Matter can be classified using physical characteristics like, color, size and shape.

Matter is broken down into two main groups – pure substances and mixtures.

A pure substance is either an element or a compound.

A substance composed of a single kind of atom is called an element. The atoms may be bonded together into molecules or crystalline solids. A compound is formed when two or more kinds of atoms bind together chemically. Bonds between atoms are created when electrons are paired up by being transferred or shared.

The simplest component of an element is an atom whereas the simplest component of a compound is a molecule.

A mixture is made up of two or more substances that can be separated by physical means.

Mixtures are divided into homogeneous and heterogeneous categories.

Homogeneous mixture contains two or more gaseous, liquid, or solid substances blended evenly throughout.

Heterogeneous mixture contains materials that are easily distinguished from one another.

Physical Changes include changes that occur in the physical form without changing the composition of the substance.

Physical properties include boiling/condensation point, freezing/melting point, density, solubility, viscosity, and conductivity.

Boiling point is the temperature at which the pressure of the atmosphere is equal to the pressure of at liquid’s vapor, and gas molecules can escape the attractive force between the molecules.

Melting point is the temperature at which a solid begins to liquefy.

Freezing point is the temperature at which liquid begins to change to a solid.

Density is the mass per unit volume.

Solubility is the maximum amount of a solute that can be dissolved in a given amount of solvent at a given temperature.

The law of conservation of mass states that matter is neither created nor destroyed during a chemical change.

Chemical changes produce new substances and can be detected by a release of energy like heat, light or sound, a change in color, an odor given off or the formation of gas bubbles.

Rust is formed when oxygen in moist air reacts with iron to form Iron (III) oxide.

Solids have a definite shape and definite, measurable volume. The particles are in a fixed position, vibrate and possess a relatively small amount of kinetic energy compared to a liquid or gas.

- Liquids have no definite shape. They take on the shape of their container and have a definite, measurable volume. Particles of a liquid move around freely, but possess some attractive forces that keep the particles from escaping (except at the surface through evaporation). Particles of a liquid possess more kinetic energy than a solid and less than a gas.
- Gas particles have no definite shape or volume, have little attractive forces and possess the highest kinetic energy.

As a gas is heated, the gas molecules start moving faster and faster (kinetic theory of matter).

Conversely, as a gas cools, the molecules slow down.

Jacques Charles (1746-1823) was a French scientist who studied various properties of gases.

According Charles’ Law, if the pressure on a gas remains constant, the volume of a gas increases (decreases) proportionally to any increase (decrease) in Kelvin temperature \( \frac{V_1}{T_1} = \frac{V_2}{T_2} \).

Robert Boyle (1627-1691) discovered that gases completely fill their containers, therefore as the volume of a container increases, the pressure the gas exerts decreases.

Boyle’s Law states that if temperature remains constant, as the volume of a gas increases (decreases) the pressure on the gas decreases (increases) proportionally.

