States of Matter

Kinetic Molecular Theory

When discussing the properties of matter, it is not enough just to classify them. We must also create a model that helps to explain the properties that we see. The most commonly used model is the Kinetic Molecular Theory, or some variation of it. This model is based on five postulates (assumptions):

1. All matter is composed of infinitesimally small atoms (size doesn't matter).
2. These particles are in constant motions and therefore have kinetic energy.
3. The particles interact with each other through repulsions and attractions and therefore have potential energy.
4. The kinetic energy is proportional to the temperature.
5. The particles transfer energy in collisions that are elastic (no energy is lost).

We can explain the properties that we observe in the various states of matter with these postulates.

Solids are a state where the potential energies are much higher than the kinetic energies. The particles in a solid are in motion but cannot pass by one another because the attractions from neighboring atoms or molecules are too strong to overcome. As a result, all solids have a definite shape and volume. Because the molecules are about as close to each other as they can be, solids cannot be compressed. However, solids can transfer heat very well for the same reason.

Liquids have approximately equal kinetic and potential energies. This allows the atoms or molecules to move around one another but remain in contact with each other. This results in liquids having a definite volume (the particles are still in close contact) but an indefinite shape (the particles can move around each other). Liquids have a little higher compressibility than liquids, because there is a little more space between particles. Because the particles have more motion, liquids still transfer heat fairly well, but not as well as solids.

The kinetic energy in gases is much higher than the potential energy. The particles are very far apart because of this higher kinetic energy. Therefore, a gas doesn't have a definite volume or shape. The large distances between particles also means that gases are very compressible. However, they are not good conductors. This is why you can put you hand in the oven without burning it as long as you don't touch the solids in the oven.

Energy

What is energy? The simplest definition of energy is the capacity to do work. There are two types of energy kinetic energy, or the energy a moving body possesses, and potential energy, the energy a body has due to its position.

There are many different kinds of energy. Some of these are mechanical, electrical, chemical, nuclear, and radiant (light). All of these kinds of energy and the two types of energy listed above are measured in the same units.

The SI units for energy can be determined by looking at the formula for kinetic energy (or any other formula for energy).

$$E_k = \frac{1}{2}mv^2$$

Here m is the mass of the object in kilograms and v is the velocity of the object in m s^{-1}. This gives a unit of energy as being kg m^2 s^{-2}. This unit is the SI unit of energy. It has been renamed the **Joule** (J) in honor of James Prescott Joule who did some of the earlier work on energy transfer and heat.

Another unit of energy that has been used and is still in use in some places, like nutrition, is the calorie. A **calorie** (cal) is defines as the amount of energy needed to increase the temperature of 1.00 g of water from 14.5 °C to 15.5 °C. 1 calorie is the same as 4.184 J.

Heat

What is heat? It is NOT the same thing as temperature. Temperature is a measure of how much heat can flow but it is not a measure of the heat itself. Heat is a form of energy. As a result, heat is measured in Joules. Heat is energy that flows due to a temperature difference and is given the symbol **q**.

Heat that is transferred during a chemical reaction is called the **heat of reaction**. These heats can flow into or out of the reaction. In order to describe this so that everyone has the same definitions, we describe the reaction as the **system**. Everything else is the **surroundings**. When we talk about the heat transferred, we always describe it from the system's point of view.

When heat flows from the system into the surroundings, the system loses heat and therefore has a negative value ($q<0$). The temperature of the surroundings will increase in this situation. This reaction would be described as being **exothermic**. If the heat flows from the surroundings into the system, it has a positive value ($q>0$). The temperature of the surroundings will decrease in this situation. This reaction would be described as being **endothermic**.

Earlier, we stated that heat is a measure of the amount of energy that can flow due to a temperature difference. Therefore, the heat is proportional to the temperature difference. It is also proportional to the mass. In order to get these proportionalities to be equalities we must insert a constant of proportionality. This constant is called the **specific heat**, **s**, of the substance and has units of J g^{-1} $°C^{-1}$. Specific heat is defined as the amount of heat required to increase the temperature of one gram of substance by one degree Celsius. The heat measured can be calculated then as

$$q = ms\Delta T$$

Phase Transitions

All elements and compounds undergo some sort of phase transition as their temperature is increased from 0 K. The points at which these phase transitions occur depend on the substance being examined. There are three types of phase transition. Two that should be familiar are melting and boiling. The third is called sublimation and is a transition from the solid phase directly into the gas phase. Each type of phase transition has associated with it a heat (ΔH).

Melting (Freezing)	ΔH_{fus} (Heat of fusion)
Boiling (Condensation)	ΔH_{vap} (Heat of vaporization)
Sublimation (Deposition)	ΔH_{sub} (Heat of sublimation)

The heat of fusion is numerically the same as the heat of freezing, only the sign is different. Similarly, for the other two, the heat of condensation has the same magnitude as the heat of vaporization but with an opposite sign.

