# REVIEW: TOPIC 5

NAME: ________

DUE: ________

# PHYSICAL BEHAVIOR OF MATTER

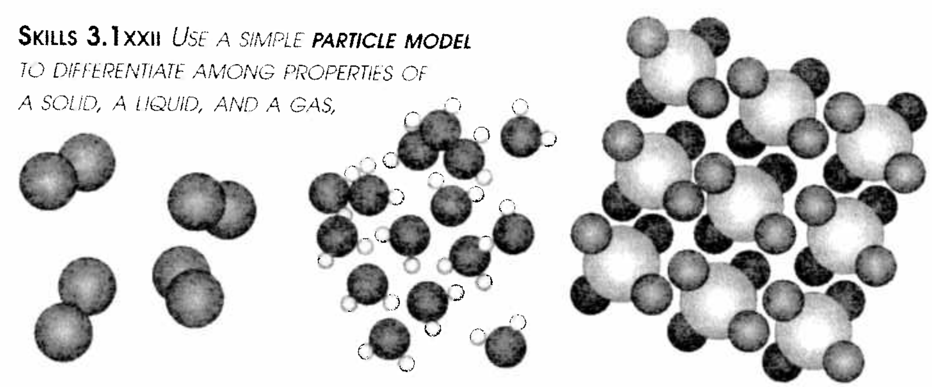
## UNIT 5 – MAJOR UNDERSTANDINGS

- ☆ 3.1q Matter is classified as a pure substance or as a mixture of substances.
- ☆ 3.1kk The three phases of matter (solids, liquids, and gases) have different properties.
- ☆ 3.1r A pure substance (element or compound) has a constant composition and constant properties throughout a given sample, and from sample to sample.
- ☆ 3.1u Elements are substances composed of atoms that have the same atomic number. Elements cannot be broken down by chemical change.
- ☆ 3.1s Mixtures are composed of two or more different substances that can be separated by physical means. When different substances are mixed together, a homogeneous or heterogeneous mixture is formed.
- ☆ 3.1t The proportions of components in a mixture can be varied. Each component in a mixture retains its original properties.
- ☆ 3.1nn Differences in properties such as density, particle size, molecular polarity, boiling and freezing points, and solubility permit physical separation of the components of the mixture.
- ☆ 3.1oo A solution is a homogeneous mixture of a solute dissolved in a solvent. The solubility of a solute in a given amount of solvent is dependent on the temperature, the pressure, and the chemical natures of the solute and solvent.
- ☆ 3.1pp The concentration of a solution may be expressed in molarity (M), percent by volume, percent by mass, or parts per million (ppm).
- ☆ 3.1qq The addition of a nonvolatile solute to a solvent causes the boiling point of the solvent to increase and the freezing point of the solvent to decrease. The greater the concentration of solute particles, the greater the effect.
- ☆ 4.1a Energy can exist in different forms, such as chemical, electrical, electromagnetic, heat, mechanical, nuclear.
- ☆ 4.2a Heat is energy transfer (usually thermal energy) from a body of higher temperature to one of lower temperature. Thermal energy is the energy associated with the random motion of atoms and molecules.
- ☆ 4.2b Temperature is a measurement of the average kinetic energy of particles in a sample of material. Temperature is not a form of energy.
- ☆ 3.4a The concept of an ideal gas is a model to explain the behavior of gases. A real gas is most like an ideal gas when the real gas is at low pressure and high temperature.
- ☆ 3.4b Kinetic molecular theory (KMT) for an ideal gas states that all gas particles: • are in random, constant, straight-line motion. • are separated by great distances relative to their size; the volume of the gas particles is considered negligible. • have no attractive forces between them. • have collisions that may result in a transfer of energy between gas particles, but the total energy of the system remains constant.

## UNIT 5 – MAJOR UNDERSTANDINGS (CONTINUED)

- ☆ 3.4d Collision theory states that a reaction is most likely to occur if reactant particles collide with the proper energy and orientation.
- ☆ 3.4c Kinetic molecular theory describes the relationships of pressure, volume, temperature, velocity, and frequency and force of collisions among gas modules.
- ☆ 3.4e Equal volumes of gases at the same temperature and pressure contain an equal number of particles.
- ☆ 4.2c The concepts of kinetic and potential energy can be used to explain physical processes that include: fusion (melting), solidification (freezing), vaporization (boiling, evaporation), condensation, sublimation, and deposition.
- ☆ 3.2a A physical change results in the rearrangement of existing particles in a substance. A chemical change results in the formation of different substances with changed properties.
- ☆ 4.1b Chemical and physical changes can be exothermic or endothermic.
- ☆ 3.1jj The structure and arrangement of particles and their interactions determine the physical state of a substance at a given temperature and pressure.
- ☆ 5.2m Intermolecular forces created by the unequal distribution of charge result in varying degrees of attraction between molecules. Hydrogen bonding is an example of a strong intermolecular force.
- ☆ 5.2n Physical properties of substances can be explained in terms of chemical bonds and intermolecular forces. These properties include conductivity, malleability, solubility, hardness, melting point, and boiling point.

**SKILLS 3.1XXII** *USE A SIMPLE* ***PARTICLE MODEL*** *TO DIFFERENTIATE AMONG PROPERTIES OF A SOLID, A LIQUID, AND A GAS.*

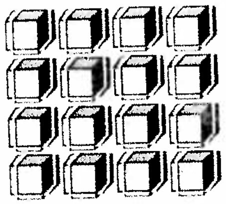


Gas - particles are random and at greater distances. | Liquid - particles are less random, closer, and better organized. | Solid - particles are arranged in a specific pattern making it rigid.

Matter is something that occupies space, has mass, and exists as a solid, liquid, or gas. Matter is classified as a substance or a mixture of substances.

# A – Phases of Matter

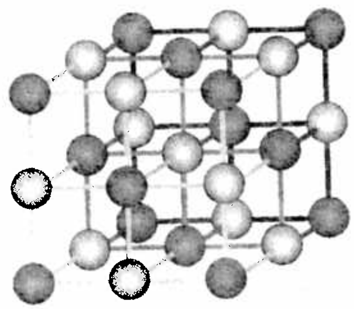
The term "**phase**" is used to refer to the gas (*g*), liquid (*l*), or solid (*s*) form of matter. Note that the word "phase" is now used instead of "state" to avoid confusion with other conditions, such as the "state of equilibrium."



Solid - particles vibrate "in place".

## Solid Phase (*s*)

In order to distinguish one solid from another, their respective densities must be compared. Solids have crystalline structures with definite shape and volume. Their characteristics are explained below.



NaCl Crystal Lattice

**Crystals** – Although the particles in a crystal are constantly vibrating, they do not change their regular positions and are arranged in a regular geometric pattern (called the crystal lattice structure).

The particles in glass and certain plastics are not arranged in a regular geometric pattern and behave as highly viscous liquids. Since all true solids have crystalline structures, these materials are often considered solids, but are really **super-cooled liquids**. Therefore, they are not considered true solids.

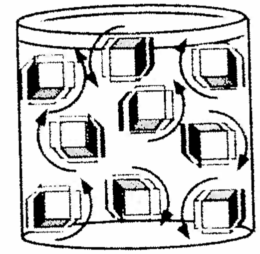

### Real World Connections

**Metallic Solids** have unique properties:

- **Strength** (as in iron and its alloys) makes metals useful in the manufacturing of cars, buildings, bridges, and machinery.
- **Malleability** (as in copper and aluminum) allows the metal to be drawn into wires. Also, since these metals have dislocated electrons, they can carry an electric current or heat energy.

## Liquid Phase (*l*)

In the liquid phase, the particles are characterized as having acquired two types of motion, vibrating and rotating. However, this phase is also considered an intermediate state having all three types of motion found in the gaseous state, but to a much more limited degree of motion with limited space between the particles. Liquid water is a good example of this phase. Liquids exhibit certain characteristics which include:



Liquid - particles vibrate and rotate.

**Vapor Pressure** – As a result of the free movement of particles in a liquid, a certain quantity of collision occurs among them, releasing energy in the process. This energy, absorbed by individual particles, allows them to acquire the translational energy with which they break those intermolecular bonds holding them to one another, and they become gaseous particles. When a liquid substance changes to a gas, the process is called **evaporation**. Evaporation tends to take place at the surface of a liquid and at all temperatures.

The term "vapor" is frequently used to refer to the gas phase of a substance that is normally a liquid or solid at room temperature. The vapor (gas) produced exerts a pressure, known as vapor pressure, which increases as the temperature of the liquid is raised and is specific for each substance and temperature.

**Boiling Point** – A liquid will boil at the temperature at which the vapor pressure equals the pressure on the liquid. The "normal boiling point" is the temperature at which the vapor pressure of the liquid equals one atmosphere (example: $H_2O$: 100°C or 373 K). Usually when reference is made to the "boiling point" of a substance, it is the normal boiling point that is indicated.

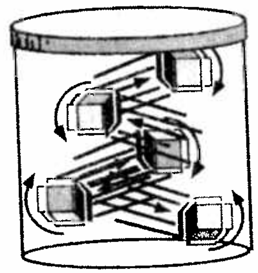
### Real World Connections

Liquids possess the following properties: Surface tension – the bonding between the water molecules on the surface of a liquid and the water molecules below them is stronger than the bond between the surface water molecules and the air molecules above them. This sets up a tight tension between the water molecules on the surface of the liquid.