- Gases under pressure require special handling, due to their potential to explode. Many gases are flammable; therefore precautions must be taken to keep flames away from them.
- Under pressure, leaks in a container can be catastrophic if the pressure is too great (similar to decompression in an airplane). Gas containers, must be stored so that the chance of openings in the container are minimized.
- As heat is added to an object, the molecules in the object start moving faster (kinetic theory of matter).
- When particles start moving faster, they collide more often with each other, causing them to separate further. The faster a particle moves, the more kinetic energy the particles have.
- Temperature is a measurement of the average kinetic energy in an object. The slower the particles move, the less kinetic energy they have, the less they hit each other causing a release of heat, the lower the temperature of the object.
• Once the velocity of the particles reaches a certain point (temperature), a change is state occurs. 
• For example, when the average kinetic energy of water particles measures 0°C, liquid water becomes ice. When the liquid reaches 100°C, the water turns to water vapor. 
• Heat of fusion is the amount of energy required to change a substance from the solid phase to the liquid phase. 
• Heat of vaporization is the amount of energy required for liquid particles to escape the attractive forces within the liquid or the energy required to change from a liquid to a gas. 
• Heating curve of water shows the temperature change of water as thermal energy or heat is added. 
• Viscosity is the resistance of a fluid to flow. (liquids and gases) 
 • The component particles of the atom are protons, neutrons and electrons. Protons and neutrons have approximately the same mass and are found in the nucleus located in the center of the atom. This is where the mass of the atom is concentrated. Protons have a positive charge and neutrons are neutral. Electrons occupy regions around the outside of the nucleus called electron clouds (energy levels), are negatively charged, and have very little mass. The mass of the electron is insignificant when considering the mass of an atom. 
• Isotopes are atoms of the same element (same number of protons) with different numbers of neutrons. Isotopes are expressed in two basic ways. One with the element’s symbol and another with the element’s name. Two classic examples are carbon-12, carbon-13, & and carbon-14 and Hydrogen-1, Hydrogen-2, & Hydrogen-3. Some isotopes have unstable nuclei and are thus radioactive. 
• Carbon is a common example of isotopes. Below are the three isotopes of carbon expressed in both forms. 
\[
\begin{align*}
12^6C &= \text{Carbon } -12 \\
13^6C &= \text{Carbon } -13 \\
14^6C &= \text{Carbon } -14
\end{align*}
\]
• The elements in the periodic table are arranged like words in a paragraph of text. The elements are arranged from left to right and when reaching the end of a row begin a new row. As this process is repeated the elements in the same group or family have very similar or predictable physical and chemical properties, such as reactivity, density, hardness, state of matter, and types of compounds formed. 
• Property trends of elements going down a group or family include increase in atomic size, increase in boiling/melting points, increase in density, reactivity changes, decrease of electronegativity, increasing energy level (shielding effect), etc. 
• Property trends of elements going across a period include decrease in atomic size, metallic/nonmetallic properties, reactivity, increase in electronegativity, etc. 
• The alkali metals are the Group I metals which are shiny, malleable, and ductile. They are softer and the most reactive of the metals. They react sometimes violently with oxygen and water. They are stored in substances that are unreactive such as oil. Francium, the last element is rare and radioactive and is found in uranium minerals. Each atom has 1 valence electron and becomes a positively charge ion. 
• The alkaline earth metals are the Group 2 metals. They are shiny, malleable, and ductile and combine readily with other elements in nature. These electrons are given up when an alkaline earth metal combines with a nonmetal. 
• The transition metals are those elements Groups 3 through 12 in the periodic table because they are considered to be a transition between groups 1and 2 and 13 through 18. They often form colored compounds and some often occur in nature uncombined so they are familiar compounds. 
• The inner transition metals are the two rows that are disconnected from the periodic table. They fit between Groups 3 and 4 and 6 and 7. To save room they are listed below the periodic table. The lanthanides are elements 58 through 71 and the actinides are 90 to 103 and are radioactive and unstable. 
• Hydrogen is in Group 1 but is a nonmetal. It is very reactive and has a single electron. It shares electrons with other nonmetals but gains an electron when it combines with alkali and alkaline earth metals. 
• The halogens are Group 17 and are all nonmetals. They are very reactive in elemental form. Halogens have 7 outer electrons and tend to gain one electron. Hydrogen, the halogens, oxygen, and nitrogen form diatomic molecules when two atoms of the same element form a covalent bond. 
• Groups 13 through 16 are mixed groups of metals, nonmetals, and metalloids. 
• Group 18 compose the noble gases which are chemically stable because their outmost energy levels are full. 
• All synthetic elements are short lived except the transuranium which are elements with atomic numbers greater than 92. 
• The metalloids are located along the stair step line with the exception of aluminum and can form ionic and covalent bonds.