When a phase transition occurs the temperature does not change. This is because the heat being added to or removed from the system is involved in changing the state of the substance, not in changing the temperature. As a result, the equation used for calculating the amount of heat transferred that we used earlier cannot be used for phase transitions. The heat transferred in a phase transition is equal to the number of moles of the substance times the heat of phase change.

$$q = \pm n\Delta H_{fus,vap,sub}$$

If we were to plot the temperature of a substance versus the heat added to the substance we would not get a constantly increasing line. There would be plateaus in the line where the phase changes.

When the temperature changes, the heat transferred can be calculated with

$$q = ms\Delta T = C\Delta T$$

When the phase changes we use

$$q = \pm n\Delta H_{fus,vap,sub}$$

We can get the total heat transferred for the entire heating process by adding the heats from the individual processes (temperature changes and phase changes).

Intermolecular forces

All substances can be liquefied or even solidified if the temperature is lowered enough. This is an indication that there are some attractive forces between all molecules and atoms. Some molecules have stronger forces between them than others. Water, for instance, has very strong forces between the molecules. Evidence for this is the fact that water has a very high boiling point for a molecule of its molar mass. All other molecules with a similar molar mass are gases at room temperature and pressure. Only water is a liquid.

The weakest of the intermolecular forces are the **London forces**. All molecules possess London forces. This force arises when two molecules get close to each other. The nuclei on one molecule will start to attract the electrons on the other molecule. This sets up an induced dipole in the molecules. The induced dipole is very temporary; it lasts for only a very small fraction of a second. This is the reason this force is so weak. A molecule that possesses only London forces must get its kinetic energy low enough (i.e., low temperature) to allow the potential energy of the London force to keep it in contact with other molecules (i.e., the liquid state). The London force increases with molar mass and molecular volume. A more spread out molecule is more easily polarized than a compact one.

The intermolecular force that comes in the middle is the **dipole-dipole interaction**. This interaction exists only between polar molecules. The positive end of one molecule will attract, through electrostatic interactions, the negative end of another molecule. If the temperature is low enough, this will serve to hold the molecules together in the liquid phase.

Hydrogen bonding is the strongest of the intermolecular forces. This force requires a molecule to possess a hydrogen atom attached to a very electronegative atom (such as oxygen, nitrogen or fluorine). The electronegative atom will tend to pull electrons away from the less electronegative hydrogen atom. This leaves the nucleus partially exposed. The hydrogen atom will then try to attract any electrons that come by. Usually, these electrons are the lone pair electrons on the oxygen, nitrogen or fluorine from another molecule. The attraction is actually a weak coordinate covalent bond between the hydrogen on one molecule and the corresponding electronegative element on another molecule.

A molecule that possesses more of these forces will have a higher boiling point than one that possesses fewer of them. This explains why water, which has all three forces, has a higher boiling point than methane, which has a similar molar mass but has only London forces.

Solids

There are two distinctly different kinds of solids. The first is the amorphous kind of solid, which is represented by plastics and glass. The second is the crystalline solid, which is represented by ionic solids. We will concentrate here on the crystalline solids.

Types of crystalline solids

There are four types of crystalline solids: ionic, molecular, metallic, and covalent network. Each of these types has different physical properties.

Ionic solids
This type of solid is held together by the attraction of positive and negative ions. As a result, ionic solids tend to be very brittle. They will shatter very easily. Although they are comprised of positive and negative ions, they will not conduct electricity, unless they are melted or dissolved in water. The melting points of ionic solids range from a few hundred to a couple of thousand Kelvin.

Molecular solids
The intermolecular forces described above hold molecular solids together. They also tend to be very brittle. Because of the weaker interactions between particles the melting points of molecular solids range from a few Kelvin to a couple of hundred Kelvin. These solids do not conduct electricity under any conditions.

Metallic solids
Non-directional bonding holds metallic solids together. Non-directional bonding results from the loosely held valence electrons of metals. Each metal atom in the solid shares its valence electrons with every other atom in the solid. This is also the reason that metallic solids are excellent electrical conductors. Metals are not brittle. They are ductile (can be drawn into wires) and malleable (they can be pounded into sheets). These properties are also explained by the non-directional bonding in metals. Metals have melting points from a couple of hundred Kelvin (Mercury and Cesium) to several thousand Kelvin (Tungsten).

Covalent Network solids
Covalent network solids are essentially giant covalent molecules. Each atom in this type of solid is covalently bonded to at least two other atoms. Covalent network solids tend to be very hard but also brittle. Diamond is an example of this kind of solid. It is the hardest natural substance known but will shatter if struck with a hammer. Some examples of this type of solid are electrical conductors, like graphite. Most do not conduct electricity. The melting points of covalent network solids are very high, several thousand Kelvin. This is because covalent bonds must be broken to turn it into a liquid.