Capillary action is the strength between the hydrogen bonds between the water molecules allows them to remain together while being absorbed by some blotting material so much so as to defy gravity. Viscosity of some liquids is made up of larger molecules with multiple hydrogen bonds among them which causes them to be more viscous (resistance to flow) than others.

## Gaseous Phase

In a gaseous phase, the molecules are all translating, as well as rotating and vibrating. When translating in the gaseous state, the molecules have broken the intermolecular bonds between themselves and other molecules, and the distance between molecules is great. In this state, the molecules have attained a greater amount of randomness. The degree of the randomness of a substance is defined as **entropy**



Gas - particles vibrate, rotate, and translate.

Compounds can be differentiated by their chemical and physical properties. **Properties of matter** are those characteristics that distinguish one kind of matter from another. There are two types of properties: physical and chemical.

### Physical Properties

Physical properties are those characteristics that can be described without changing the identity of the material. Physical properties are described as

- **Extensive Properties** – These are dependent on the amount of material in a system and include volume, mass, length, width, and height (i.e. the energy content in a pot of tea is greater than a cup of tea).
- **Intensive Properties** – these are *not* dependent on the quantity of material present in a system. They include melting and boiling point, color, and density (also refracture index, crystalline shape ductility and malleability).

### Chemical Properties

**Chemical Properties** describe the behavior of a substance when it reacts with other substances. In gases, oxygen supports combustion but does not itself burn. Hydrogen burns but does not support combustion.

## Physical & Chemical Change

**Physical Change** occurs when the property of matter changes but its identity does not. **Chemical Change** occurs when the identity and composition of matter is altered. For example, burning wood in air produces gases and ash which are altogether different in appearance than the wood.

As energy is supplied in a system, the atoms become more active and their movements increase to break the ridged lattice structure from just vibrations to vibration and rotational movements and become a liquid. The atomsthen proceed to transform into a vapor or gas phase where they add the movement of translation to the movements of vibration and rotation and break their inter molecular bonds and move further away from each other.

## Energy & Chemical Change

Energy is absorbed by molecules when chemical bonds are broken and liberated when bonds are formed. Usually, this energy occurs in the form of heat. During a chemical reaction, when the bonds of molecules are broken and new ones are formed, the net energy absorbed or given off depends on the strength of the bonds broken, as compared to the strength of the bonds formed. Reactions involving heat energy are classified as:

- **Exothermic reactions** – When the energy required to break the existing bonds is less than the energy given off as new bonds are formed, the excess heat energy is given off. This reaction is called an exothermic reaction, and the container in which the exothermic reaction is taking place becomes warm. For example:

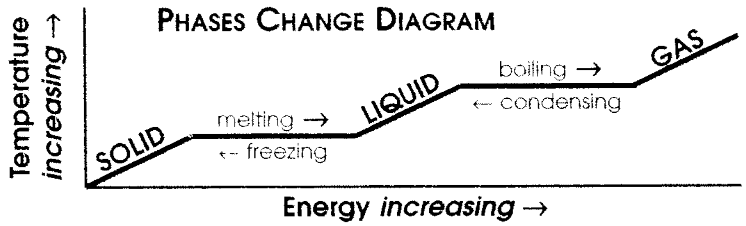
$$C + O_2 \rightarrow CO_2 + 418\ kJ$$

- **Endothermic reactions** – When the energy required to break the existing bonds is greater than the energy given off in the formation of new ones, heat energy is absorbed into the reaction. This reaction is called an endothermic reaction, and the container in which the reaction is taking place becomes cool. For example:

$$50.4\ kJ + H_2 + I_2 \rightarrow 2\ HI$$

# B – Change of Phases

The change of phase is accompanied by the absorption (called an **endothermic process**) or the release (called an **exothermic process**) of heat energy. The phase that a substance is in is dependent on temperature and pressure. Also, the phases of matter are characterized by the type of motion that the particles are undergo. **Note**: In equations that include heat, the phase of each species should be specified, such as (***g***) for gas, (***l***) for liquid, and (***s***) for solid.




Reading from left to right, the diagram shows the phase change of a solid as heat energy is added. As the temperature of the substance increases, it reaches its melting point. While a phase change takes place, the temperature remains the same (horizontal lines) until all of the sample has melted. The temperature begins to rise until the boiling point is reached. During this second phase change, the temperature remains constant until all of the sample has changed from a liquid phase to a gaseous phase.

**Melting point** – The "normal melting point" is the temperature at which a solid substance will change to a liquid at 1 atmosphere pressure. Melting point may also be defined as the temperature at which the solid and liquid phases can exist in equilibrium. At the melting point temperature, all of the particles once in a definite geometric pattern have acquired rotational energy and are no longer a part of a lattice structure. Melting points can be determined from cooling curves which are obtained experimentally.

**Heat of fusion** – The energy required to change a unit mass of a solid to a liquid at constant temperature is called its heat of fusion. Keep in mind that fusion is another word for melting. As found in *Reference Table B – Physical Constants for Water*, the heat of fusion of water is 334 J/g (joules per gram) of water. **Note**: When a gram of water is frozen into ice, the same amount of heat is given off.

**Table B**
**Physical Constants for Water**

| | |
|---|---|
| Heat of Fusion | 334 J/g |
| Heat of Vaporization | 2260 J/g |
| Specific Heat Capacity of $H_2O$ ($\ell$) | 4.18 J/g•°C |

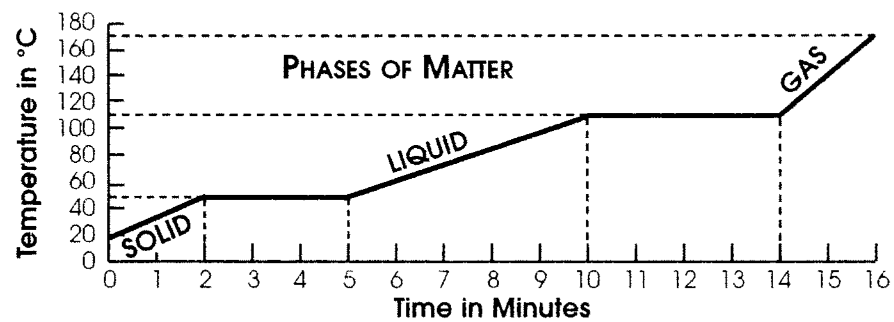
**Sublimation** – Sublimation is a change from the solid phase directly to the gas phase without passing through an apparent liquid phase. (**Note**: The opposite is **deposition**, moving from a gas to a solid.) Solids which sublime have high vapor pressures and low intermolecular attractions. Examples of solids that sublime at room temperature are solid carbon dioxide (dry ice), iodine crystals, and naphthalene.

**Heat of Vaporization** – The energy required to vaporize a unit mass of a liquid to a gas at constant temperature is called its heat of vaporization. The energy which is involved in the change of phase is needed to overcome the forces of attraction between particles (potential energy) and does not increase their average kinetic energy. Since temperature is defined as the measure of average kinetic energy in a system, there is no increase in temperature during this phase change.

As found in *Reference Table B*, the heat of vaporization of water is 2260 Joules/gram of water to be vaporized. **Note**: When a gas is changed to a liquid by condensation, the same quantity of heat is given off.

### Sample Problem:

If heat energy is being applied at the rate of 210 kJ/minute, how many calories does it take for the sample to melt? (Refer to Phases of Matter illustration below). Record the time that it takes for the sample to melt. Multiply the time by the melting rate, which is 210 kJ/minute, by the recorded time.




**Solution**: 3 min x 210 kJ/min = 630 kJ

**Note**: Both plateaus on the *Phases of Matter* graph above indicate a change of phase is taking place, and the temperature does not change during a phase change. The energy added is being used to change the potential energy of the molecules; therefore, the kinetic energy (temperature) remains constant.

Multiplying the time factor by the constant or average rate of heat energy supplied determines the joules of heat energy used during any portion of the graph. **Note**: The *Phases of Matter* graph could read from right to left as a cooling curve.

## C – Substances

A **substance** is defined as homogeneous matter when it has identical properties and composition. For example, all samples of a particular substance have the same heat of vaporization, melting point, boiling point, and other properties related to composition. These properties can be used for the identification of the substance.

Substances include:

**Elements** – All samples of an element are composed of atoms of the same atomic number and are considered substances that cannot be decomposed by chemical means. Although the periodic table of elements names just one hundred and nine elements, more have recently been discovered or made.

**Compounds** – Compounds are two or more different elements chemically combined in a definite ratio by weight. Therefore, all samples of a compound have identical composition and can only be decomposed by chemical change. Those compounds, which are made up of just two elements, are called binary compounds. In addition, those compounds that are made up of three elements are called ternary compounds.

The properties of a compound are quite different from the separate elements which make them up. For example, sodium is a soft, very active, metal that must be stored under benzene so that it does not react with air. Chlorine is a green, deadly gas. They combine in a definite weight ratio of 23g of sodium to 35.5g of chlorine. When combined, they form the compound sodium chloride (common table salt) which is stable in air and is required, to a limited extent, for the body to function normally.

## D – Mixtures

**Mixtures** are combinations of varying amounts of two or more distinct substances (either elements or compounds) that differ in properties and composition. Mixtures may be homogeneous or heterogeneous.

