Chapter 10 – States of Matter

Chapter 10: 1 – 19, 21, 23, 25, 28, 29, 34, 44, 46, 51

Section 10.1 – The Nature of Gases

Practice Problems

1. What pressure, in kilopascals and in atmospheres, does a gas exert at 385 mm Hg?

$$385\ mm\ Hg\left(\frac{1\ atm}{760\ mm\ Hg}\right) = 0.507\ atm;\ 0.507\ atm\left(\frac{101.3\ kPa}{1\ atm}\right) = 51.3\ kPa$$

2. The pressure at the top of Mount Everest is 33.7 kPa. Is that pressure greater or less than 0.25 atm? Explain with mathematical proof.

$$33.7\ kPa\left(\frac{1\ atm}{101.3\ kPa}\right) = 0.333\ atm$$ 33.7 kPa is less than 0.25 atm

Section Review 10.1

3. According to kinetic theory, how do the particles in a gas move?

The particles in a gas move rapidly in constant random motion, traveling in straight paths, moving independently until they collide with other particles, or the sides of the container. As a matter of interest, oxygen molecules (O_2) travel at 1700 km/ h at 20℃!

4. Use kinetic theory to explain what causes gas pressure.

Since gas particles are always in motion, when large numbers of the particles collide on an object, they can produce a measurable force. Gas pressure is defined as the force exerted by a gas per unit surface area of an object, and is expressed as: *P = F/A*, where P is pressure, F is the force applied by the particles colliding, and A is the surface area over which these act.

5. Express the pressure 545 mm Hg in kilopascals.

$$545\ mm\ Hg\left(\frac{101.3\ kPa}{760\ mm\ Hg}\right) = 72.6\ kPa$$

6. How can you raise the average kinetic energy of the water molecules in a glass of water?

By heating up the water, the water absorbs the energy, and then the energy is converted into the energy of motion, or kinetic energy.

7. A cylinder of oxygen is cooled from 300 K (27℃) to 150 K (-123℃). By what factor does the average kinetic energy of the oxygen molecules in the cylinder decrease?

Since the Kelvin temperature of a substance is directly proportional to the average kinetic energy of its particles, if the temperature of the oxygen sample is cut in half (going from 300 K to 150 K), then the kinetic energy of the oxygen molecules is also cut in half.

Section 10.2 – The Nature of Liquids

Section Review 10.2

8. Describe the nature of liquids. Refer to the role of attractive forces in your answer.

Liquid molecules, unlike gas particles, are attracted to one another by way of intermolecular forces. The intermolecular forces are great enough to prohibit the escaping of particles to a gaseous state, but not so great that they cannot slide past one another.

9. Use kinetic theory to explain the differences between the particles in a gas and those in a liquid.

Since the kinetic energy of gas particles is much higher than that of liquids, there are no intermolecular attractive forces between them – as a consequence, the distance between the particles is vast (which is why EVERY gas, regardless of its makeup, at STP, occupies 22.4 L of volume). By contrast, the kinetic energy of liquid particles is much less, making their motion slow enough to become attracted to one another, creating very little distance between them.

10. Use kinetic theory to explain the difference between evaporation and boiling of a liquid.

The evaporation of a liquid, which occurs at its surface, can occur at a temperature much lower than that of the liquid's boiling point. When a liquid boils, the vapor pressure is just equal to that of the external pressure, resulting in bubbles of vapor forming throughout the liquid, rising to the surface, and escaping.

11. Use Figure 10.11 to determine the boiling point of each liquid.

a. ethanoic acid at 200 mm Hg

$200\ mm\ Hg\ \left(\frac{101.3\ kPa}{760\ mm\ Hg}\right) = 26.6\ kPa$ so boiling point is about 76 ℃

b. chloroform at 600 mm Hg

$600\ mm\ Hg\ \left(\frac{101.3\ kPa}{760\ mm\ Hg}\right) = 80.0\ kPa$ so boiling point is about 54 ℃

c. ethanol at 400 mm Hg

$400\ mm\ Hg\ \left(\frac{101.3\ kPa}{760\ mm\ Hg}\right) = 53.3\ kPa$ so boiling point is about 62 ℃

12. Explain why the boiling point of a liquid varies with atmospheric pressure.

At the boiling point of a liquid, the vapor pressure equals the external pressure. As external pressure increases (or decreases), the temperature needed to produce an equivalent vapor pressure, or the boiling point, must similarly increase (or decrease).

