point**.

The boiling point of a liquid depends on the atmospheric pressure. The higher the atmospheric pressure, the higher the boiling point. Conversely, the lower the atmospheric pressure, the lower the boiling point.

The **normal boiling point** for a liquid occurs when the atmospheric pressure above the liquid is one atmosphere (101.3 kilopascals or 760 mm of mercury) – the standard pressure at Earth's surface.

Vapor Pressure Charts

The temperature dependence of vapor pressure can be represented graphically. Plots of vapor pressure versus temperature for several liquids are shown below.



From this chart you can determine the vapor pressure at a given temperature, or the boiling point at a given atmospheric pressure.

Example 1
Using the vapor pressure chart, estimate the vapor pressure of ethyl alcohol at 60°C.

Example 2
Using the vapor pressure chart, determine the normal boiling point of each liquid.

Sublimation

Sublimation is the process by which a solid changes directly to a gas without first becoming a liquid. Dry ice, moth balls, and solid air fresheners are examples of solids that are able to sublimate.

Solids exhibit vapor pressures just as liquids do, but the vapor pressure of a solid is typically quite low. In order to change into the gaseous state, the vapor pressure of the solid must equal or exceed the atmospheric pressure. For most solids, this is unlikely to occur unless the atmospheric pressure is quite low (e.g. the solid is on the moon).

Heating a solid will increase its vapor pressure. As the solid is heated, one of two things will happen.

1. The temperature of the solid will reach its melting point before the vapor pressure is equal to the atmospheric pressure. In this case, the solid will melt and enter the liquid phase.

2. The vapor pressure of the solid will become equal to the atmospheric pressure before the temperature reaches its melting point. In this case, the solid will skip the liquid phase and go directly to the gas phase. This is sublimation.

Most solids will not easily sublimate. However, solids with a high vapor pressure will sublimate relatively easily. Most of the examples of solids that sublimate belong to the molecular solids (ice, carbon dioxide, etc.). This is because molecular solids have the weakest intermolecular forces (and, therefore, the highest vapor pressures).

Many common solids will sublimate. Ice cubes left in the freezer will shrink as they sublimate. Snow on the ground in winter will gradually disappear even if the temperature is below freezing.

Sublimation is used in freeze drying. Fresh food is frozen and placed in a container that is attached to a vacuum pump. As the pressure is reduced, the ice sublimates and is removed from the container. Once the water is removed from the food, it becomes much lighter, and can be stored for long periods of time.

Worksheet #3

1. Define each of the following.

    a) evaporation
    b) sublimation
    c) boiling
    d) melting

2. Is melting endothermic or exothermic? Explain.

3. Explain what happens, at a molecular level, when a solid melts.

4. Explain how it is possible for a liquid to change to a gas at temperatures below its boiling point.

5. Why does water evaporate?

6. Define vapor pressure.

7. What is the vapor pressure of water at 100°C? How do you know?

8. Identify two factors that affect the vapor pressure of a gas.