Chemistry

States of Matter
Lesson 9
Lesson Plan
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States of Matter
The Nature of Gases
Objectives: Describe the motion of gas particles according to the kinetic theory; Interpret gas pressure in terms of kinetic theory.

- The Nature of Gases
  - Kinetic Theory – what is the difference between water as ice, as a liquid and as a gas? Why is that when a lady puts on perfume we can detect it across a room? It all has to do with particle motion and the model called kinetic theory will help us find the answer
  - What is kinetic?
    - Refers to motion. The energy an object has because of motion is called **kinetic energy**
    - The **kinetic theory** states that all matter is made of particles that are in constant motion
  - Gases and the kinetic theory
    - A gas is composed of particles (usually molecules or atoms) that are considered small hard spheres that have insignificant volume, relatively far apart and have no attraction or repulsion for each other.
    - Gas particles move rapidly in constant random motion, in straight lines, independent of each other. They fill their container – regardless of the space. They change direction only upon striking the wall of the container or another particle. Did I mention fast? Average speed of an oxygen molecule at 20˚C is 1660 km/hr, which would let it travel from Washington, DC to Mexico City in about 90 minutes!
    - All collisions are perfectly elastic – that is during a collision kinetic energy is transferred without loss from one particle to another and the total kinetic energy remains the same.
  - Gas Pressure
    - While the mass of a gas particle is very small, and therefore the force that such a particle exerts on the walls of a container is very small, the force of billions of such particles is significant and measurable.
    - **Gas pressure** is defined as the force exerted by a gas per unit of surface area of an object.
    - If there are no gas particles, there can be no collisions and therefore no pressure – such a condition is called a **vacuum**.

        - **Air pressure** is caused by the collision of the molecule that form the atmosphere with objects – you, me, trees, the Earth.
        - **Barometers** are devices used to measure atmospheric pressure. While there is an average atmospheric pressure for the Earth, the local pressure is affected by the weather.
        - The SI for air pressure is called the pascal (PA). At sea level, atmospheric pressure is 101.3 kilopascals (kPa). Older units include 1 atm (atmosphere) and 760mm of Hg.
            - Galileo had a barometer of water that extended through the roof of his house. People could walk by and see it and get a feel for upcoming weather.
            - The Hg barometer works by filling a tube (one end closed) with Hg and turning it upside down in a bowl of Hg. The Hg in the tube will fall down in the tube according to the atmospheric pressure pushing back on the Hg. As noted, at sea level this level is 760 mm. As atmospheric pressure rises, it pushes the Hg back up the tube, and as atmospheric pressure falls the Hg in the tube also falls since there is less pressure on the Hg.
            - Most barometers today do not contain Hg (health hazard), but rather are aneroid barometers. These work by having a fix air pressure in a flexible metal bellows. A needle indicator moves as increasing or decreasing air pressure flexes the bellows.
            - Air pressure is measured at 0˚ C or 273˚K
    - Kinetic Energy and Kelvin Temperature
        - We just mentioned 273˚K. Hold on a minute and we will describe what Kelvin is all about
        - As you add heat to a gas, some of the energy goes is stored in the particle (called potential energy), the rest of the energy makes the particle move faster – that is increases the kinetic energy.
        - In a container of gas particles, not all the particles have the same kinetic energy – there is an average kinetic energy. Some will have more, some will have less, be most will be around the average.
        - As heat is removed from a container of gas, the average kinetic energy is reduced and the particles move slower. If we were to remove all the kinetic energy the particles would stop moving all together. This is called **absolute zero** and is 0˚K
        - The closest man has gotten to absolute zero is .0001˚K
        - The last thing to note about the Kelvin scale is that every time the temperature doubles, the kinetic energy a particle has also doubles

The Nature of Liquids

Objectives: Describe the nature of a liquid in terms of the attractive forces between the particles; Differentiate between evaporation and boiling point of a liquid, using kinetic energy.

- A Model for Liquids
  - In describing the composition of a gas we spoke of free roaming particles with no attraction or repulsion for each other.
  - The particles in a liquid are also in motion, but have an attraction for each other. They are free to slid by each other.
  - The force of attraction between these particles is called **intermolecular forces.** These force cause liquid particles to be fairly close together, thus raising the density of a liquid compared to a gas. Accounts for why an increase in pressure on a liquid or gas does little to change it's volume.
  - The particles in a liquid spin and vibrate as they move from place to place, contributing to the average kinetic energy of the particles.
  - For a particle to escape the liquid phase to go to the gas phase, it must have enough kinetic energy to overcome the intermolecular forces that hold the particles together.
- Evaporation
  - You know from experience that if you leave a glass of water out for several days it all disappears. The process of changing from a liquid to a gas is called **vaporization**. When boiling is not involved, it is called **evaporation**.
  - Evaporation takes place when a particle on the surface of the liquid has enough kinetic energy to overcome the intermolecular forces and leave the liquid. You should note that some of these particles collide with air molecules and end up rebounding right back into the liquid.
  - Heating a liquid adds heat to the liquid, which also increases the average kinetic energy of the particles, making it more likely that a particle may reach the escape velocity needed to break away from the surface of the liquid.
  - Evaporation is a cooling process – why do you sweat?, why are you cool when getting out of a swimming pool on a hot day when the wind blows?. If a particle with a high kinetic energy escapes from a liquid, then the average kinetic energy of the liquid is lowered.
  - When you have a closed container of liquid, a different situation arises. Particles with sufficient kinetic energy will escape the liquid to become a gas. These particles of gas collide with the sides of the container creating pressure, which we call **vapor pressure**.
  - Once the air in the container becomes **saturated,** then it can hold no more gas particles of the liquid. Particles will continue to escape the liquid, but some of the gas particles will condense back into the liquid. When the rate of evaporation = the rate of condensation, then we say it has reached an **equilibrium.**
  - If we were to increase the temperature of the contained liquid, the vapor pressure would increase. This happens because with a higher temperature, more particles will have enough kinetic energy to escape the liquid and therefore more particles collide with the walls of the container.
  - Vapor Pressure can be measured with a **manometer**. This is a simple U shape of glass tubing, open on one end and the other end connected to a

