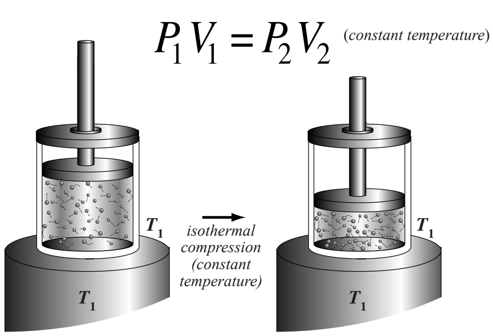
# States of Matter
## Answers and Explanations

**1. D**

Boyle's Law governs isothermal (constant temperature) transformations. At constant temperature the pressure of an ideal gas sample is inversely proportional to its volume, i.e. $PV$ = constant. If the volume of a particular sample of gas is halved the pressure doubles (if constant temperature is maintained).

The pressure increases in an isothermal compression not because the gas particles are moving more vigorously but because they have become more crowded, and a surface area in contact with the compressed gas experiences more collisions with gas particles in a given amount of time.




**2. B**

An ideal gas occupies 22.4 L at STP. A real gas will not deviate too greatly from this, especially a noble gas like helium.

$$1\text{ L}\ \frac{(1\text{ mol})}{(22.4\text{ L})}\ \frac{(6.02 \times 10^{24}\text{ molecules})}{(1\text{ mol})}$$

$$\sim\ \frac{6 \times 10^{24}\text{ molecules}}{2 \times 10^{1}}$$

$$\sim\ 3 \times 10^{23}\text{ molecules}$$

**3. D**

As can be seen in the Van der Waals equation of state, the pressure of real gases will tend to be slightly *lower* than predicted by the ideal gas law.

$$\left(P + \frac{a\,n^2}{V}\right)\left(V - n\,b\right) = n\,R\,T$$

$a$ and $b$ are specific constants for each gas.

**4. B**

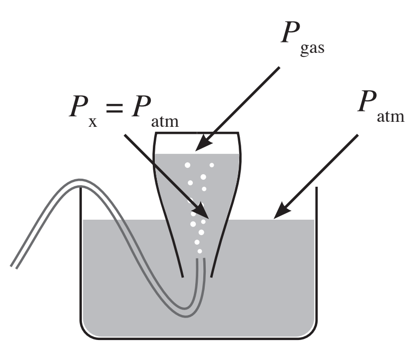
A gas takes the volume of its container, so each component has an equal volume, the volume of the container.

**5. D**

The pressure in a static fluid is the same everywhere in the fluid at the same level, even when the pressure is due to different causes. In other words, the water in contact with the trapped gas is at the same level as the water in contact with the atmosphere. Therefore, the pressure of the gas must be equal to atmospheric pressure. Note that choice 'A' is incorrect. Even though 760 torr is equal to 1 standard atmosphere (1 atm), the actual atmospheric pressure will be somewhat greater or less depending on the weather.

**6. A**

The figure below depicts the apparatus after collection of gas is first initiated.




Because the pressure in a static fluid is the same everywhere at the same level, $P_x$ must equal $P_{atm}$. Furthermore, $P_x$ is at a greater depth in the fluid than the surface of the fluid in contact with the gas, which is at the same pressure as the gas. Therefore, $P_x < P_{gas}$.

**7. D**

At the time collection is completed, the water in contact with the trapped gas is at the same level as the water in contact with the atmosphere. Therefore, the total pressure of the gas above the water must be equal to atmospheric pressure. However, the gas above the water is a mixture of argon and water vapor. The partial pressure of the water vapor is equal to its vapor pressure. At 33°C this is 37.7 torr.

Dalton's Law of Partial Pressures tells us that the total pressure of a gaseous mixture equals the sum of the partial pressures of the gaseous components of the mixture. The total pressure is equal to the atmospheric pressure, 760 torr more or less, so the partial pressure of argon is approximately 722 torr.

The partial pressure of any gas in a mixture is the total pressure multiplied by the mole fraction of that gas. Another way to say this is that the partial pressure fraction of a gas is equal to its mole fraction. The mole fraction of argon is 0.95.

**8. A**

Rearranging the ideal gas law we can see that if pressure and temperature are the same, the volume will be directly proportional to the number of moles of a gas. It's the number of particles that determines the macrosate behavior of a gas, in other words, not the size of the individual particles.

$$PV = nRT$$

$$V = n\,\frac{RT}{P}$$

Student A collected a gram of methane. Student B collected a gram of propane. Because the molecular weight of methane is lower than that of propane, a gram of methane will have more particles. With fewer grams per mole, there are more moles per gram. The volume of 1 gram of methane will be greater than 1 gram of propane.

**9. A**

Collection is completed when the level of the water in the flask equals the level in the tank. The volume is the same in both experiments, and the pressure of the gas equals atmospheric pressure. To have the same pressure and volume at a greater temperature, there must be fewer moles of gas.

