

1.6 STATES OF MATTER

1.6.1 Phases of Matter

A material can exist in different **phases** distinguishable by some physical property such as the density and compressibility. For instance, water can exist in solid, liquid and gas forms. These phases are distinguishable by the density of water in each phase. When water exists simultaneously in two or all three of these phases, they separate into different regions, some parts exist in one phase and other parts in other phases.

Many materials exist in more than one solid phase. The solid phases usually differ from each other in crystal structures and other physical properties. For instance at least eight different phases of ice are found at high pressures. The density-based phases are not the only types of phases of interest to physicists. Magnetic materials and dielectric materials also exhibit phases with different regions of the sample organized with different magnetic or electric properties. To be specific, here, we will focus on density-dependent phases although our discussion can be applied to other types of phases as well.

1.6.2 Phases diagram

The equilibrium conditions under which a substance can exist in the solid phase, the liquid phase, the gas phase, or a mixture of phases, are usually displayed in the temperature-pressure plane. Each point on this plane corresponds to an equilibrium state of the material as shown in Fig. 1.12. The resulting picture is called a **phase diagram** of the material.

You can see in Fig.1.12 that most of the points in the phase diagram correspond to equilibrium conditions in which the material exists in only one phase. The entire region of conditions labeled solid corresponds to the equilibrium conditions under which the entire sample is in solid phase. Similarly, for the regions labeled liquid and gas.

The curves separating the pure phase regions are called coexistence lines. For instance, the line OA separates the regions of solids and gas. If you create the condition of a point on this line, then the material will exist in both solid and gas phase. By increasing or decreasing the energy of the system you can change the proportion of each phase in the material from 100% solid to 100% gas without

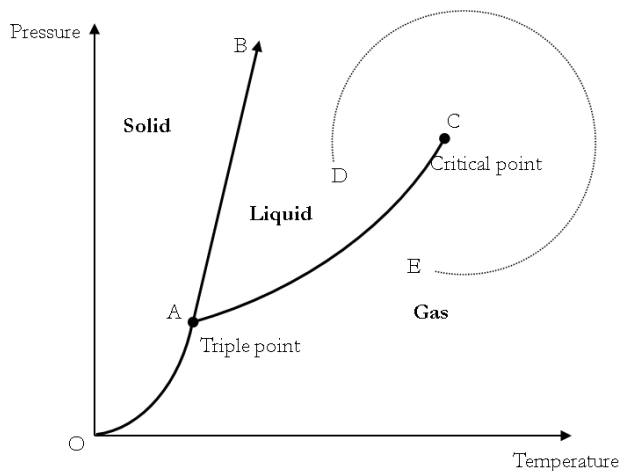


Figure 1.12: Phase diagram. Lines OA, AB and AC separate phases. The dotted line DE shows a process that converts liquid into gas without any phase separation step.

any liquid present.

For instance, if you start to a little left of the curve OA the sample will be completely in the solid phase. As you increase the temperature of the system while keeping the pressure constant, you will reach the OA line. At the OA line, the solid will be sublimated to the gas without melting. Once all the solid has been transformed into gas phase more energy will increase the temperature of the gas and you will enter the pure gas phase region of the phase diagram.

Similarly, the equilibrium curve AC between liquid and gas phases denote the boiling and condensation points, and the equilibrium curve AB between solid and liquid phases correspond to melting and freezing points. The three curves, the sublimation curve OA, the melting curve AB and the evaporation curve AC meet at a point called the triple point where all three phases coexist in equilibrium. Triple points of several substances, such as hydrogen (13.80 K, 7.04 kPa), neon (24.56 K, 43.2 kPa), nitrogen (63.18 K, 12.5 kPa) and water (273.16 K, 0.610 kPa) are used for the calibration of Thermometer.

The equilibrium curve between the liquid and gas phases terminates at the critical point beyond which the separated identities of liquid and gas phases disappear. For instance, if you start with state D which is all in liquid phase and change pressure and temperature along the dotted curve shown in Fig. 1.12, the material will undergo a phase transformation from all liquid at D to all gas phase at E without any phase separation; there would be no point at which you would have two separate densities in the sample. Since, critical pres-

tures and temperatures are well above normal temperatures for most common substances, therefore, we do not encounter strange behavior associated with the critical point in everyday life. For instance critical point of water/steam system is (647.4 K, 22,120 kPa). Boilers of electric generating plants reach well above the critical point of steam.

1.6.3 Latent Heat

Put some ice and water in a metal container so that you have an ice and water mixture at the end. Place a Thermometer in the container. It would read around 0°C . Now, if you heat the mixture, you will find that its temperature does not rise until all the ice has melted. All the thermal energy goes into changing ice to liquid water.

The same phenomenon occurs at the coexistence curve between any two phases. To develop a more general understanding of changing phases let us consider the constant pressure process $a-b-c-d-e$ shown in the phase diagram 1.13. At point a , the sample is entirely in solid. It has temperature T_a and pressure p_a . Now, we introduce some energy in the sample while keeping the pressure fixed. This can be done, for example, by heating the sample. The temperature and volume of the sample will change but we keep the pressure unchanged. If the process is conducted slowly then the sample will be in equilibrium and the temperature will change along the line ab in the diagram.