**Homogeneous mixtures** have a **uniform** intermixture of particles and are called solutions. A **solution** is produced when one substance dissolves or dissociates in another. Examples include:

- gas in gas – such as air
- solid in liquid – such as salt dissolved in water [$NaCl(aq)$]
- solid in solid – an alloy, such as brass, a combination of copper and zinc
- liquid in liquid – such as alcohol and water, in which the components are considered miscible [$C_3H_5OH(aq)$]

**Heterogeneous mixtures** have uniformly dispersed ingredients. Examples include:

- iron and sulfur
- oil and water
- concrete
- sand and water

Mixtures differ from compounds in that the amounts of the different substances which make up mixtures are not in a fixed ratio by weight, whereas, in compounds they are. For example, a sand and water mixture can contain various quantities of sand and water and still be considered a mixture of sand and water. Also, the components of a mixture can be separated by physical means and do not lose their identity. The sand and water mixture would be separated by either boiling off the water or by filtering out the sand. These are both physical means of separation.

Finally, compounds are made up of elements, but mixtures can be made up of either elements or compounds.

- **Filtration** – Soluble substances can be separated from insoluble substances by passing the mixture through a filter.
- **Distillation** – A liquid is vaporized and re-condensed. Using this procedure can find the percent of water in vinegar.
- **Density** – A lead shot and water mixture can be separated by filtration or by decanting the water off the top of the lead shot.
- **Particle Size** – Sight and hand can be used to separate objects.
- **Magnetic separation** – A magnet can be used separate iron ore from sulfur.
- **Boiling Point** – A salt and water mixture can be separated by boiling off the water, leaving the salt behind.
- **Freezing Point** – Cooling a salt water solution below 0°C will cause the water to freeze into ice, separating it from the salt.
- **Solubility** – A salt and sand mixture can be placed in water, allowing the salt to dissolve and the sand to settle to the bottom to be filtered off.

**Emulsifiers** – An **emulsion** is a heterogeneous suspension of a liquid in a liquid such as vinegar in oil as in mayonnaise or water in cream as in ice cream. The agent that allows them to bind together is called an **emulsifier**. In mayonnaise and ice cream the emulsifier is eggs added to the liquids and mixed vigorously until it has a creamy consistency.

**Alloys** – Alloys are industrial metals which contain significant amounts of other elements. Natural impurities (metals) are found in combined form with oxides, sulfides, or carbonates. Artificial impurities (as found in alloys) are produced by melting the metal together with others, then cooling the mixture. The resulting alloys often have properties unlike those of the separate elements. An alloy

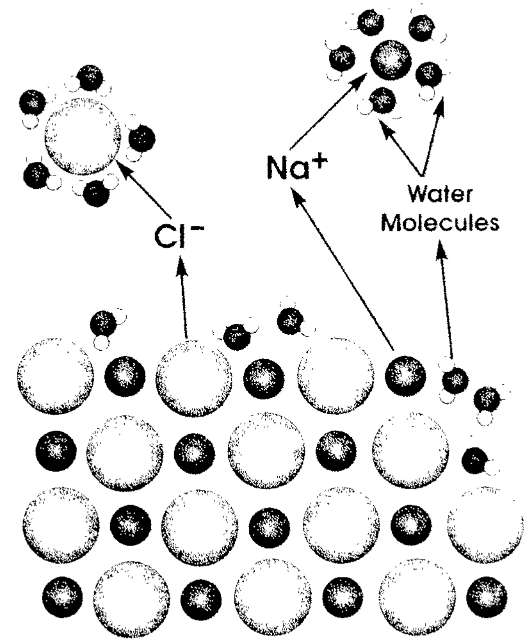
- tends to be harder than its component parts (silver-copper alloy is used to make sterling silver);
- tends to have a lower melting point than its component parts (Woods' metal composed of tin, cadmium, and lead melts in hot water but none of its components do);
- tends to be a poor conductor of electricity (Nichrom, a nickel-chromium alloy has such a great resistance to electric current flow, it is used as a heating element in toasters);
- tends to be less active than its separate elements (stainless steel is very slow to act with most chemicals). Some examples of alloys with their component parts include bronze w/ copper and tin; brass w/ copper and zinc; stainless steel w/ iron, chromium, carbon, manganese, and nickel; gold coins w/ copper and gold.

**Colloids** – Colloids are 1 to 1,000 nanometers size particles, suspended in a media which are not detected by the naked eye and appear to be homogeneous, but the dispersions are readily detected when a light is scattered by it. This scattering of light by colloidal dispersion is called the **Tyndall Effect** (e.g., fog is a liquid in gas; smoke is a solid in a gas; aerosol foam is a gas in a liquid; liquid emulsion is a liquid dispersed in a liquid).

# E – Solutions

A **solution** is a homogeneous mixture of two or more substances, the composition of which may vary within limits. The component of a solution which is usually a liquid and is present in excess is called the **solvent**, while the other component which is dissociated (dissolved) in the solvent is called the **solute**. The dissociation of ionic solute particles by a solvent is called **solvation**.

Most solutions dealt with in beginning courses in chemistry are aqueous solutions. When water is the solvent, the dissociation of solute particles is called **hydration**.




Solvation – When NaCl (ionic crystal) is added to water, the $Na^+$ ions attract the negative end of the water dipole and the $Cl^-$ ions attract the positive end. This results in dissociation.

**POLAR SOLUTES DISSOLVE IN POLAR SOLVENTS**

**NONPOLAR SOLUTES GENERALLY DISSOLVE IN NONPOLAR SOLVENTS**

## Methods of Indicating Concentrations

The main way of measuring and describing the concentration of solutions is through **molarity**. The molarity (M) of a solution is the number of moles of solute contained in a liter (1000 mL) of solution. The formula is:

$$\textbf{molarity (M)} = \frac{\textbf{number of moles of solute}}{\textbf{1 liter of solution}}$$

Therefore, a two molar (2M) solution contains 2 moles of solute per liter of solution and 0.1 molar solution, (0.1M) contains 0.1 mole of solute per liter of solution. By rearranging the above formula, it can also be state that the concentration in moles per liter multiplied by the volume in liters equals the number of moles of solute in the solution, or

$$\textbf{moles of solute} = \textbf{molarity} \times \textbf{volume in liters}$$

**Sample Problem 3**: What is the molarity of a solution of KOH if a 500 mL of the solution contains 5.6 grams of KOH?

**Solution**: First, find the moles of KOH by substituting in the equation:

$$\textbf{Grams of solute} = \textbf{no. of moles} \times \textbf{mass of one mole}$$

$$5.6\text{ g} = X \times 56\text{ g/mole}$$

$$\frac{5.6\text{ g}}{5.6\text{ g/mole}} = X$$

$$0.1\text{ mole} = X$$

Now, find the molarity by substituting in the formula:

$$\text{Molarity (M)} = \frac{\text{moles of solute}}{\text{liter of solution}}$$

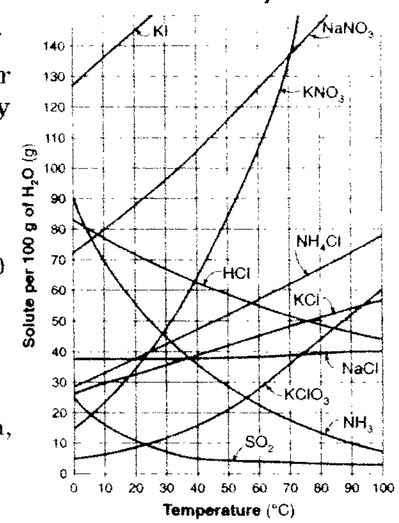
$$X = \frac{0.1\text{ mole}}{0.5\text{ liter}} = 0.2\text{ M}$$

## Other Descriptor for Concentration of Solutions

- A **miscible solution** is a solution of liquid solutes that are soluble in liquid solvents, such as alcohol in water; whereas an **immiscible solution** is a solution of liquid solutes that are insoluble in liquid solvents, such as oil in water.

- A **dilute solution** is a solution in which a large amount of solvent is required to dissolve a small amount of solute. For example, at 30°C, only about 12 grams of Cesium Sulfate will dissolve in 100 grams of water. However, a **concentrated solution** is a solution in which a large amount of solute can be dissolved in a small amount of solvent. For example, at 30°C, about 97 grams of Sodium Nitrate will dissolve in 100 grams of water.

Table G Solubility Curves




- **Percent by volume** refers to the percentage of solute per volume of solution or percentage of solute per volume of solvent, and **percent by mass** refers to the number of grams of solute in 100 grams of solvent, usually water. It is represented by a solubility curve whose vertical axis is grams of solute/100 grams of solvent and horizontal axis is temperature change.

- **Saturated solution** is a solution, which under specific conditions, holds all of the solute that it is capable of holding in a dissolved state. At this point, the liquid and solid phase are in a state of equilibrium. For example, according to *Reference Table G*, any points on the line graphs represent saturated solutions at that temperature.

- **Unsaturated Solution** is a solution in which less solute is dissolved than is capable of being dissolved under specific conditions. For example, at a specific temperature, a point below any line in *Reference Table G* indicates that the solution is not in equilibrium and, therefore, unsaturated at that temperature.

- **Supersaturated Solution** is a solution in which more solute is dissolved than can be dissolved under specific conditions. For example, at a specific temperature, a point above any line graph on *Reference Table G* indicates that the amount of solute dissolved in solution is greater than is normally dissolved and is called supersaturated.