• Ionic bonds are bonds that involve the transfer of electrons from one atom to another. Metals typically lose (give away) electrons to attain stability and nonmetals gain (take) electrons in order to become stable. Upon losing one or more electron(s) the metal attains a positive charge. The nonmetal on the other hand attains a negative charge upon gaining one or more electrons. The opposite charges of the metal and nonmetal cause a very strong force of attraction called ionic bonding.
  • Covalent bonds form by sharing electrons. This typically involves nonmetals. One, two, or three pairs of electrons are shared between nonmetals to attain a noble gas electron configuration (8 valence electrons). These bonds are typically weaker than ionic bonds.
  • Examples of ionic compounds are metal-nonmetal compounds or compounds containing polyatomic ion (NaCl, CaCO₃, NH₄NO₃).
  • Examples of covalent compounds are those which contain only nonmetals (CO₂, H₂O, C₆H₁₂O₁₁).
  • An oxidation number is the number of electrons that an atom gains, loses or shares. For ionic substances this is the same as the charge of the ion. Compounds are neutral; therefore the sum of the oxidation numbers in a compound is equal to zero. The symbol and number of each element in the compound is represented in the chemical formula, which contains the symbols of the elements and subscripts to indicate the ratio of the elements in the compound.
  • Properties of ionic substances include high melting and boiling points, conductivity in molten or dissolved state, crystalline and soluble in water.
  • Properties of covalent compounds include low melting and boiling points, low or no conductivity in molten or dissolved state and noncrystalline. Polar covalent compounds are soluble in water, whereas nonpolar covalent compounds are insoluble in water.
  • Polar substances include ionic compounds and covalent compounds that share electrons unequally resulting in polarity. Nonpolar substances are covalent compounds whose members share electrons equally and do not exhibit polarity.
  • Polar solutes dissolve in polar solvents and nonpolar solutes dissolve in nonpolar solvents. Polar and nonpolar substances do not mix.
• Carbon forms a tremendous number of compounds with itself, often in long branched chains of connected carbon atoms with attached hydrogen. As carbon bonds to another carbon it has four bonding sites. It may form a single covalent bond (sharing of one pair of electrons), a double covalent bond (sharing of two pairs of electrons) or a triple covalent bond (sharing of three pairs of electrons).
  • Saturated hydrocarbons are carbon-hydrogen compounds with containing only single bonds.
  • Unsaturated hydrocarbons are carbon-hydrogen compounds that contain double or triple bonds.
  • Polymers are large molecules of repeating units (called monomers) connected together.
  • Common synthetic polymers include polyethylene (shopping bags and plastic bottles), polypropylene (glues and carpet) and polystyrene (CD cases and Styrofoam).
  • Common biological polymers include protein (amino acid monomers); nucleic acid (nucleotide monomers); and starches (glucose monomers).
  • Proteins are large organic polymers formed from organic monomers called amino acids. Proteins come in numerous forms and make up many of the tissues in the organisms including muscles, hair and fingernails. The identity of a protein depends upon the exact sequence of amino acids. Cells of an organism manufacture proteins by assembling the appropriate sequence of amino acids.
  • Carbohydrates such as starches are polymers composed of monomer units called simple sugars (such as glucose). When starches break down during digestion, this releases the simple sugars that your body converts to usable energy.
  • Lipids are polymers that contain monomers called fatty acids. These are much more difficult for the body to break down into usable forms of energy and are often stored in the body.
• An iron atom loses three electrons therefore attaining a +3 charge. An oxygen atom gains two electrons attaining a –2 charge. When iron is exposed to air a chemical reaction takes place and a new substance is formed. Chemical changes require that a new substance is formed. The new substance is rust or iron oxide. In order for the chemical reaction to take place there must be a certain number of units of iron and a certain number of units of oxygen to form the stable compound rust. The following equation shows the number of units of each to form 2 units of iron oxide. This is a good example of a chemical reaction and of a balanced equation that is very common. The steps to balancing equations involve writing the word equation correctly, then writing the formulas correctly, and then by balancing each element one at a time using a number called a coefficient. These are the numbers that are written in front of each substance.
  • Word Equation; Iron plus oxygen produces iron oxide.
  • The balanced synthesis reaction of the rusting of iron is as follows: \[ 4\text{Fe} + 3\text{O}_2 \rightarrow 2\text{Fe}_2\text{O}_3 \]
  • Prohibiting the exposure of iron to oxygen may prevent rusting. This may be done by maintaining an oxygen-free environment or by coating the iron with a substance to keep the oxygen from reacting with the iron.
  • Rusting has many negative impacts on society. The rusting of iron destroys the hardness, shininess,
• Another example of a chemical reaction is the digestion of food. During digestion food is broken down into a smaller, more usable form that the body may either use or store. (e.g. amylase breaks complex carbohydrates into simple sugars).