13. Why does evaporation lower the temperature of a liquid?

When the molecules within a liquid, having the highest kinetic energy escape, the average kinetic energy within the liquid is lowered.

Section 10.3 – The Nature of Solids

Section Review 10.3

14. Explain the nature of solids and tell why they differ from liquids. Refer to the organization of particles in your answer.

Particles in solids are packed tightly together (and often highly organized) and vibrate about fixed points. In liquids, particles are freer to move relative to one another.

15. How does the crystal lattice of a solid differ from its unit cell?

A crystal lattice is a solid, regular array of positive and negative ions that project out along three dimensions. A unit cell is the smallest group of particles that is representative of the crystal's shape.

16. How do allotropes of an element differ?

Allotropes of a given element have different crystalline structures. As an example, carbon comes in many allotropes, such as diamond (wherein each interior carbon is bonded to four others), and graphite (carbon atoms are linked in widely spaced layers of hexagonal arrays). Although both diamond and graphite are comprised of carbon atoms, the differing crystalline structures make them very different.

Section 10.4 – Changes of State

Section Review 10.4

17. What general information can you get from a phase diagram for water at various temperatures and pressures?

The actual state of water (solid, liquid, gas) at given temperatures and pressures, and the conditions at which phase changes occur.

18. Describe the process of sublimation. What is a practical use of this process?

During sublimation, a substance changes directly from a solid to a vapor (skipping the liquid phase). This is useful for drying substances (removing any liquid) without heating.

19. Explain triple point.

The triple point defines the conditions of temperature and pressure, at which all three phases of a substance are at equilibrium.

Chapter 10 Review

21. List the various units used to measure pressure, and identify the SI unit. 10.1

Pascal (Pa), millimeters of mercury (mm Hg), and atmosphere (atm) are all units used to measure pressure. The SI unit for pressure is the pascal.

23. Convert 190 mm Hg to the following. 10.1

a. kilopascals

$(190\ mm\ Hg)\left(\frac{101.3\ kPa}{760\ mm\ Hg}\right) = 25.3\ kPa$

b. atmospheres of pressure

$(190\ mm\ Hg)\left(\frac{1\ atm}{760\ mm\ Hg}\right) = 0.250\ atm$

25. Explain the relationship between the absolute temperature of a substance and the kinetic energy of its particles. 10.1

Kinetic energy is directly proportional to absolute temperature.

28. Express standard temperature in kelvins and standard pressure in kilopascals and in millimeters of mercury. 10.1

Standard temperature in kelvins is 270K; standard temperature is 101.3 kPa and 760 mm Hg.

29. What is significant about the temperature absolute zero? 10.1

At absolute zero, the average kinetic energy of particles is zero.

34. Explain vapor pressure and dynamic equilibrium. 10.2

Vapor pressure results from collisions of vapor particles with the container's walls. A dynamic equilibrium exists when the rate of evaporation of the liquid equals the rate of condensation of the vapor.

44. Describe what happens when a solid is heated to its melting point. 10.3

When a solid is heated to its melting point, the particles within, gain sufficient kinetic energy to overcome attractive forces.

46. When you remove the lid from a food container that has been left in a freezer for several months, you discover a large collection of ice crystals on the underside of the lid. Explain what has happened. 10.4

Moisture in the food has sublimed and then resolidified on the container lid.

51. Why is the equilibrium that exists between a liquid and its vapor in a closed container called a dynamic equilibrium?

Although there is an ongoing exchange of particles moving from the liquid phase to the gaseous phase; and from the gaseous phase to the liquid phase, the net amounts of vapor molecules and liquid molecules remain constant.