**Note**: **Parts per million (ppm)** refers to the amount of solute per million parts of solution.

# F – Effect of Solute on Solvent

The presence of dissolved particles affects some properties of the solvent. Properties which depend on the relative number of particles rather than on the nature of the particles are called **colligative properties**. Colligative properties, as related to solutions, include changes in boiling point, freezing point, vapor pressure, and osmotic pressure.

The effect on the boiling point and the freezing point on a solvent by the addition of a solute is measured by knowing the **molality** of the solution. The molality of a solution is an expression of the solution's concentration and is defined as the number of moles of a solute dissolved in 1,000 grams (1kg) of solvent. It is arrived at by using the following formula:

$$\textbf{molality (m)} = \frac{\textbf{moles of solute}}{\textbf{kg of solvent}}$$

**Boiling Point Elevation** – The presence of a nonvolatile solute raises the boiling point of the solvent. The amount of increase is proportional to the concentration of dissolved solute particles.

One mole of particles per 1000 grams of water raises the boiling point of water 0.52°C. The relationship between moles of solute and grams of solvent is expressed in the concentration unit, molality.

**Freezing Point Depression** – The presence of a solute lowers the freezing point of the solvent by an amount that is proportional to the concentration of dissolved solute particles. One mole of particles per 1,000 grams of water lowers the freezing point of water 1.86°C.

**Abnormal Behavior of Electrolytes** – Non-electrolytes dissolve in solution to form molecules which are not charged and will not conduct an electric current.

**Electrolytes** dissolve in solution and dissociate into ions that carry an ionic charge and, therefore, they do conduct an electric current.

A mole of sugar ($C_{12}H_{22}O_{11}$), which is a non-electrolyte, will dissolve in 1 kg of water to give a mole of sugar molecules. A mole of salt (NaCl), which is an electrolyte, will dissolve in the same amount of water and also dissociate to give two moles of ions (1 mole of $Na^+$ plus 1 mole of Cl). Therefore, a 1 mole solution of NaCl will increase the boiling point of the solvent by +1.04°C. Whereas, a 1 mole solution of sugar will increase the boiling point of the solvent by +0.52°C. The same solution of NaCl will lower the freezing point of the solvent by -3.72°C. Also, a 1 mole solution of sugar will lower the freezing point of the solvent by -1.86°C. This behavior of electrolytes in solution gives evidence of the existence of ions.

## Real World Connections

### Winterizing for Safety

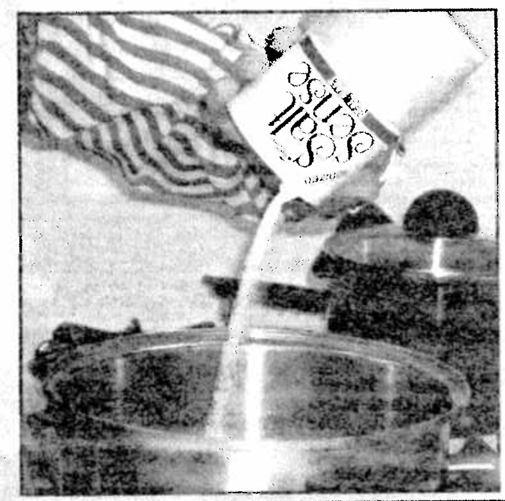
When salt is applied to an icy surface, the freezing point of the water is decreased, and the ice turns to liquid. Ice cream manufacture in the home is accomplished by the addition of salt to the ice surrounding the bowl which holds the cream, milk, sugar, and flavoring. This causes the ice to melt the water and reach temperatures colder than the ice. In the process, this causes the bowl to become cold enough to allow the ice cream to form. this principle is also used to antifreeze/engine coolant and airplane deicing solutions containing ethylene glycol as a solute instead of salt.

N&N's Ed Stich: "Softening" Water
By decreasing the freezing point of water, ice is phase changed into liquid
©Stich

## Real World Connections

### Fire Burn & Cauldron Bubble

One application of adding salt to water and bringing it to a boil is in the cooking of pasta or potatoes. The water boils at a higher temperature and cooks the pasta or potatoes at a higher temperature, therefore they will cook faster.



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# G – Energy

The ***Law of the Conservation of Energy*** states:

*Energy May Be Converted From One Form To Another But Is Never Created Or Destroyed.*

## Forms of Energy

Mechanical energy, heat energy, radiant energy (such as light, radio waves, and all other forms of electromagnetic radiation), chemical energy (derived from movements of electrons in forming bonds), and nuclear energy (as in fission or fusion reactions) are examples of energy. Energy is broken down into two types: Potential and kinetic.

**Potential energy** is the energy of position (such as, dammed up water). **Kinetic energy** is the energy of motion (such as, a moving truck).

The classical example is the process of starting up a car. In this process, energy is converted from electrical energy from the battery, to chemical energy from combustion, and then to mechanical energy.

**Heat** is the result of a transfer of thermal energy (which is associated with the random collisions of atoms or molecules) between two systems and always flows from a body of higher temperature to a body of lower temperature.

**Temperature** is a measure of the heat intensity of a body and defined as the average kinetic energy of the particles of a system.

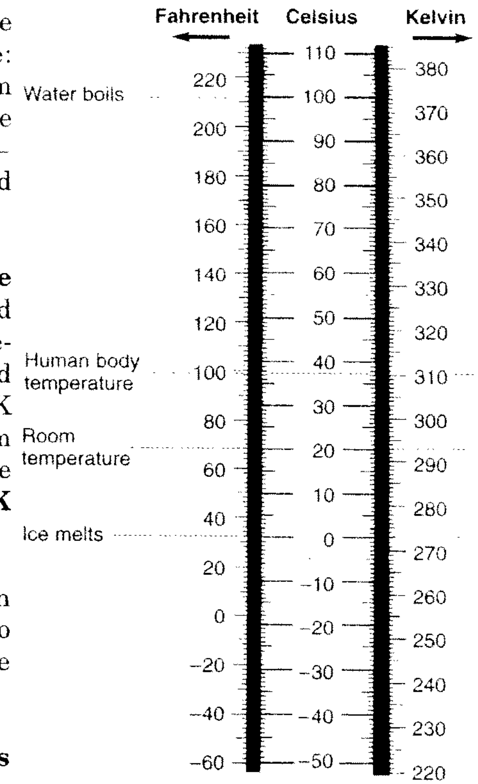
**Thermometers** are instruments used to measure temperature. Most contain liquid mercury, which has the advantage of remaining liquid over a wide range of temperatures, and also has the advantage of expanding and contracting evenly. There are two measurement scales frequently used by scientists in calibrating thermometers.

The **celsius** (°C) temperature scale has fixed points. They are: 0°C at the ice-water equilibrium temperature and 1 atmosphere pressure, and 100°C at the steam–water equilibrium temperature and 1 atmosphere pressure.

The **Kelvin** (**K**) or **absolute temperature scale** also has fixed points. They are: 273 K at the ice-water equilibrium temperature and 1 atmosphere pressure, and 373 K at the steam–water equilibrium temperature and 1 atmosphere pressure. **Absolute zero or zero K (kelvin) is equivalent to -273°C.**

When the temperature is given in degrees celsius, but one needs to find the temperature in kelvin, one should use the following equation:

**Kelvin = 273 + degrees Celsius**




## Specific Heat

The **specific heat** of a material is the amount of heat energy required to raise the temperature of 1 gram of the material one degree Celsius.

For determining reaction heats in calories, the value of the specific heat of water is considered a standard. It is 1 calorie per gram of water, per degree Celsius or cal/g°C. One calorie is equivalent in energy to 4.18 joules. When determining the reaction heat in joules, the value of the specific heat of water is 4.18 joules per gram of water, per degree Celsius or 4.2 J/g°C (rounded).

# H – Kinetic Molecular Theory for Ideal Gases

Studies of gas behavior have led to a model referred to as the "Ideal Gas Model." It is based on several assumptions.

- A gas is composed of individual particles which are in continuous, random straight line motion.
- Gas particles are separated by great distances relative to their size. Therefore, the volume of gas particles is considered negligible.
- Gas particles are considered as having no attraction to each other.
- The *Collisions Theory* states that a reaction is most likely to occur if the reactant particles collide with the proper energy and orientation.

Ideal Gas Models can be useful in the study of the behavior of gases. It should be emphasized that a model is only an approximation and is only as good as its ability to predict behavior under new conditions.

## Deviations from the Gas Laws

The **Ideal Gas Model** does not exactly represent real gases under all conditions. Hydrogen ($H_2$) and Helium (He) are the two most ideal gases. No real gas follows the ideal model under all conditions of temperature and pressure. Deviations from the gas laws occur because the model is not perfect. That is, gas particles have volume and exert some attraction for each other.

**Note**: These factors become significant under conditions of relatively high pressure, low temperature, and decreased velocity due to increased molecular mass.

## Gases

The space between molecules in a gaseous phase is about 1,000 times greater than in a liquid or solid phase. Molecules possess greater kinetic energy and have overcome the attractive forces that hold them together. At 0°C and 1 atmosphere they independently travel in random directions at a speed of about $10^3$m/sec. At this speed, it is estimated that they travel about $10^{-7}$m before they collide with each other or the walls of their container. the frequency at which they collide has been estimated to be $5 \times 10^9$ per second. Therefore, the density of a gas is lower as compared to a liquid or solid.

