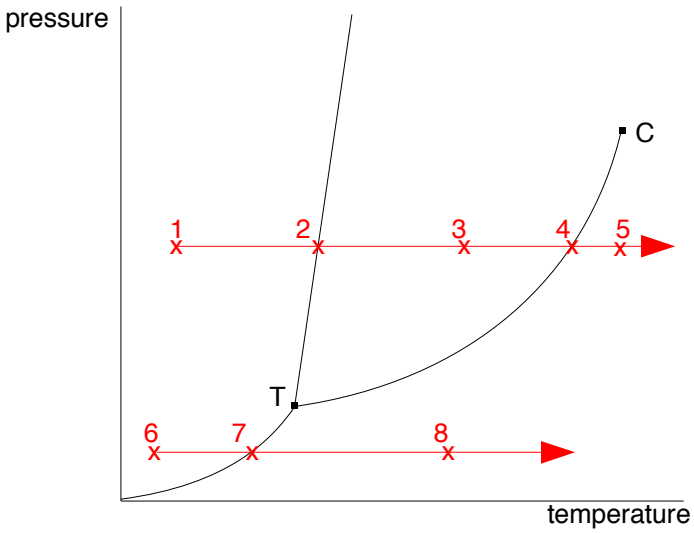
PHASE DIAGRAMS FOR PURE SUBSTANCES

1.



a) solid

b) point 2: solid and liquid (melting)
point 3: liquid
point 4: liquid and vapour (boiling)
point 5: vapour (or gas)

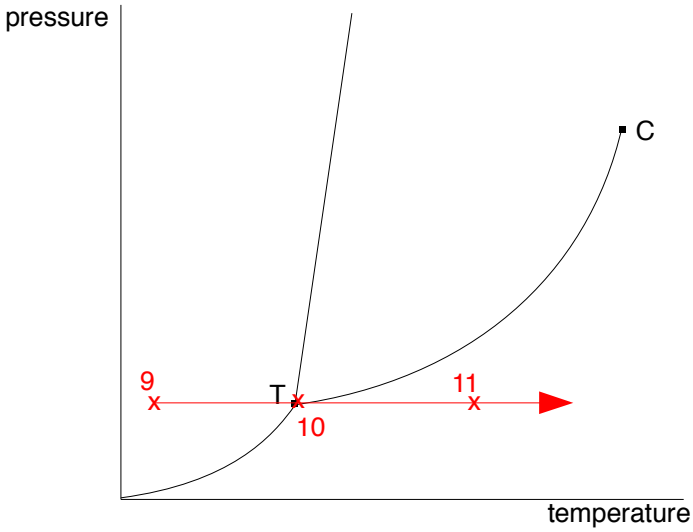
c) At 1 atmosphere, the temperature at point 2 would be the normal melting point of the substance. The temperature at point 4 would be the normal boiling point.

d) point 6: present as a solid
point 7: solid and vapour are in equilibrium and sublimation is occuring
point 8: present as a vapour

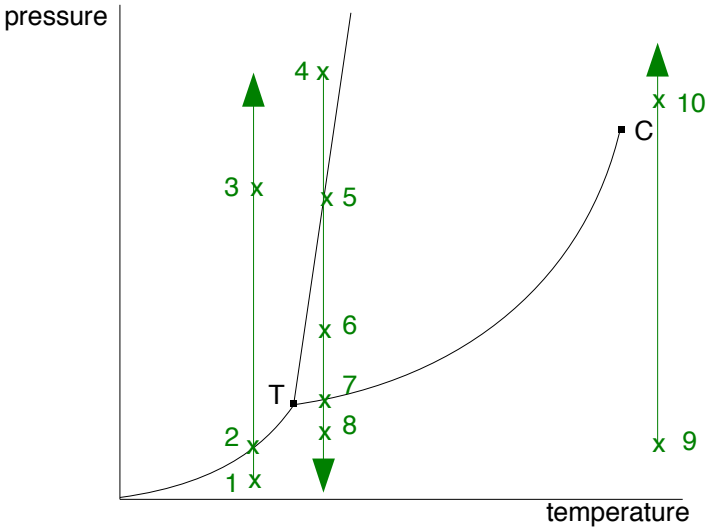
e) solid

f) point 10: solid, liquid and vapour
point 11: vapour

g) T is the Triple Point. It is the only set of conditions where all three states are found in equilibrium with each other.