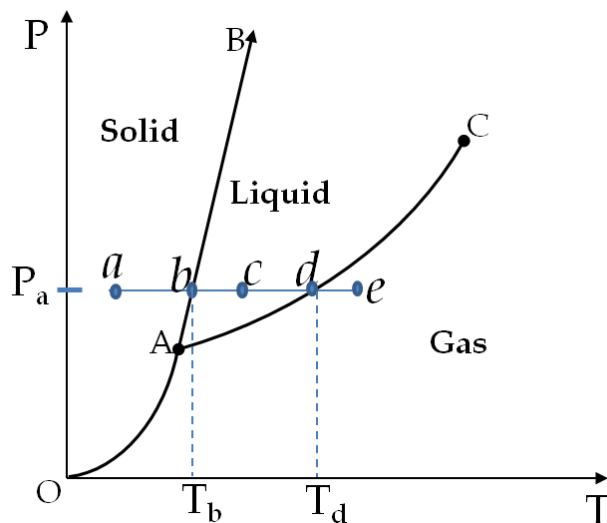


Figure 1.13: A constant pressure process in a phase diagram $a-b-c-d-e$ illustrates the phase change at the co-existence points b and d along the way in the process.

The temperature reaches T_b with pressure p_a is an equilibrium

condition for the coexistence of solid and liquid phases. As more energy is put into the system, the energy goes into loosening the molecules rather than raising the temperature. Note that raising the temperature is not possible since at point $p = p_a$ and $T > T_b$ the material does not exist in the solid phase. Therefore, at point b as we put more energy in the system, the sample must all first melt into the liquid phase. The energy that goes into transforming the substance from the solid phase to the liquid phase per unit mass of the material is called the **specific latent heat of melting**.

Once the material has completely melted at point b any additional energy in the material would raise the temperature of the material. since pressure is kept constant, the subsequent equilibrium conditions of the material will follow the path bcd in the liquid phase. During the bcd the temperature of the liquid will rise from T_b to T_d but the pressure will still be p_a . The liquid will normally expand during this process.

At the point d , the situation is similar to the point b . At the right of the point d , the material cannot exist in the liquid phase. Therefore, when we put more energy in the sample in the liquid phase at point d , the material transforms from the liquid phase into the gas phase. This transformation continues with each additional energy till all the material has been transformed from the liquid phase into the gas phase. The energy for the transformation from the liquid phase to the gas phase per unit mass is called the **specific latent heat of evaporation**.

After the material has turned completely into the gas phase at the point d , any additional energy put in the system would raise the temperature of the system. The system will successively progress in the line de of equilibrium states on the phase diagram. The energetics of the entire process $a - e$ for ice-to-steam is shown in Fig. 1.14 where heat is used to input energy in the system consisting a fixed amount of water in different phases at constant pressure.

Specific latent heats of some common substances are given in Table 1.3. If the specific latent heat of a substance for a particular transformation is l J/kg then, to transform m kg of the material heat Q involved is given by the following formula.

$$\boxed{Q = ml} \quad (1.15)$$

Heat Q will be positive if heat is being supplied to the system and negative if taken away from it.

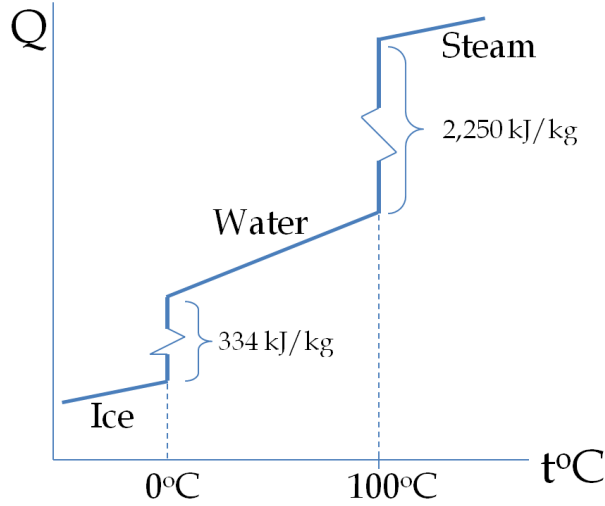


Figure 1.14: Converting ice from below freezing to steam above boiling at constant pressure. The temperature of the ice changes approximately linearly with heat input when ice is below freezing, i.e., below 0°C at 1 atm pressure. At freezing temperature, the temperature does not rise until all ice is converted into liquid water at freezing temperature. The temperature of the liquid water rises from 0°C to 100°C with input heat. At the boiling point, the temperature does not rise until all liquid is converted into steam. The slopes of Q vs temperature for ice, water and steam lines correspond to their specific heats.

Table 1.3: Specific latent heats of some common substances. The melting and boiling temperatures are at 1 atm pressure

Substance	Latent heat of fusion (kJ/kg)	Melting temperature (K)	Latent heat of evaporation (kJ/kg)	Boiling temperature (K)
Water	334	273.15	2258	373.15
Nitrogen	25	-210+273.15	199	-196+273.15
Mercury	11	-39+273.15	294	357+273
Hydrogen	60	-259+273	449	-253+273