## Combined Gas Laws

In studying the gases, three variables are concerned: volume, pressure, and temperature.

The combined gas law equation may be written:

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$P_1$, $V_1$, and $T_1$ are the original conditions of pressure, volume and kelvin temperature. $P_2$, $V_2$, and $T_2$ are the corresponding values of the final conditions. **Note**: The definitive units given for both pressures must be the same, and both temperatures must be in kelvin degrees.

**Sample Problem**: A gas that behaves ideally at STP occupies 1000 milliliters. What volume will it occupy at 546 K and 0.5 atmospheres?

**Solution**: First, make up a list of the known items (the facts):

| | |
|---|---|
| $P_1$ = 1 atm | $P_2$ = 0.5 atm |
| $V_1$ = 1000 mL | $V_2$ = $X$ |
| $T_1$ = 273 K | $T_2$ = 546 K |

Next, use the Combined Gas Law equation as noted in *Reference Table T*:

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

Finally, put the known values into the equation and solve for $X$.

$$\frac{1 \text{ atm} \times 1000 \text{ mL}}{273 \text{ K}} = \frac{0.5 \text{ atm} \times X}{546 \text{ K}}$$

$$\frac{1 \text{ atm} \times 1000 \text{ mL} \times 546 \text{ K}}{273 \text{ K} \times 0.5 \text{ atm}} = X$$

$$4{,}000 \text{ mL} = X \text{ (answer)}$$

## Standard Temperature & Pressure (STP)

As noted above, there are three variables which we must consider when we study gases. They are volume, temperature, and pressure. Since the volume of a given mass of gas varies with changes in temperature and pressure, cannot be given the volume of a gas without specifying its temperature and pressure. Therefore, gas volumes are usually calculated to an arbitrary standard, which are abbreviated as STP. Standard temperature and pressure (STP) of a gas are defined as 0°C (273 K) and 101.3 kPa or 1 atmosphere pressure. Standard temperature for those phases other than a gas is 25°C (298 K).

## Avogadro's Hypothesis

Avogadro's hypothesis states that equal volumes of all gases under the same conditions of temperature and pressure contain equal numbers of particles. For example, at the same temperature and pressure, the number of particles in 1 liter of hydrogen is the same as the number of particles in 1 liter of oxygen although the individual particles of oxygen are heavier (by a ratio of 16 to 1) and larger than the individual particles of hydrogen.

The amount of matter than contains $6.02 \times 10^{23}$ (Avogadro's number) particles is called a **mole of matter**. One mole of any substance contains as many particles (molecules, atoms, ions, etc.) as there are atoms of carbon-12 in 12.000g of carbon-12 isotope.

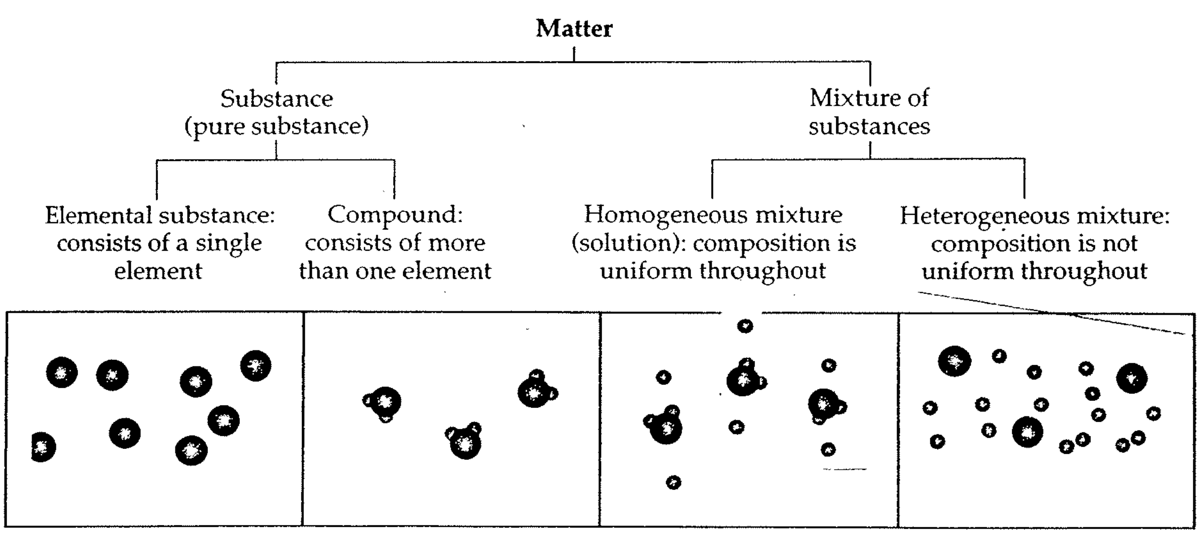
Since it is inconvenient to work with individual particles, chemists have chosen a unit containing many particles for comparing amounts of different materials. The mole is a unit which contains $6.02 \times 10^{23}$ particles. A mole of particles of any gas occupies a volume of 22.4 liters at STP and is called a **molar volume**.

...Unit 5 PHYSICAL BEHAVIOR OF MATTER

Key Concepts:

- Matter
- Element
- Compound
- Homogeneous Mixture
- Heterogeneous Mixture
- Kinetic Molecular Theory
- Avogadro's Law
- Pressure, Temperature and Volume Relationships
- Gas Law Calculations
- Heating/Cooling Curves
- Fusion
- Deposition
- Sublimation
- Vaporization
- Boiling
- Vapor Pressure
- Solubility
- "Like dissolves Like"
- Freezing Point Depression
- Boiling Point Elevation
- Concentration; ppm, Molarity
- Temperature
- Heat Transfer
- Energy Calculations
- Significant Figures

I. ***Matter*** is anything that occupies space and has mass. Matter exists as a solid, liquid or gas; and can be classified as a pure substance or a mixture of substances. A pure substance (element or compound) has its own unique properties and composition throughout. ***Elements*** are composed of a single atom and cannot be broken down by chemical means. ***Compounds*** are two or more different elements chemically combined in a set ratio by mass. Mixtures are combinations of varying amounts of two or more substances that differ in properties and composition. Mixtures are ***homogeneous***, uniform throughout or ***heterogeneous***. The proportions of the substances that make up the mixture remain.



## II Separation of Mixtures

Physical properties can be used to separate mixtures.

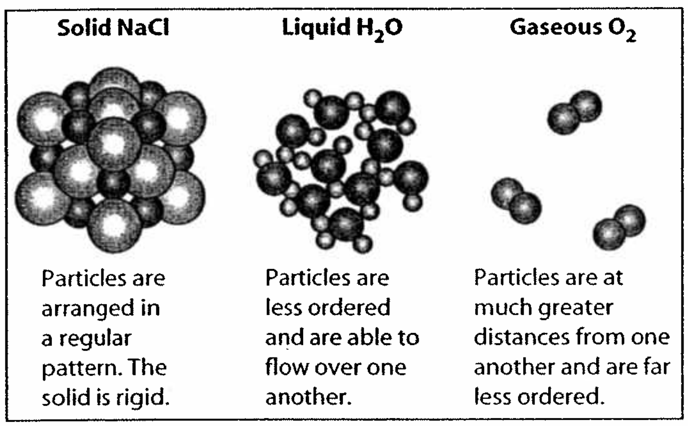
B. filtration separates insoluble substances from a solution

C. distillation separates based on differences in boiling point or freezing point, the lower boiling point substances boil out of the solution first

D. chromatography separates particles based on attractive forces and polarity, more attractive forces the less distance they will move.

## III Phases of Matter

The three phases of matter (solid, liquid, and gas) can be recognized by characteristics of the substances such as volume, shape and molecule spacing. Solids have a definite shape and volume and the molecules are closely packed. Liquids have no definite shape but do have a definite volume; their molecules are loosely packed. Gases have neither definite shape nor volume their molecules are very loosely packed.




## IV GASES

The ***Kinetic Molecular Theory*** explains the behavior of gases, it states:

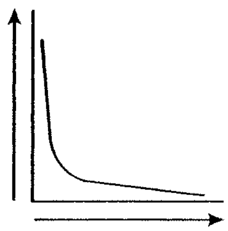
- gas molecules are in constant straight line motion
- the volume of gas molecules is negligible * assumption*
- gas molecules have no attractive forces * assumption*
- collisions are elastic
- pressure is caused by collisions

***Avogadro's Law*** also relates to gas behavior, it states

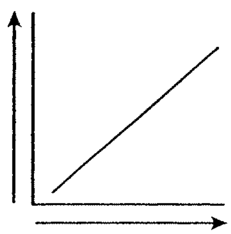
- Equal volumes of gases at the same temperatures and pressure contain an equal number of particles.

These two laws can be used to explain the relationships between pressure, temperature and volume of a gas.

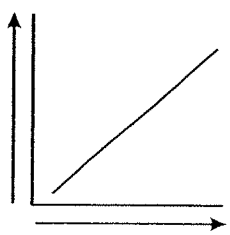
***Pressure/ Volume*** - indirect



***Temperature/Pressure*** - direct



***Temperature/Volume*** - direct



***Combined Gas Law***

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

example: A sample of gas has a volume of 12 liters at 0°C and 380 torr. What will be its volume when the pressure is changed to 760 torr at a constant temperature?