**10. B**

Choice 'B' is not necessarily true. The gas is the same temperature as the atmosphere, and given the current state of the manometer, the gas must have a greater pressure than atmospheric pressure. For it to have a greater pressure at the same temperature, it must have a greater *molar density* than the atmosphere. At the same temperature molar density goes up with pressure.

$$\frac{n}{V} = \frac{P}{RT}$$

However, if the molecular weight of the gas is significantly less than the average molecular weight of the $N_2$ and $O_2$ comprising the atmosphere (if the gas were helium or neon, for example) its density (mass per unit volume) might actually be less than the density of the atmosphere despite the fact that its molar density is greater.

**11. B**

This is a straightforward ideal gas law problem, but first we need to compute the moles of $N_2$.

$$14\text{ g}\ \frac{1\text{ mol}}{28\text{ g}} = 0.5\text{ mol}$$

Now we can compute the volume.

$$PV = nRT$$

$$V = \frac{nRT}{P}$$

$$V = \frac{(0.5\text{ mol})(0.08\text{ liter atm mol}^{-1}\text{ K}^{-1})(298\text{K})}{(0.8\text{ atm})}$$

$$V = 14.9\text{ L}$$

**12. C**

It helps us to find a simple path for this problem to think about how in the ideal gas law a pressure-volume product corresponds to a number of moles of gas at a certain temperature.

$$PV = nRT$$

This means that the sum of the pressure-volume products of the two gases before admixture will equal the sum of their moles times the temperature, and, furthermore, will equal the pressure-volume product of he mixture.

$$P_{\text{mixture}}(200\ \text{ml}) = (200\ \text{torr})(100\ \text{ml}) + (250\ \text{torr})(200\ \text{ml})$$

$$P_{\text{mixture}} = \frac{(200\ \text{torr})(100\ \text{ml}) + (250\ \text{torr})(200\ \text{ml})}{(200\ \text{ml})}$$

$$P_{\text{mixture}} = 350\ \text{torr}$$

**13. D**

The boiling point of a liquid is the temperature at which its vapor pressure is equal to the pressure of the gas above it, the external pressure, which is usually atmospheric pressure. The normal boiling point is the temperature at which the vapor pressure of a liquid becomes equal to 1 atm pressure. In comparing molecules of the same size, if the intermolecular forces are small, the liquid has a high vapor pressure and a low boiling point. Conversely, liquids in which the intermolecular forces are strong will have a low vapor pressure and a high boiling point.

**14. C**

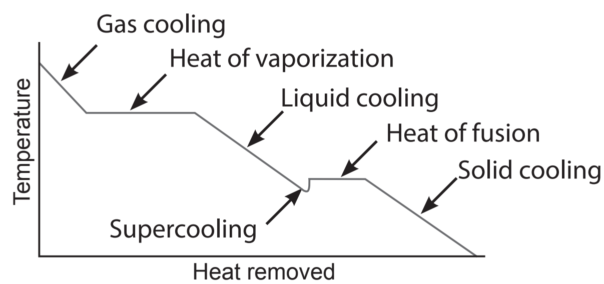
It is a practically unique characteristic of water that its solid phase is less dense than its liquid phase. This causes the equilibrium to shift from the greater volume solid phase towards the lower volume liquid phase at high pressures. Unlike with other substances, the line separating the solid and liquid phases on the phase diagram of water leans leftward.

**15. A**

The Van der Waals equation of state adjusts the ideal gas law based on the deviation of real gas behavior from ideal gas behavior. The addition to the pressure term in the ideal gas law of $an^2/V^2$ in the Van der Waals equation of state accounts for the fact that real gas molecules do in fact interact with each other and therefore show different compressibility than ideal gases. Because of intermolecular interaction the pressure is a little bit lower than one would expect based on ideal gas behavior. The value of the constant $a$ depends on the gas.

**16. C**

The challenge in this question is to recognize the narrative for the substance as heat flows out.




Supercooling is the cooling of a liquid below its freezing point without it becoming solid.

**17. A**

In this context we refer to the attractive forces between molecules of the same substance as cohesive forces. The forces between molecules of different substances are called adhesive forces. In case of mercury, the cohesive forces are stronger than the adhesive forces. As a result mercury detaches from the glass surface and its level goes down in the capillary tube.

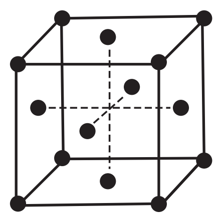
**18. B**

In a molecular solid the positions in the lattice are occupied by molecules. A molecular solid is held together by intermolecular forces.

**19. A**

Even though crystal structure is not mentioned on the AAMC outline, we've included a couple of questions to provide a very simple orientation because it would not be out of character mention of these ideas within a passage, even though there would not be an explicit penalty for a lack of foreknowledge.

The face-centered cubic structure is also called cubic closest packing and can utilize seventy-four percent of available space.



**20. B**

The body centered cubic unit cell has an atom at each of its eight corners, which it shares with eight other unit cells (8 × ⅛ atom = 1 atom). Additionally there is an atom in the center of the cell, so the total is 2 atoms per unit cell.