2.



a) point 1: vapour
point 2: solid and vapour (sublimation)
point 3: solid

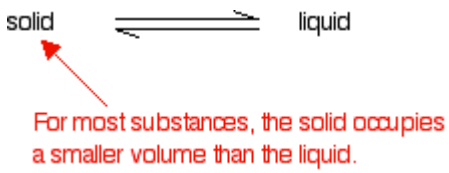
b) point 4: solid
point 5: solid and liquid (melting)
point 6: liquid
point 7: liquid and vapour (boiling)
point 8: vapour

c) point 9: gas
point 10: gas

(Because these are above the critical temperature, we normally use the term gas rather than vapour.)

d) C is the Critical Point. The temperature at this point is called the critical temperature. Above that temperature, it is impossible to condense a gas to a liquid just by increasing the pressure.

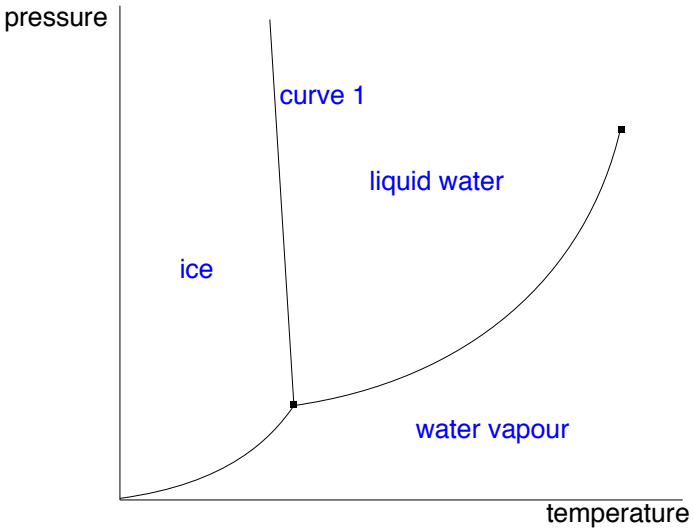
3. a) At the melting point there is an equilibrium between the solid and the liquid, and the solid usually takes up slightly less space than the liquid because the particles are slightly more closely packed.



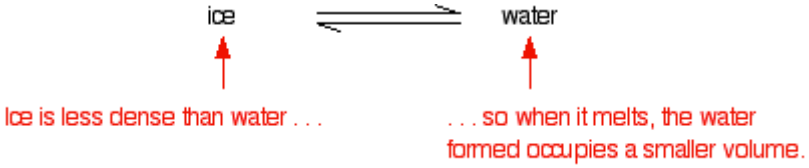
According to Le Chatelier, if you increase the pressure, the equilibrium will respond by moving in such a way as to decrease the pressure again. That means that it will move to form the solid because it takes up less room. Therefore the solid is no longer melting. To make it melt again at that pressure, you will have to increase the temperature, and so the melting point is higher at the increased pressure. The line is steep because there is very little difference between the volumes occupied by the solid and the liquid, and so the effect of pressure is fairly small.

b) At the boiling point, there is an equilibrium between liquid and vapour, but this time there is a huge increase in volume from liquid to vapour. Applying the same reasoning as in part (a), an increase in pressure will strongly favour the formation of the liquid, so that it no longer boils at the same temperature. In order to make it boil again, you need to increase the temperature by much more than in the melting case.

4. a)



b) Curve 1 represents the effect of pressure on the melting point of ice. At higher pressures, the melting point falls. This is because when ice melts, the volume of water formed is less than the ice it came from.



If you have this equilibrium and increase the pressure on it, according to Le Chatelier's Principle the equilibrium will move to reduce the pressure again. That means that it will move to the side with the smaller volume. Liquid water is produced. To make the liquid water freeze again at this higher pressure, you will have to reduce the temperature. Higher pressures mean lower melting (freezing) points.

5. The phase diagram shows that if you have solid carbon dioxide at 1 atmosphere pressure, as you increase the temperature, it will sublime at -78°C. No liquid carbon dioxide is possible at pressures of less than 5.11 atmospheres. At 1 atmosphere pressure, there is never any liquid carbon dioxide present whatever the temperature - cold solid carbon dioxide ("ice") is always "dry".