$P_1$ = 380 torr $P_2$ = 760 torr

$T_1$ = 0°C = 273 K $T_2$ = 0°C = 273 K

$V_1$ = 12 liters $V_2$ = X

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$\frac{380(12)}{273} = \frac{760(X)}{273}$$

X= 6 liters

## V SOLUTION

A solution is an example of a homogeneous mixture and consists of a solute (the substance being dissolved) in a solvent (the substance doing the dissolving) The solubility of a substance depends on the temperature, pressure and chemical nature of the solute and solvent.

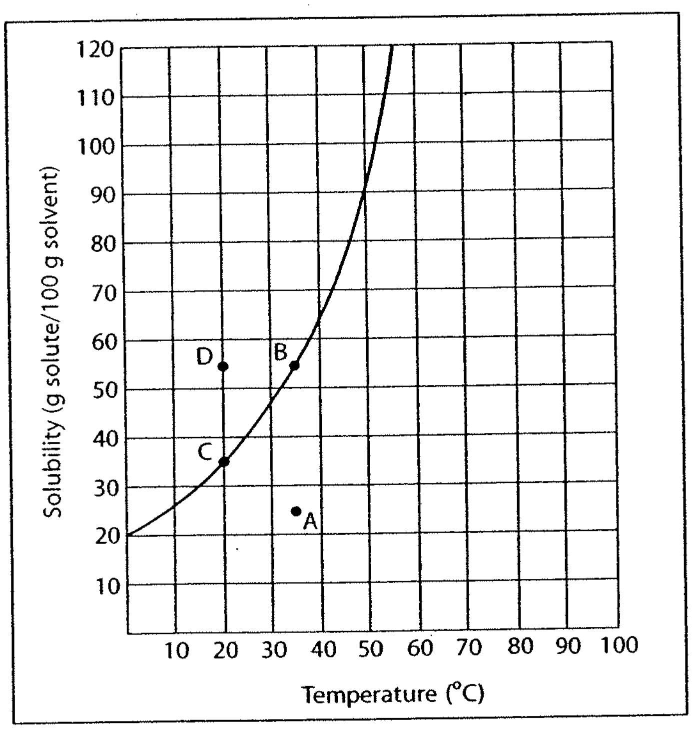
The general rule for solubility is polar dissolves polar nonpolar dissolves nonpolar or simply ***"Like dissolves Like"***. Reference Table F can be used to predict solubility.

calcium chloride- soluble, chlorides are and Ca is not an exception

sodium sulfide - soluble, sulfides are insoluble but sodium is an exception

Reference Table G you can determine how much of a substance can be dissolved in a given amount of water. In reading Table G you can determine if a solution is just saturated ( a point on the line), if a solution is super saturated( a point above the line with no residue ) or if the solution was unsaturated ( a point below the line)

Examples; A solution contains 14 g of KCl in 100.g of water at 40°C. What is the minimum amount of KCl that must be added to make this saturated solution? Need 39g have 14 g need 25 g more.




- A unsaturated
- B saturated
- C saturated
- D supersaturated

***Concentration*** : Concentration of solutions and be expressed in molarity (M), percent by volume, percent by mass or parts per million (ppm) Reference Table T has formulas for calculations.

Approximately .0043 of oxygen can be dissolved in 100 mL of water at 20°C.

examples:

Express this in terms of parts per million.

***ppm*** = $\frac{\text{grams of solute}}{\text{grams of solution}}$ X 1,000,000 ppm

$.0043/100.0043 \times 1{,}000{,}000 = 43$ ppm

What is the molarity of an $H_2SO_4$ solution if .25 L of the solution contains .75 mol of $H_2SO_4$ ?

***Molarity*** = Moles of solute/Liter of solution

.75 mol/.25 L = 3M

Effect of Solute on Solvent: The addition of a nonvolatile solute to a solvent causes the ***boiling point to increase and the freezing point to decrease***. The greater number of particles the greater the effect

VI ENERGY

Energy can exist in different forms (chemical, electrical, electromagnetic, heat, mechanical, and nuclear). Heat is a form of l energy. Heat always transfers from a body of higher energy to one of lower energy.

***Temperature*** is the measure of the average kinetic energy of a sample. It is not a form of energy but a measure of the energy. As the kinetic energy increases the temperature increases.

***Heat Calculations:***

The amount of heat involved in a reaction can be calculated using the formula on Table T; $q=mc\Delta T$

Example: How many joules are absorbed when 50.0 g of water is heated from 30.2 °C to 58.6 °C?

$q= (50.0\ g)(4.18 J/g \bullet °C) (58.6\ °C - 30.2\ °C)$

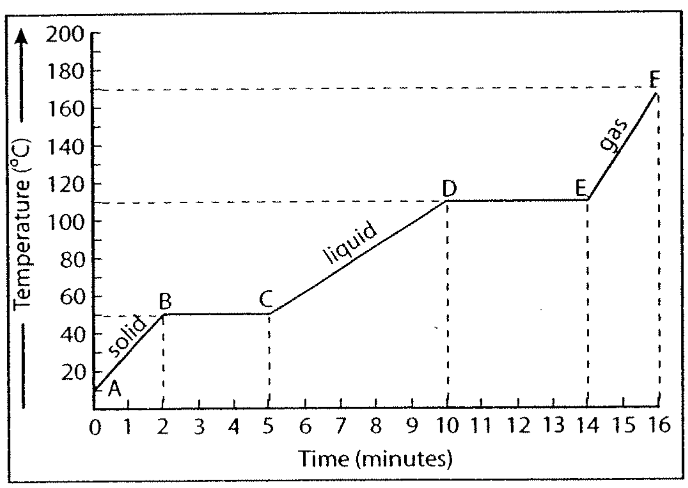
$q= 5936\ J = 5.94\ kJ$

Table I shows the heat given off (exothermic) or absorbed (endothermic) for a variety of reactions.

VII Phase Changes

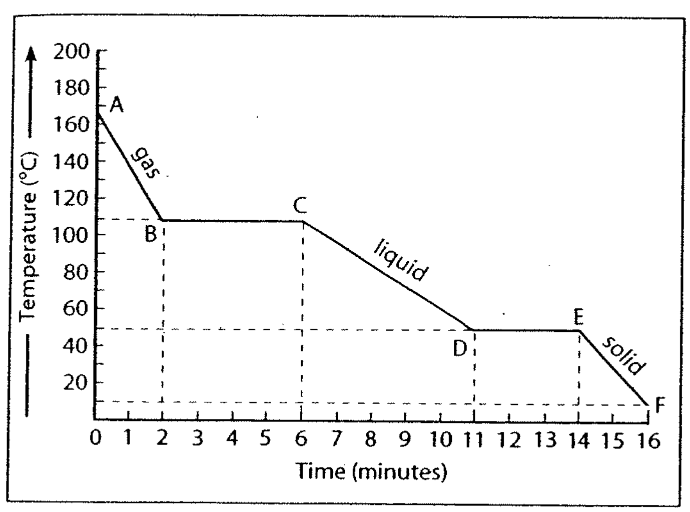
A. Heating Curve

The diagram below shows a typical heating curve. Heating curves are endothermic because heat must be added to change form s→ l → g. ***Fusion*** is the technical term for melting; ***vaporization*** describes the process of converting a liquid into a gas. During a phase change there is no cahnge in kinetic energy only potential energy, There is no temperature cahnge. The melting point is the first plateau and the boiling point is the second plateau. ***Boiling*** is special because it can occur at any temperature as long as the vapor pressure equals the atmospheric pressure. Therefore according to table H water boils at 100 ° C at standard pressure but at 50 kPa it would boil at about 82 °C. Sublimation is going directly from the solid to liquid phase without becoming a liquid.




E. Cooling Curve

The diagram below shows a typical cooling curve. Cooling curves are exothermic because heat is released as a substance from g→ l → s. ***Condensation*** is going from a gas to a liquid. ***Deposition*** is going from a gas to a solid.



***Energy Calculations*** during Phase Changes:

Vaporization: $q = mH_V$ Heat of vaporization = 2260 J/g

Table T Table B

How many joules of energy are required to vaporize 423 g of water at 100 °C and 1 atm?

$q= (423\ g)(2260\ j/g)$

$q = 9555{,}980\ J = 956\ kJ$

Fusion: $q = mH_f$ Heat of Fusion = 334 J/g

Table T Table B

How many joules are required to melt 255 g of ice at 0 °C?

$g= (255\ g)(334\ J/g)$

$q= 85{,}170\ J = 85.2\ kJ$

## VIII ***Significant Figures***

Rules

1. All non zeros are significant. 33 7, 3 significant figures
2. All zeros between significant numbers are significant 303, 3 significant figures.
3. All zeros at the end of the number and after the decimal place are significant. 3.400, 4 significant figures
4. Zeros that are placeholders are not significant. 400, 1 significant figure

Addition/Subtraction

Keep the least number of decimal places.

Multiplication/Division

Keep the least number of significant figures.

Examples:

23.7 + 34.342 + 18.10 = 76.142 rounded to proper significant figures 76.1 since 23.7 has only one decimal place.

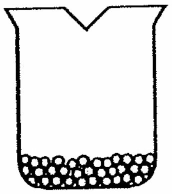
23.7 x .0032 ÷ 120 = .000632, rounded to significant figures is .00063 since both .0032 and 120 have only two significant figures.