closed container of gas. Mercury is placed in the U. As vapor pressure rises, it will push the Hg away from the flask. The rise (or fall) of the Hg can be measured to give vapor pressure at the current temperature of the container.

- Boiling Point
  - You know that the rate of evaporation of a liquid in an open container increases when you heat the liquid. As noted before, heating raises the average kinetic energy of the particles in the liquid, allowing more to reach escape velocity.
  - When you heat a liquid to a high enough temperature, many of the particles in the liquid (not just on the surface) reach the escape velocity. This occurs at the **boiling point**, which is when the temperature at which the vapor pressure of the liquid is just equal to the external pressure. Bubbles of vapor form in the liquid, rise to the surface and escape into the air.
  - Since the boiling point occurs when the vapor pressure is equal to the outside pressure, what happens as the outside pressure changes?
    - Boiling points are normally given at sea level. If you go to the top of a mountain, the air pressure is less, so the required vapor pressure is less, which means the average kinetic energy is less, which means the temperature is less.
    - Like evaporation, boiling is a cooling process. When a particle reaches escape velocity and leaves the liquid, the average temperature of the liquid decreases. Liquids can never be heated above a certain temperature at a given pressure.
    - Pressure cookers use this fact to work. They can reach higher temperatures that boiling in an open pan, since as the pressure rises inside the cooker, the temperature must also rise to raise the average kinetic energy of the liquid particles before boiling can occur.
    - Graphs of vapor pressure vs. temperature exist to help scientist and manufacturers to know at what pressure a chemical will boil.
  - Since a liquid particle of water and gas particle of water are at the same temperature at evaporation, which one would cause a more serve burn? Gas – higher kinetic energy.

The Nature of Solids

Objectives: Describe how the degree of organization ofparticles distinguishes solids from gases and liquids; Distinguish between a crystal lattice and a unit cell; Explain how allotropes of an element differ.

- A Model for Solids
  - While particles in liquids are relatively free to slide past each other, particles in solids are not. Rather the motion of particles in a solid is a vibration around a fixed point. In most solids, particles are packed tightly in highly organized patterns. It is due to these fixed positions that solids retain their shapes and do not take the shape of their container.

- When you heat a solid, its particle vibrate more rapidly as their kinetic energy increases. Eventually the kinetic energy overcomes organization of the particles in the solids and the solid melts – becomes a liquid. This point is called the **melting point**.
- The process can be reversed by lowering the temperature and the liquid becomes a solid again – freezing. It should be noted that the melting point and the freezing point are the same temperature for a substance

- Crystal Structure and Unit Cells
  - Most solid substances are crystalline – that is a repeating 3-D pattern of a regular shape. The shape of the crystal reflects the arrangement of the particles within the solid.
  - The type of bonding that exists between atoms determines the melting point of crystals
  - Crystals have faces and angles, which the faces intersect.
    - The seven types of crystals are cubic, tetragonal, Orthorhombic, Monoclinic, Triclinic, Hexagonal and Rhombohedral
    - The smallest group of particles within a crystal that retains the shape of the crystal is called a **unit cell**
- Some solids can exist with more than one crystalline shape
  - Carbon can exist as carbon (graphite in pencils), diamonds (a girl's best friend) and in a recently discovered form as a buckyball (buckminsterfullerene)
  - **Allotropes** are two or more different molecular forms of the same element in the same physical phase.
- **Amorphous solids** have no crystalline structure.
  - Rubber, plastic, asphalt and glass are such examples
  - The atoms are randomly arranged
  - Glasses are sometimes called super cooled liquids
    - They do not have a melting point, rather they just soften

Changes of State

Objectives: Interpret the phase diagram of water at any give temperature; Describe the behavior of solids that change directly to the vapor state and recondense to solids without passing through the liquid state.

- Phase Diagrams
  - As we have seen, the relationship between a solid, liquid and gas, in a sealed container, depends on the temperature and pressure. This can best be represented by a single graph called a **phase diagram**.
  - As seen in the diagram, all three phases of water are shown, along with lines indicating a division between the phases.