Chemguide – answers

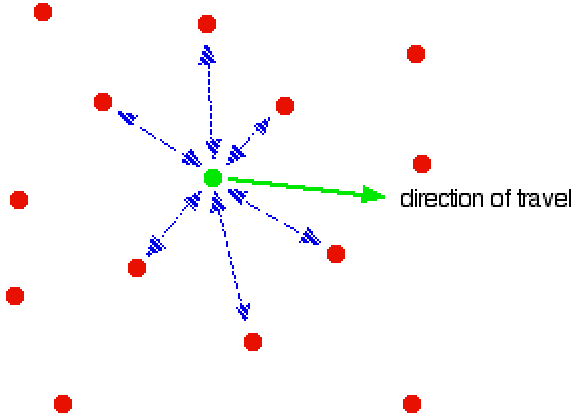
REAL GASES

1. a) The graphs would be the same horizontal straight line with a compression factor of 1 at all temperatures. For an ideal gas, $pV = nRT$. So pV/nRT will always be exactly 1.

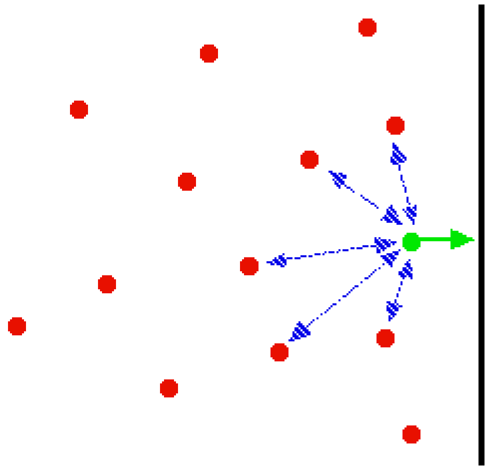
   b) An ordinary lab pressure would be around 1 bar, and the temperature would probably be a bit less than 300 K. Under those conditions, the graphs show that the compression factor is close to 1, and so $pV = nRT$ is a reasonable approximation. If a gas obeys the ideal gas equation, then we must count it as ideal under those conditions.

   c) In working out the ideal gas equation, Kinetic Theory assumes that the whole volume of the container is available to the molecules to move around in. In a real gas, some of the space is taken up by the molecules themselves. That means that the measured volume of the container that you put into the expression pV/nRT is bigger than the actual space available. As you compress the gas, the molecules take up an increasing percentage of the available space, and so the error gets greater as pressure increases.

   d) Intermolecular forces have no net effect on molecules in the body of the gas. Attractions in all directions cancel each other out, and the motion of the molecules is unaffected overall.



But just before the molecule hits the wall of the container, the net intermolecular forces are back into the gas, slowing the molecule down.



That means that it hits the wall more gently and so exerts less pressure than it would do if there were no intermolecular forces.

If the measured pressure is less than it would have been if the gas was ideal, then putting the smaller value into pV/nRT means that the value of the compression factor is also less than it should be.

e) Increasing the pressure brings the molecules closer together. That means that the effect of the intermolecular attractions will be greater, because they are only effective over a fairly short distance. The greater the effect of the intermolecular forces, the more the molecules will be slowed just before they hit the walls, and the more the pressure will be lowered relative to an ideal gas. Putting a lower pressure into the pV/nRT term obviously lowers that as well.

f) The effect of the volume of the molecules in increasing pV/nRT at high pressures becomes more important than the effect of the intermolecular attractions decreasing it. (There is no simple way of explaining why this, is. It is something you just have to accept.)

g) At lower temperatures, the molecules are moving more slowly on average. Any pull they feel back into the gas will have relatively more effect on a slow moving particle than a faster one.