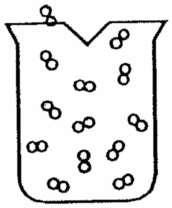
# QUESTIONS

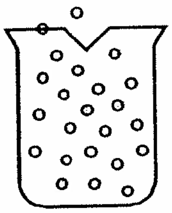


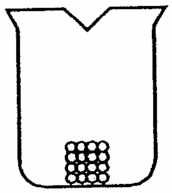
Name: ______________________________

1) Which particle diagram *best* represents a diatomic gas?

A) 

C) 

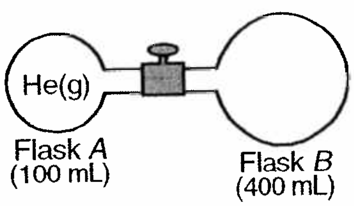
B) 

D) 

2) A 1-gram sample of which substance in a sealed 1-liter container will occupy the container completely and uniformly?

A) $Hg(\ell)$ C) $H_2O(g)$
B) $H_2O(\ell)$ D) $Ag(s)$

3) The diagram below shows two flasks connected by a stopcock. Flask *A* contains helium gas. Flask *B* contains a vacuum.




What final volume will the gas occupy after the stopcock is opened?

A) 500 mL C) 400 mL
B) 100 mL D) 300 mL

4) Under the same conditions of temperature and pressure, a liquid differs from a gas because the particles of the liquid

A) take the shape of the container they occupy
B) have stronger forces of attraction between them
C) are in constant straight-line motion
D) have no regular arrangement

5) Which sample contains particles arranged in a regular geometric pattern?

A) $CO_2(\ell)$ C) $CO_2(aq)$
B) $CO_2(g)$ D) $CO_2(s)$

6) At what point do a liquid and a solid exist at equilibrium?

A) vaporization point
B) boiling point
C) sublimation point
D) melting point

7) What term represents the change of a substance from the solid phase to the liquid phase?

A) evaporation C) condensation
B) vaporization D) fusion

8) A mixture of ice and water is in equilibrium at standard pressure. The temperature of the mixture must be

A) 212°C C) 273°C
B) 0°C D) 100°C

9) What conditions of pressure and temperature exist when ice melts at its normal melting point?

A) 1 atm and 0°C
B) 760 atm and 273 K
C) 760 atm and 0°C
D) 1 atm and 0 K

10) At 1 atmosphere, which substance will sublime when heated?

A) $CH_4(g)$ C) $HCl(aq)$
B) $H_2O(\ell)$ D) $CO_2(s)$

11) What process occurs when dry ice, $CO_2(g)$, is changed into $CO_2(s)$?

A) deposition C) condensation
B) fusion D) vaporization

12) The boiling point of water at standard pressure is

A) 373 K C) 100. K
B) 273 K D) 0 K

13) A gas is *most* likely to change to the liquid phase when the pressure on the gas

A) increases and its temperature decreases
B) decreases and its temperature decreases
C) increases and its temperature increases
D) decreases and its temperature increases

14) As the temperature of liquid water decreases, its vapor pressure

A) increases
B) decreases
C) remains the same

15) What is the first phase change that is *most* likely to occur as the pressure on nitrogen gas is increased and its temperature is decreased?

A) evaporation
B) condensation
C) solidification
D) crystallization

16) According to the *Vapor Pressure of Four Liquids* chemistry reference table, what is the vapor pressure of water at 105°C?

A) 68 kPa
B) 120 kPa
C) 105 kPa
D) 110 kPa

17) Which liquid has the *highest* vapor pressure at standard temperature?

A) ethanoic acid
B) ethanol
C) propanone
D) water

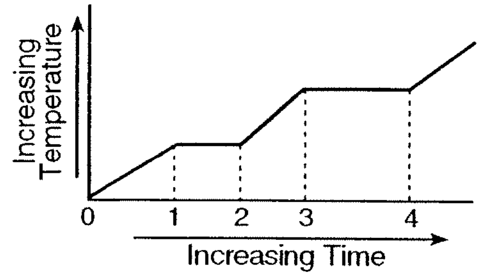
18) Which phase change is endothermic?

A) gas to solid
B) liquid to solid
C) liquid to gas
D) gas to liquid

19) Which change results in a release of energy?

A) the melting of $H_2O(s)$
B) the condensation of $H_2O(g)$
C) the boiling of $H_2O(\ell)$
D) the evaporation of $H_2O(\ell)$

20) The graph below represents the relationship between the temperature and time for a substance that was heated uniformly starting at $t_0$. The substance was in the solid phase at $t_0$.




During what time interval does the heat absorbed by the substance represent the heat of fusion of the substance?

A) $t_3$ to $t_4$
B) $t_1$ to $t_2$
C) $t_2$ to $t_3$
D) $t_0$ to $t_1$

21) The temperature of a substance is a measure of its particles'

A) average potential energy
B) entropy
C) enthalpy
D) average kinetic energy

22) In which beaker would the particles have the *highest* average kinetic energy?

A) 

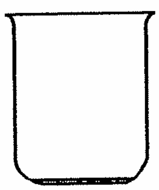
400 milliliters
0.1 M HCl at 15°C

B) 

50 milliliters
0.1 M HCl at 10°C

C) 

200 milliliters
0.1 M HCl at 30°C

D) 

10 milliliters
0.1 M HCl at 20°C

23) Two pure water samples held in separate containers at 1 atmosphere pressure must have molecules possessing the same average kinetic energy if the samples have the same

A) temperature
B) density
C) mass
D) volume

24) At 1 atmosphere of pressure, the fixed temperature points on a Celsius thermometer are located on the basis of

A) the ice/water equilibrium temperature, only
B) both the ice/water and the water/steam equilibrium temperatures
C) neither the ice/water nor the water/steam equilibrium temperatures
D) the water/steam equilibrium temperature, only

25) Human body temperature is 37° Celsius. What temperature does this correspond to on the Kelvin scale?

A) -236 K
B) 310 K
C) 98.6 K
D) 236 K

26) What Kelvin temperature is equal to -33°C?

A) 33 K
B) 240 K
C) -33 K
D) 306 K

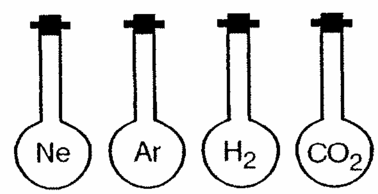
27) If 4 grams of water at 1°C absorbs 8 calories of heat, the temperature of the water will change by

A) 1°C
B) 2°C
C) 3°C
D) 4°C

28) Equal volumes of all gases at the same temperature and pressure contain an equal number of

A) protons
B) molecules
C) atoms
D) electrons

29) The diagrams below represent four 500-milliliter flasks. Each flask contains the gas represented by its symbol. All gas samples are at STP.



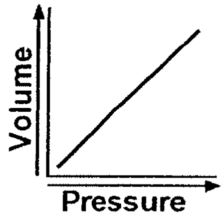

Each flask contains the same number of

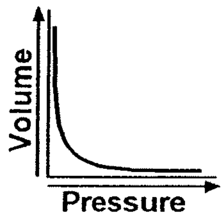
A) atoms, only
B) molecules, only
C) atoms and molecules
D) atoms, but different numbers of molecules

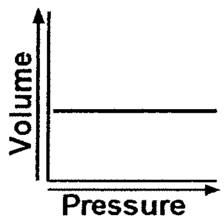
30) As the pressure of a gas at 101.3 kPa is changed to 50.65 kPa at constant temperature, the volume of the gas

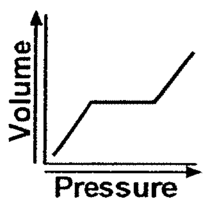
A) remains the same
B) decreases
C) increases

31) Which graph *best* shows the change in the volume of 1 mole of nitrogen gas as pressure increases and temperature remains constant?

A) 


B) 

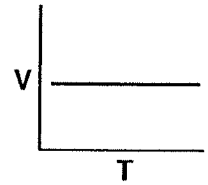

C) 

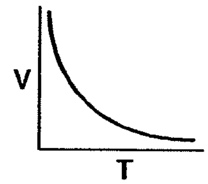

D) 

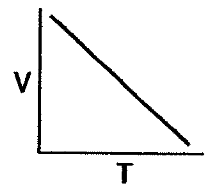

32) The volume of a given mass of an ideal gas at constant pressure is

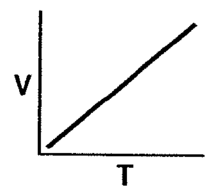
A) inversely proportional to the Kelvin temperature
B) directly proportional to the Celsius temperature
C) directly proportional to the Kelvin temperature
D) inversely proportional to the Celsius temperature

33) At constant pressure, which graph shows the correct relationship between the volume of a gas (*V*) and its absolute temperature (*T*)?

A) 


B) 


C) 


D) 


34) A 100.-milliliter sample of helium gas is placed in a sealed container of fixed volume. As the temperature of the confined gas increases from 10.°C to 30.°C, the internal pressure

A) remains the same
B) decreases
C) increases

35) Which gas under high pressure and low temperature has a behavior *closest* to that of an ideal gas?