• Simple sugars such as glucose is assimilated and stored in organisms in the form of glycogen in animals and in the cell walls of plants. Proteins are made of simple units (monomers) called amino acids. The exact sequence of amino acids and the shape of the molecule identify the protein. Cells of an organism assemble proteins by assembling the amino acids in the proper sequence through the code found on DNA.

• Evidence of a chemical reaction (or change) includes color change, production of a gas, energy released as light and heat and production of a precipitate. Words such as grow, ripen, spoil, decay and burn are words that describe everyday chemical reactions that occur in our lives.

• The beginning material(s) in a chemical reaction is/are called the reactant(s). The changed or new material(s) in the chemical reaction is/are called the product(s). The reactants are separated from the products by an arrow called the yield sign (→). The substances are represented by their symbols. If there is more than one reactant or product, they are separated by a plus sign (+).

• The Law of Conservation of Matter states that matter cannot be created nor destroy but is conserved in any change. Therefore in a chemical reaction the mass of the reactant (actual mass or number of atoms) is equal to the mass of the products.

• In order to obey this law in a chemical reaction, coefficients must be provided to show that the number of atoms of reactant is equal to the number of atoms of product. A coefficient is a large whole number placed in front of a substance (reactant or product) in order to balance the total number of atoms of reactant with the total number of atoms of product. Refer to II.C.1.a to see an example of the balanced equation for the rusting of iron.

• Chemical equations may be classified by type.

• A synthesis reaction (also called combination or composition reaction) is a reaction with only one product. The rusting of iron is an example of this.

• A decomposition reaction is a reaction in which there is only one reactant. Peroxide decomposition is an example of this: 2H₂O₂ → 2H₂O + O₂

• A single replacement reaction is one in which one member of a compound is replaced by a single similar substance to produce a new compound and a new single substance. The reaction of aluminum and hydrochloric acid is an example of this: 2Al + 6HCl → 2AlCl₃ + 3H₂

• A double replacement reaction is one in which two compounds exchange like members. This is often compared to square dancing in which like partners exchange places. Many double replacement reactions occur in solutions and may produce a precipitate (an insoluble product). An example of this reaction is as follows: Pb(NO₃)₂ + 2 KI → 2KNO₃ + PbI₂ (yellow precipitate).

• A complete combustion reaction usually involves a hydrocarbon or carbohydrate that is burned to produce carbon dioxide and water.

• An incomplete combustion is a combustion reaction performed with limited oxygen producing carbon monoxide instead of carbon dioxide.

• Many reactions require an initial amount of energy to start (activate) a reaction. Once a reaction begins it will either release or consume (absorb) energy. If the reaction releases energy, it is classified as exothermic. If a reaction absorbs energy, it is classified as endothermic.

• The burning of isooctane is an example of an exothermic reaction. 2C₈H₁₈ + 25O₂ → 16CO₂ + 18H₂O + Energy

• Photosynthesis is an example of an endothermic reaction. 6CO₂ + 6H₂O + Light Energy → C₆H₁₂O₆ + 6O₂

• Decreasing the particle size by pulverizing/crushing/powdering will increase the surface area of the particles and thus increase the rate of reaction.

• Stirring or agitation will increase the rate of reaction.

• A higher concentration will increase the rate of reaction. A high concentration of oxygen increases the rate of combustion more than a lower concentration of oxygen.

• The presence of a catalyst will speed up a reaction. They are not involved in the reaction as reactants or products but simply speed up a reaction. Most solid catalysts, such as those in an exhaust system of a car, provide a large surface area where reactants can collect and react. Hydrogen peroxide decomposes very slowly naturally. In the presence of manganese dioxide, this normally slow reaction takes place spontaneously. 2H₂O₂ → 2H₂O + O₂

• Substances that slow down a reaction are called inhibitors.
• Increasing the temperature, increasing the surface area by grinding, powdering, etc, and stirring or agitating a mixture will increase the rate in which a substance dissolves.
  • The terms “concentrated” and “dilute” are relative, qualitative (rather than quantitative) terms. A “concentrated” solution is one that contains much solute contrasted with a “dilute” solution that contains relatively little solute.
  • The higher the concentration of an ionic solution, the higher the conductivity and the lower the concentration of an ionic solution, the lower the conductivity. Concentrated and dilute polar covalent solutions have little effect on the conductivity of a solution.
  • Colligative properties of a solution are those properties of a solution that are dependent upon the amount of dissolved solute. These include melting point depression and boiling point elevation.
  • Polar substances include ionic compounds and covalent compounds that share electrons unequally resulting in polarity. Nonpolar substances are covalent compounds whose members share electrons equally and do not exhibit polarity.
  • Polar solutes dissolve in polar solvents and nonpolar solutes dissolve in nonpolar solvents. Polar and nonpolar substances do not mix.