A) $O_2(g)$
B) $H_2(g)$
C) $CO_2(g)$
D) $NH_3(g)$

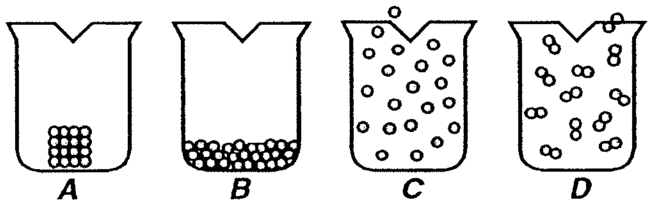
36) The volume of a 1.00-mole sample of an ideal gas will decrease when the

A) pressure decreases and the temperature increases
B) pressure increases and the temperature decreases
C) pressure increases and the temperature increases
D) pressure decreases and the temperature decreases

37) A 0.500-mole sample of a gas has a volume of 11.2 liters at 273 K. What is the pressure of the gas?

A) 0.500 atm
C) 1.00 atm
B) 273 atm
D) 11.2 atm

38) The particle diagrams below represent elements at STP.




Which particle diagram *best* represents a monoatomic gas?

A) *A* B) *B* C) *C* D) *D*

39) The table below shows the temperature, pressure, and volume of five samples.

| Sample | Substance | Temperature (K) | Pressure (atm) | Volume (L) |
|---|---|---|---|---|
| *A* | He | 273 | 1 | 22.4 |
| *B* | $O_2$ | 273 | 1 | 22.4 |
| *C* | Ne | 273 | 2 | 22.4 |
| *D* | $N_2$ | 546 | 2 | 44.8 |
| *E* | Ar | 546 | 2 | 44.8 |

Which sample contains the same number of molecules as sample *A*?

A) *E* B) *B* C) *C* D) *D*

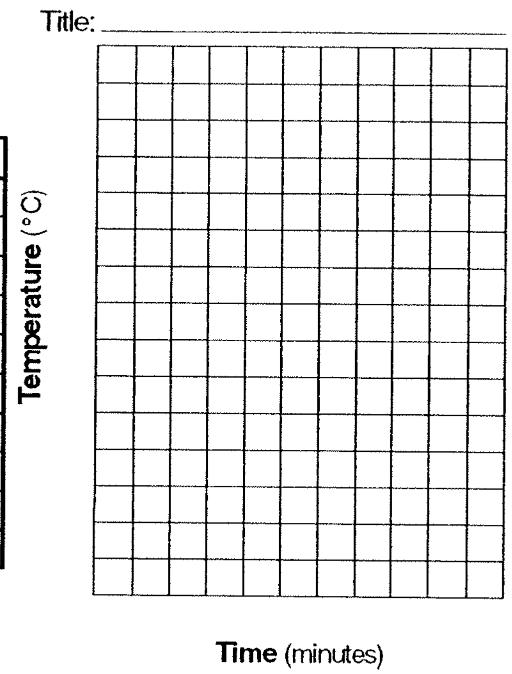
Constructed Response

Questions 40 and 41 refer to the following:

A student performed a laboratory investigation to determine the melting point of solid p-Dichlorobenzene (moth crystals). The student filled a test tube with solid p-Dichlorobenzene then gently heated the test tube in a beaker of boiling water. As the student heated the p-Dichlorobenzene, temperature readings were taken every half-minute. The following data was collected.

DATA TABLE

| **Time** (minutes) | **Temperature** (°C) |
|---|---|
| 0.0 | 20.0 |
| 0.5 | 35.0 |
| 1.0 | 50.0 |
| 1.5 | 53.0 |
| 2.0 | 53.0 |
| 2.5 | 53.0 |
| 3.0 | 53.5 |
| 3.5 | 55.0 |
| 4.0 | 60.0 |
| 4.5 | 70.0 |




40) (a) Construct a line graph on the grid above according to the following directions:

(1) Mark an appropriate scale on the axis labeled "Temperature (°C)".

(2) Mark an appropriate scale on the axis labeled "Time (minutes)".

(3) Plot the data from the data table. Surround each point with a small circle and connect the points.

**EXAMPLE:** 

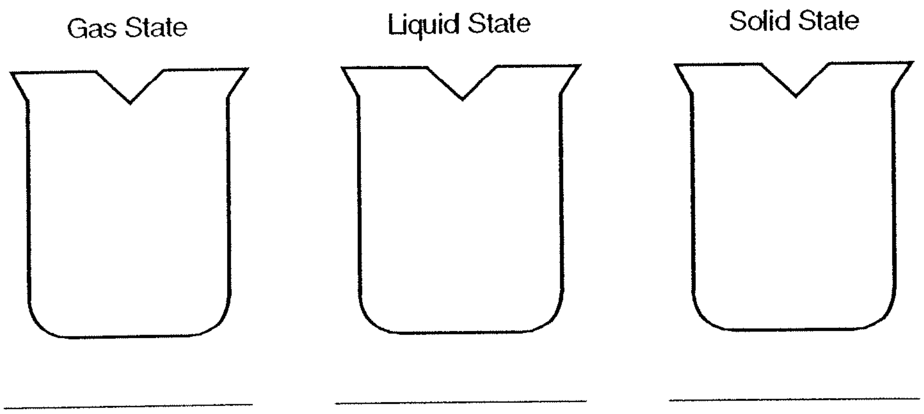
(b) Write an appropriate title on the graph.

(c) From the data table and the graph, determine the melting point of p-Dichlorobenzene.

41) Explain the plateau in temperature between 1.5 and 3.0 minutes in terms of kinetic and potential energy.

Questions 42 through 44 refer to the following:

The Halogen Family (Group 17) contains elements that exist in the gas, liquid, and solid states at room temperature and standard pressure.




42) (a) Using  to represent one molecule of a halogen, draw a particle diagram to represent a halogen in the gas, liquid, and solid states in the containers above. [*Each diagram should contain at least 6 halogen molecules*.]

(b) On the line below each container, identify (by name and\or symbol) a halogen which exists in the corresponding phase at room temperature and standard pressure.

43) Describe the difference between a halogen in the gas, liquid, and solid state in terms of particle arrangement.

44) Explain, in terms of intermolecular forces of attraction, why different halogens under the same conditions of temperature and pressure can exist in three different phases.

45) When a chemical cold pack is activated, a chemical reaction occurs that causes the pack to feel cold.

(a) Is the reaction that takes place inside a chemical cold pack exothermic or endothermic?

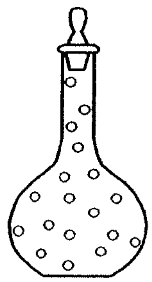
(b) Describe the transfer of energy between a cold pack and the air around it.

46) Jacques Charles (1746-1823) was a French physicist and hot air balloonist. He is credited with stating the relationship between the temperature and volume of a gas.

(a) The air inside a hot air balloon is heated by a gas burner. What effect does increasing the temperature have on the volume of the gas inside the balloon?

(b) As a hot air balloon rises to higher altitudes, atmospheric pressure against the balloon decreases. What effect does the decrease in pressure have on the volume of the gas inside the balloon?

47) The particle diagram below represents a sample of a gas sealed in a 1.0 liter flask. The sample was heated gently and the gas pressure was measured over a range of temperatures as reported in the data table.



| Temperature (K) | Pressure (kPa) |
|---|---|
| 300 | 101.3 |
| 310 | 104.7 |
| 320 | 108.1 |
| 330 | 111.4 |

(a) State the relationship between the temperature and pressure of the gas when the volume remains constant.

(b) Explain the increase in pressure of the sample at higher temperatures in terms of kinetic energy and collisions of the gas particles.

(c) Predict the gas pressure at 340 K. [*Write the correct formula. Show all work. Indicate the correct answer with an appropriate unit.*]

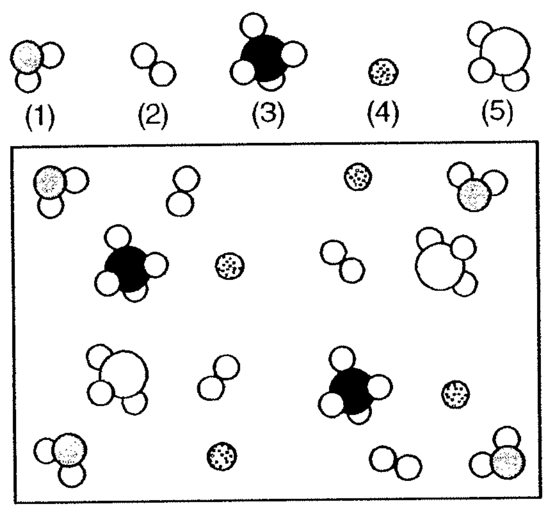
48) Ammonia ($NH_3$) is commonly used in the manufacture of nitric acid, synthetic fibers, and fertilizers. It is marketed as $NH_3(\ell)$ in steel containers or as $NH_3(aq)$ in bottles or drums.

(a) What is the difference in composition between a sample labeled $NH_3(\ell)$ and a sample labeled $NH_3(aq)$?

(b) In the boxes below, draw a particle diagram to support your answer to *part (a)*.

| $NH_3(\ell)$ | $NH_3(aq)$ |
| --- | --- |

49) The diagram below represents a gaseous mixture of the substances labeled 1 through 5.




(a) Which of the five substances are compounds? [*Explain why*.]

(b) Which of the five substances are elements? [*Explain why*.]

(c) Does the diagram represent a homogeneous or a heterogeneous mixture? [*Give one reason to justify your answer*.]