• The physical and chemical properties of acids include: sour taste, donate hydrogen (H\(^+\)) ions or hydronium (H\(_3\)O\(^+\)) ions when dissolved in water, turns blue litmus paper red, can burn the skin, and will react with many metals (corrosion).
  • The physical and chemical properties of bases include bitter taste, donate OH- ions when dissolved in water, turns red litmus paper blue, can damage skin, and feel slippery when dissolved in water.
  • Acids typically begin with a hydrogen atom, such as HCl, H\(_2\)SO\(_4\), HBr. (H\(_2\)O and H\(_2\)O\(_2\) are not acids.) Bases have an OH- group attached. The OH- group is typically attached to a metal or positive polyatomic ion (NH\(_3\)). This should not be confused with alcohols that contain –OH groups attached to a carbon-hydrogen molecule.
  • Acids react with metals to produce hydrogen gas that is explosive. An example of this reaction is as follows: 2Al + 6HCl → 2AlCl\(_3\) + 3H\(_2\)
  • pH values correspond to the concentration of hydronium ions (H\(_3\)O\(^+\)) that are produced when an acid is added to water. The lower the pH the more acidic the substance is. pH ranges from 0-14 with 0 being the most acidic/least basic and 14 being the most basic/least acidic. A pH of 7 indicates a neutral substance – neither acidic nor basic. A pH closer to 7 is considered a weak acid or base and a pH farther from 7 is considered a strong acid or base. Small changes in pH may translate into large changes in hydronium concentration. Each unit of pH represents a factor of 10 in hydronium ion concentration. For example, a substance with a pH of 4 is ten times more acidic than a substance with a pH of 5 and a hundred times more acidic than a substance with a pH of 6.
  • The pH of a substance may be determined by several methods. Litmus paper can give a general idea of pH as blue litmus paper turns red in an acid and red litmus paper turns blue in a base. Universal pH paper can be used to determine a more accurate pH. Indicators turn a variety of colors that may be matched to a chart on the side of the container. A pH meter is an electronic device that can determine exact pH values. These instruments can be very accurate, but must be carefully calibrated. CBLs with pH probes are also very accurate and versatile, but must also be carefully calibrated. Some other substances can change color to indicate pH. Phenolphthalein turns pink in a portion of the basic range. Red cabbage juice, grape juice, carrot juice and other common substances may change color to indicate pH.
  • Soaps are made from basic substances. Ammonia and bleach are both bases. fruits and fruit juices are acidic. Antacids such as milk of magnesia are basic to neutralize excess stomach acid. Cosmetics, shampoos and skin conditioners must be slightly acidic to neutral pH in order to match the pH of our skin. Many foods are acidic such as vinegar based foods, milk, and fruits. Drain cleaners are strongly basic.

Fission Reactions are the result of the splitting of the nucleus of an atom. Below is an example of this reaction:

\[
\begin{align*}
235 {}^92\text{U} + {}^1\text{n} & \rightarrow 142 {}^56\text{Ba} + 91 {}^3\text{Kr} + 3 {}^1\text{n} + \text{energy}
\end{align*}
\]
• Fusion reactions fuel the sun and stars. Fusion is the process in which light nuclei combine at extremely high temperatures, forming heavier nuclei and releasing energy. This multi-step process is illustrated below:

\[
\begin{align*}
&\text{H} + \text{H} \rightarrow \text{H} + \text{two particles} \\
&2 \text{H} + \text{H} \rightarrow 3 \text{He} + 0 \gamma \\
&3 \text{He} + 3 \text{He} \rightarrow 4 \text{He} + \text{H} + \text{H}
\end{align*}
\]

• Refer to related standard II.A.3.a. above.
• Fusion reactions fuel the sun and stars. Fusion is the process in which light nuclei combine at extremely high temperatures, forming heavier nuclei and releasing energy.

\[
\begin{align*}
&\text{H} + \text{H} \rightarrow \text{H} + \text{two particles} \\
&2 \text{H} + \text{H} \rightarrow 3 \text{He} + 0 \gamma \\
&3 \text{He} + 3 \text{He} \rightarrow 4 \text{He} + \text{H} + \text{H}
\end{align*}
\]